

Atomic Spectra & Electron Energy Levels

OBJECTIVES:

- To measure the wavelength of visible light emitted by excited atoms
- to calculate the energy of that emitted radiation
- to calculate, for hydrogen, the energy levels occupied by the excited electrons before they returned to lower energy levels.

DISCUSSION:

We have obtained much of the information concerning the arrangement of electrons within atoms by *spectroscopy*, the examination of the light absorbed or emitted when atoms undergo a change of energy state. Atoms may absorb or emit energy at many frequencies, corresponding to visible light, infrared, ultraviolet, or X-rays. In this lab exercise, you will observe the visible light emitted by excited atoms. Visible light constitutes only a very narrow range of the much wider electromagnetic spectrum. This spectrum also includes radio, microwaves, infrared, ultraviolet, X-radiation, and gamma radiation.

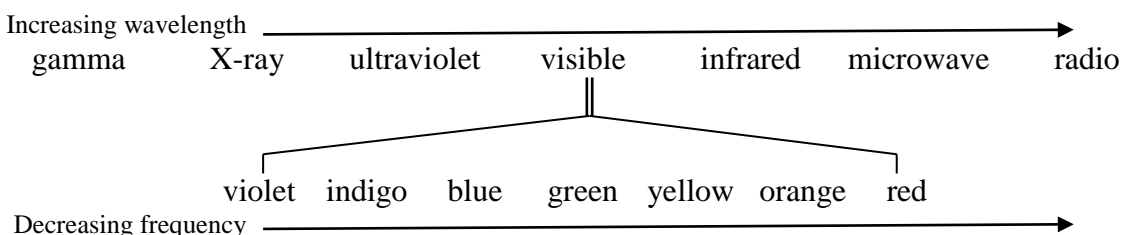


Figure 1: The electromagnetic spectrum

We describe electromagnetic radiation as an oscillating disturbance in the electric and magnetic fields, traveling through space at the speed of light, $c = 2.998 \times 10^8$ m/s. We can specify the oscillation using either the wavelength, λ (*lambda*) or the frequency, ν (*nu*). The *wavelength* gives the distance between equivalent points on sequential waves (see Figure 2) and *frequency* gives the number of waves per second that pass a reference point.

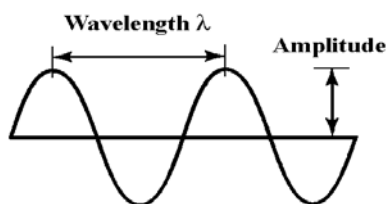


Figure 2: Wavelength and amplitude

Because light travels through space at a constant speed, the wavelength and frequency are inversely proportional:

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

As one increases, the other must decrease. Reconsider Figure 1 as you think about this relationship.

The energy of the light emitted by an atom is directly proportional to the light's frequency:

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

where Planck's constant, $h = 6.626 \times 10^{-34}$ J·s for each emitting atom.

How does the nature of light give information about the electronic structure of the atom? The energy absorbed by an atom boosts an electron to a higher energy state. As the electron returns to lower energy states it releases the energy it absorbed as electromagnetic radiation, though not necessarily all in one step. Some of this energy may lie in the X-ray, some in the ultraviolet, and some in the visible, infrared, or μ -wave range.

If the emitted visible light passes through a diffraction grating, the light path bends. The angle of the bend depends on the frequency (color) of the light, so in the lab you will observe several separated bright images of the atomic light source, each of a different color. The series of lines constitutes the atomic emission spectrum of that particular element. Each line represents a different frequency of radiation. Each frequency has associated with it a definite amount of energy, as given by Equation 2. The observation of discontinuous spectra for atoms of every element has led to the idea that electrons within an atom occupy definite energy levels. The spectra result from the radiation emitted by the excited electrons as they lose their excess energy and go from higher to lower energy levels, eventually returning to the unexcited ground state (see Figure 3). After measuring the frequency (or wavelength) of the emitted light, you can use Equation 2 to calculate the *energy differences* between the various energy levels. You will observe that no two elements have identical emission spectra. What does this mean about the energy levels in atoms of different elements?

