#  **E8-ATOMIC EMISSION SPECTRA**

Knowledge of the arrangement of electrons around the nuclei of atoms has been obtained by examining the light emitted by **excited** atoms. Atoms become excited when they **absorb** energy; they then **emit** energy in the form of light as they return to a less excited state. Under some conditions, the light emitted may be **visible.**

 When emitted light is reflected from a diffraction grating, only a few bright lines can be observed. These lines make up the atomic emission spectrum of the atom being studied. Each individual line represents a definite wavelength of light emitted by the excited atom.

 The diffraction grating is a device which separates light into a spectrum of its individual wavelengths. A prism behaves similarly. Gratings are made by etching fine, parallel, equally-spaced grooves on a glass plate. The diffraction gratings we will use are important parts of the spectroscopes, to be described below. Light is electromagnetic radiation and can be described in terms of waves with characteristic wavelengths. The relation between the wavelength (**λ**), speed (**c**), and frequency (**υ**) is given by the equation:

**λ ⋅ υ = c (1)**

where **c** is the speed of light in a vacuum, 2.998 x 108 m/s. The wavelength range for **visible** light is about 400 nm (violet) to 780 nm (red) (remember that 1 nm = 10–9 m), corresponding to frequencies of about 1015 s–1.

 In this experiment, we will examine the light emitted when hydrogen or helium absorbs energy from an outside source and then emits energy in the form of light. The two elements are contained in glass tubes and a high-voltage power supply creates an electric discharge through the gas in the tube. In the case of hydrogen, the electric discharge dissociates the H2 molecules into H atoms and excites the hydrogen electron.

 According to Bohr's theory of the hydrogen atom and also in quantum mechanical theory, there exists around the nucleus a series of energy levels in which the electron(s) may reside. In the ground (unexcited) state of an atom, the electrons occupy the lowest available set of energy levels. Some electrons may be excited to higher levels by the absorption of specific amounts of energy. As the electrons return to their ground states, the absorbed energy is released. If an atom has only 1 electron, like hydrogen, then only a few energies are emitted, each corresponding to a specific wavelength (and color in the visible spectrum). Overall, the excited gas has an apparent color that is a mixture of its emitted colors. If the sample contains many excited atoms, all possible energies of light for that atom will be emitted. And if an atom has many excitable electrons, each with a variety of energies it can emit, the overall emission may be virtually a complete visible spectrum, or “white light”.

We will focus here on simple atoms with defined emission spectra. The possible emitted energies can be calculated from the difference of two electronic energy levels. The relationship specifically for hydrogen between the energy, frequency and wavelength of the emitted light is given by the following equation:

 **Efinal – Einitial = ΔE = hυ = hc/λ** (2)

The quantity **h** is Planck's constant and is equal to 6.626 x 10–34 J ⋅ s (J = joule, s = second). The emission spectrum consists of discrete lines corresponding to the differences in energy levels characteristic of, and unique to, the atoms of the element. The lines are normally tabulated by wavelength. The pattern of lines in an atom’s emission spectrum serves as a “fingerprint” for the element and allow identification of that element by spectroscopy. This is the technique used to determine the composition of our Sun and distant stars and for other analyses.

 Bohr found that, for a one-electron atom or ion, the energy of an electron occupying an orbit with principal quantum number **n** is proportional to 1/n2. For hydrogen, this relation is expressed by the equation:

 **En = -k(1/n2)** (3)

where k for the hydrogen atom is the proportionality constant known as the Rydberg constant and has the value 2.179 x 10–18 J; **n** is the principal quantum number and can be any integer from 1 to infinity. When an electron drops from ninitial to nfinal, the energy change is given by the expression:

**ΔE = Efinal – Einitial = [ -k(1/n2)final ] – [ -k(1/n2)initial ] = k [ (1/n2)initial – (1/n2)final ]**  (4)

 From equation (2) we can calculate the wavelength or frequency corresponding to any energy difference. The values for Planck's constant and the Rydberg constant given above lead to a wavelength in meters. For convenience, we usually express the wavelength in nanometers.

