Experiment 6
Electromagnetic Radiation and Atom Interaction

B OBJECTIVES

To be familiar with the relationship between emission line spectra and the energy levels of electrons in various atoms.

B INTRODUCTION

Our understanding of atomic structure has come from studies of the properties of light (or radiant energy) and emission and absorption spectra. Radiant energy has two separate characteristics, that of a wave and that of a particle. As a wave, light can bend around corners (diffraction) and change direction through a lens (refraction). As a particle, light is a discrete entity called a photon, and can interact with other particles, like electrons. As a wave, light obeys the equation:

\[ c = \lambda \cdot \nu \]  

where \( c \) is the speed of light (3.00 \( \cdot \) 10\(^8\) m/s in vacuum), and \( \lambda \) (lambda) is the wavelength of the light in meters (although customary measurements are in nm; 10\(^{-9}\) m) and \( \nu \) (nu) is the frequency of the light in cycles per second (Hz).

A particular source of radiant energy may emit a single wavelength as in the light from a laser, or many different wavelengths, as in the radiation from an incandescent light bulb or a star. Radiation composed of a single wavelength is termed monochromatic. When the radiation from a source such as the Sun or other star, is separated into its components, a spectrum is produced. This separation can be achieved by passing the radiation through a prism. Each component of the polychromatic radiation is bent to a different extent by the prism, as shown in Figure 3.1. This rainbow of colors, containing light of all colors, is called a continuous spectrum. The most familiar example of a continuous spectrum is the rainbow, produced by the dispersion of sunlight by raindrops.
Albert Einstein theorized that energy is absorbed by atoms based on the frequency of light, not intensity. A minimum energy is needed according to the equation: 
\[ E = \lambda \cdot \nu \]
where, \( E \) is the energy of the photon, the constant \( h \) is Plank’s constant (6.63 x 10^{-34} \text{joule-seconds [J-s]}) and \( \nu \) is its frequency. Einstein used his equation to explain the photo-electric effect, a puzzle to scientists in the early 1900’s. Combined with equation [1], the energy of light is based on its wavelength or color.

\[ E = h \cdot \nu = \frac{hc}{\lambda} \]  \[2\]

Most substances will emit light energy if heated to a high enough temperature. For example, a fireplace poker will glow red if left in the flame of a fire long enough. Similarly, neon gas will emit bright red light when excited with a sufficient high electrical voltage. If you have ever spilled table salt [NaCl] into a flame, you have seen the characteristic yellow emission of excited sodium atoms.

Excited atoms do not emit a continuous spectrum, but rather emit radiation at only certain discrete, well-defined, fixed wavelengths. For example; if the light emitted by the atoms of a particular element is viewed in a spectroscope, only certain bright-colored lines are seen in the spectrum. This is called line spectra.

That atoms absorb and emit radiation with characteristic wavelengths was one of the observations that led the Danish physicist Niels Bohr to develop a model for a theoretical explanation of line spectra. The electron of the hydrogen atom moves about the central proton in a circular orbit. Only orbits of certain radii and having certain energies are allowed. In the absence of radiant energy, an electron in an atom remains indefinitely in one of the allowed energy states or orbits. When electromagnetic energy impinges upon the atom, the atom may absorb energy, and in the process an
electron will be promoted from one energy state to another. According to Bohr’s theory, if an electron were to move from an outer orbit to an inner orbit, a photon of light should be emitted having the energy:

$$\Delta E = E_{inner} - E_{outer} = -R_H \left( \frac{1}{n_{inner}^2} - \frac{1}{n_{outer}^2} \right)$$ \[3\]

In the Bohr model, the radius of the orbit is related to the principal quantum number, n:  

$$\text{radius} = n^2 \times (5.3 \times 10^{-11} \text{ m})$$

and the energy of the electron is also related to n. The constant $R_H$ is called the Rydberg constant; it has the value $2.18 \times 10^{-18} \text{ J}$. Thus, as n increases, the electron moves farther from the nucleus and its energy increases. Figure 3.2 shows Bohr's model depicting the changes in energy levels of an electron.

$$E_n = -R_H \left( \frac{1}{n^2} \right)$$ \[4\]

Figure 3.2

Electrons can be excited (absorb energy) to a higher energy level or move to a lower unoccupied energy level releasing energy in the form of photons (light).
An instrument used for studying line spectra is called a spectroscope. A spectroscope contains the following: a slit for admitting a narrow beam of light, which sharpens the images; a diffraction grating, which acts as a prism, for dispersing the light into its components; an eyepiece for viewing the spectrum; and an illuminated scale against which the spectrum wavelengths may be determined.

EXAMPLES

3.1 What is the frequency that corresponds to a wavelength of 500 nm? See equation [1]

\[ \nu = \frac{c}{\lambda} = \frac{3.00 \times 10^8 \text{ m/s} \cdot (10^9 \text{ m/nm})}{500 \text{ nm}} = 6.00 \cdot 10^{14} \text{ s}^{-1} \]

or: \( 6.00 \cdot 10^{14} \text{ Hz} \).