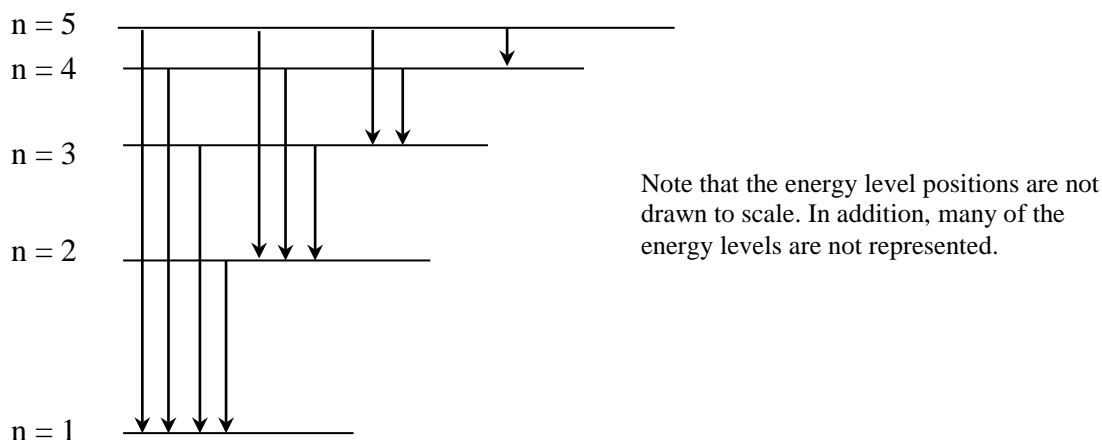


Figure 3: Representation of H-atom electron energy levels

We may calculate the H-atom energy level numbers using an equation empirically determined by Rydberg. Nils Bohr derived the same equation (3) theoretically in his early model of the electronic structure of the hydrogen atom.

$$\frac{1}{\lambda} = \frac{R_H}{hc} \left(\frac{1}{n_{lower}^2} - \frac{1}{n_{upper}^2} \right) \quad (3)$$

Where:

λ : wavelength in m

R_H : Rydberg constant for hydrogen ($R_H = 2.180 \times 10^{-18}$ J)

n_{lower} = energy level the electron goes to as it loses energy. For H-atom **visible** light, $n_{lower} = 2$. If $n_{lower} = 1$, the transition lies in the ultraviolet range.

n_{upper} = energy level the electron comes from

h and c represent Planck's constant and the speed of light, as given above

In today's lab, you will observe the visible lines in the emission spectra of both hydrogen and helium. You can determine the wavelengths of these lines by measuring the apparent positions of

the lines of the emission spectrum (images of the lamp) viewed through a diffraction grating. Figure 4 shows the equipment setup schematically, as viewed from above.

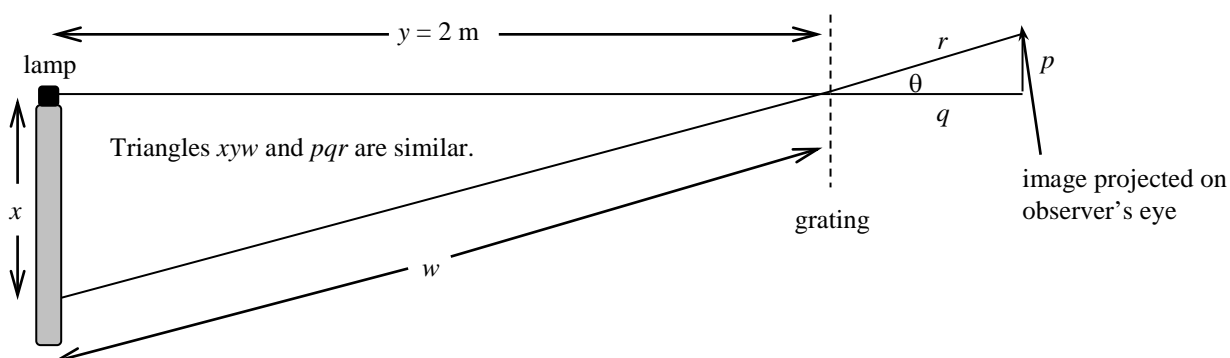


Figure 4: Experimental apparatus

In this configuration, the angle of diffraction, θ (*theta*), depends on the wavelength of the light and the grating space of the diffraction grating, d , according to the approximate relationship:

$$\lambda \approx d \sin \theta \quad (4)$$

Because $\sin \theta = \frac{\text{opposite}}{\text{hypotenuse}}$, then $\sin \theta = \frac{x}{w}$ for the larger similar triangle, and therefore

$$\lambda \approx d \frac{x}{w} \quad (5)$$

You will set up the apparatus so that $y = 2.000$ m, and for each line in each spectrum you will measure the length x with a meter stick. Calculate the length w by the Pythagorean theorem for right triangles, $w = \sqrt{x^2 + y^2}$. Substitute this expression into Equation 5 to give the useful relation;

$$\lambda \approx d \frac{x}{\sqrt{x^2 + y^2}} \quad (6)$$

Our gratings are specified at 750 lines per millimeter, so the grating space $d = 1.33 \times 10^{-6}$ m. Your measured lengths x and y will allow you to calculate the approximate wavelengths of the several visible lines that appear in the spectrum of each element.

PROCEDURE:

Caution: Do not touch lamps, wires, transformers or connections once the equipment is turned on. The lamp will be very hot, and the power supplies have high voltages. Do not move equipment. Do not touch the transparent parts of the diffraction grating—handle the grating only by its mount.

1. Your instructor will divide the class into teams as the equipment availability allows.
2. Mount the diffraction grating 2.000 meters from the light source. Align the zero of the side-mounted meter stick with the center of the light source and perpendicular to the line from the source to the grating.
3. Switch on the light source power supply, and switch off the room lights.

4. One student should sight through the diffraction grating and locate the images of the colored lines out to the side of the gas discharge tube. You may see lines on both sides of the tube; choose the most convenient side for your measurements. If the lines appear as horizontal lines above the gas tube, rotate the diffraction grating 90° before proceeding.
5. Another student should slide a card or ruler along the side-mounted 2-meter stick to measure the apparent location of each line to the nearest millimeter (0.1 cm). Record the distance a (between the center of the lamp and the image of the line as it appears to the viewer) on the data sheet. Proceed in this fashion until you have measured three visible lines for hydrogen and seven visible lines for helium.

From measurements of x and y and the given value of d you can calculate approximate values for wavelength (in meter and nanometer units), frequency (in hertz or s^{-1} units), energy differences (in joule units) between energy levels. In addition, you will determine the upper energy level number (principal quantum number n_{upper} for the higher level) for the transitions in the hydrogen atom. Use the Bohr and Rydberg equations to calculate the quantities required. Note that you can calculate quantum numbers only for hydrogen lines, since the Bohr and Rydberg equations given apply only to hydrogen. Enter the results of your calculations in the data table.

Name _____ Date _____ Score _____

Pre-laboratory Assignment

1. Arrange X-rays, microwaves, infrared, ultraviolet, gamma rays and visible light in order of:
 (a) decreasing wavelength, (b) decreasing frequency and (c) decreasing energy.

| Decreasing Frequency | Decreasing Wavelength | Decreasing Energy |
|----------------------|-----------------------|-------------------|
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FOR FULL CREDIT, SHOW DETAILED CALCULATION SETUPS. REMEMBER TO FOLLOW THE SIGNIFICANT FIGURES CONVENTION, AND TO SHOW MEASUREMENT UNITS FOR EACH QUANTITY.

2. Calculate the wavelength (in m and in cm) of light with frequency of $4.85 \times 10^{15} \text{ sec}^{-1}$

_____ m

_____ cm

3. Calculate the wavelength of the light in question 2 in units of nm.

4. Calculate the energy associated with the light in question 2 in units of joule/atom.

Name _____ Date _____ Score _____

Data Sheet and Calculation Results

| Element | Line | x (m) | λ (m) | λ (nm) | Frequency (s^{-1}) | Energy (J/atom) | n_{lower} | n_{upper} |
|---------|------------|------------|------------------|-------------------|---------------------------|--------------------|-------------|-------------|
| H | red | | | | | | 2 | |
| H | blue-green | | | | | | 2 | |
| H | indigo | | | | | | 2 | |
| He | | | | | | | | |
| He | | | | | | | | |
| He | | | | | | | | |
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Calculations (continue on the back of this sheet):

FOR FULL CREDIT, SHOW DETAILED CALCULATION SETUPS. REMEMBER TO FOLLOW THE SIGNIFICANT FIGURES CONVENTION, AND TO SHOW MEASUREMENT UNITS FOR EACH QUANTITY.