3.2 What is the energy of an electron when \( n = 3 \)? When \( n = 2 \)? See equation [4].

\[ E_3 = (-2.18 \cdot 10^{-18} \text{ J}) \cdot \left( \frac{1}{3^2} \right) = -2.42 \cdot 10^{-19} \text{ J} \]

\[ E_2 = (-2.18 \cdot 10^{-18} \text{ J}) \cdot \left( \frac{1}{2^2} \right) = -5.45 \cdot 10^{-19} \text{ J} \]

Thus an electron moving from \( n = 3 \) to \( n = 2 \) would emit light of energy:

\[ \Delta E = E_2 - E_3 = (-5.45 \cdot 10^{-19} \text{ J}) - (-2.42 \cdot 10^{-19} \text{ J}) = -3.20 \cdot 10^{-19} \text{ J} \]

The negative value means energy is released.
B  **PROCEDURE:**

Part A.- Energy levels for Hydrogen Gas.

Since the energy levels of each atom are different, the emission spectrum of an individual atom is unique.

From the class demonstration each student will record the color and wavelength band of the hydrogen emission on their lab report sheet.

Part B. Identification visible band in Helium and Mercury.

Students will be working in teams of two in the laboratory. Each student will observe one of the two gases while the other member of the team will record the color bands and wavelength emission of the gas on their lab report sheet. Team members will switch roles on the second gas.

¢  With the power supply turned off, position the slit opening of the spectroscope directly in front of the hydrogen lamp.

¢  DO NOT TOUCH ANY PORTION OF THE POWER SUPPLY, WIRE LEADS, OR LAMPS WHILE THE POWER SUPPLY IS TURNED ON.

¢  Ultraviolet radiation is damaging to your eyes. Wear your safety goggles at all times during this experiment, since they will absorb some of the ultraviolet radiation.

¢  Do not look directly at any of the lamps while they are illuminated.

¢  DO NOT let the power supply or lamp touch the spectroscope. With your TA permission, turn on the power supply switch to illuminate the gas lamp. Look into the eyepiece and adjust the slit opening so as to maximize the brightness and sharpness of the emission lines on the scale. Once adjusted do not move the spectroscope.
Part C. Flame test for some common metals.

A number of metals can be identified by their unique color of flame. Each team of students will record color of the flame for lithium, calcium, barium, sodium, strontium, and magnesium in their lab report sheet.

You will need a burner as an excitation source and a wire loop to pick up a drop of the metal-ion solution and then place the drop into the flame for vaporization. You may need to repeat this as it produces only a brief burst of color before the sample evaporates completely. Obtain a wire and about 10 mL of 6 M HCl. To clean the wire loop, dip the loop into the 6 M HCl then rinse the wire loop in distilled water and heat the loop in the hottest part of the flame until no color is imparted to the flame by the wire.

Once the wire loop is clean, dip it into a solution of metal-ion and heat the wire loop in the hottest part of the flame. Record the gross color of the flame. Repeat the above for each of the other metal-ion solutions. Remember to clean the wire loop for each new metal-ion.

You will find attached to this procedure, emission spectra of the metal-ions.

Part D. Identifying a mixture of two metals.

This part of the experiment can be done outside of the laboratory. From the emission spectrum of individual metals, found in back of this procedure, identify the composition of a mixture. Your TA will assign to you which mixtures you are to identify. Record the mixture assignments and your answers in your lab report sheet.


Line spectrums obtained from University of Oregon, Department of Physics webpage.
Part C  Flame Test Spectra

Lithium

Sodium

Magnesium

Calcium

Strontium

Barium
Part D  Known Metals: 1 - 5

Aluminum

Cadmium

Cobalt

Gold

Lead
Part D (cont) Unknown Metal Pairs

Unknown A10

Unknown B4

Unknown C20

Unknown D1

Unknown E16
Part D (cont) Unknown Metal Pairs

Unknown F17

Unknown G19

Unknown H2

Unknown J3

Unknown K24
Part D (cont) Unknown Metal Pairs

Unknown L11

Unknown M14

Unknown N7

Unknown O5

Unknown P21
Part D (cont) Unknown Metal Pairs

Unknown Q12

Unknown R23

Unknown S8

Unknown T9

Unknown U22
REPORT SHEET          EXPERIMENT  3

Electromagnetic Radiation and Atom Interaction

Part A.  Emission Spectrum of Hydrogen

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Part B.

Emission Spectrum of __________________________

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Emission Spectrum of __________________________

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Part C. Color of flame

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Part D. Identification of a mixture of two metals.

Unknown numbers ______ and ______

What two metals are in your mixtures?

Mixture 1 : _____________________
Mixture 2 : _____________________
Mixture 3 : _____________________

QUESTIONS:

1. What is the purpose of the slit in the spectroscope?

2. Why is the spectroscope scale illuminated?

3. a. Calculate $E$ (change in energy) of the visible lines of the hydrogen atom observed from the spectroscope using equation [2]. Show your set up.

3. b. Calculate the energy levels ($n_i$ to $n_f$) likely responsible for these $E$ ‘s, using equation [3]. Show your set-up.
4. In addition to the spectral lines that you observed in the emission spectrum of hydrogen, several other lines are also present in other regions of the spectrum. Calculate the wavelength of the \( n = 4 \rightarrow n = 1 \) and \( n = 4 \rightarrow n = 3 \) transitions and indicate in which regions of the spectrum these transitions would occur.

5. The minimum energy required to break the oxygen-oxygen bond in \( \text{O}_2 \) is 495 kJ/mole. What is the longest wavelength of radiation that possesses the necessary energy to break the \( \text{O} \rightarrow \text{O} \) bond? What type of electromagnetic radiation is this?

6. Compare your observation of the gross flame test with the line spectrum of each metal-ion. Do your observations correspond to the line spectrum? Explain your answer.