 Figure 7 shows some of the electronic transitions possible for a hydrogen atom, and the type of light associated with the transition. Since **n** can take any value from 1 to infinity, there is an infinite number of transitions possible. Note that the energy levels are closer together as **n** increases, so there is a practical limit to the number of distinct lines that can be resolved. And, of course, we can only directly observe the spectral lines in the visible portion of any spectrum. Emissions in other regions of the electromagnetic spectrum require instrumentation sensitive to that band for observation.



*Figure 7: Approximate Electron Transition Energy‑level Diagram for the Hydrogen Atom*

Some thought questions, based on the diagram in Figure 7:

Transition of the excited hydrogen electron **from** which initial **to** which final Principal Quantum Number has the greatest energy of all?

For the Balmer series that we will observe in lab, which transition produces the lowest energy light? Is the color of that spectral line red or violet?

If you could safely touch the hydrogen lamp while it is turned on, you would find it quite warm. Why? (BUT DON”T; the power source is high voltage and could electrocute a careless user! Still, you will observe this same characteristic for all now obsolete conventional light bulbs.)

# **THE SPECTROSCOPE**

 The light source and spectroscope are shown schematically in Figure 8. The leads from the hydrogen or helium tube are connected to the terminals of a power source. There is a potential difference of 5000 volts between the terminals. Keep your hands away from the lamp and power source. **Work only when the instructor is present. This spectroscope is found in the Chemistry 60 labs and when you have time, collect data for comparison to your results.**

Light emitted from the tube enters the spectroscope through a narrow adjustable slit. By using the narrowest slit, a fine beam of light is examined. The beam is focused on a diffraction grating which separates it into its component wavelengths and the resulting visible, colored lines are superimposed on a fixed scale calibrated to read in any convenient units. Our spectroscopes read in cm; that is, 4.25 on the scale means 4.25 x 10–5 cm. The conversion factors 1 cm = 10–2 m and 1 nm = 10–9 m can be used to change our reading to nm: 4.25 x 10–5 cm = 4.25 x 102 nm.



*Figure 8: Measurement of Emission Spectrum Using a Spectroscope*

**PROCEDURE**

You will build your own version of a spectroscope and use direct measurements, plus some physical laws of optics, to determine the emission line wavelengths of helium and hydrogen.

1. Build your spectroscope as directed in figure 9. Using the setup in the laboratory with a diffraction grating and emission tube, and working with a partner, calculate the wavelengths of spectral lines from meter-stick measurements. Determine the wavelengths of the helium lines by carefully observing their positions on the meter stick scale. To avoid parallax error, be sure to look at each line in turn **directly** while reading the scale for that line. Record data for all of the wavelengths that are visible to you in your lab notebook. Some lines are very faint. Some of the wavelengths for helium are given in the table below. Sometimes you may see the color pattern repeat again at the far edge of the diffraction grating. Read only the first, brightest, set of lines.

Some wavelengths for visible lines in the helium spectrum:

| Color | Wavelength in nm |
| --- | --- |
| Red | 667.8 |
| Yellow | 587.6 |
| Green | 492.1 |
| Blue-violet | 447.1 |

2. Use your calibrated “spectroscope” to determine the wavelengths of the three readily visible hydrogen lines. A fourth line may be barely visible, depending on your own eye; record its wavelength if you see it. Compare the wavelengths you observe with the wavelengths calculated for the Balmer Series from equations (4) and (2). See Table 2 in the Pre-Laboratory Assignment.

3. Next, by asking the instructor to change your lamp bulb, determine the wavelengths for the mercury gas lamp. The number of lines to be calculated will depend on the element being studied; consult the instructor. The setup is diagrammed in Figure 9.



*Figure 9: Line Observed with Diffraction Grating*

***DO NOT LOOK DIRECTLY AT THE LIGHT SOURCE****.*

Use a meter stick to measure distances “a” and “x” in Figure 9. “x” is the distance from the emission tube (the source) to a line in the spectrum. One partner, the observer, looks through the grating in order to see a line, and the other partner measures the distance to the point on the meter stick where his partner sees the line superimposed. The value of L can be calculated from:

 L2 = a2 + x2 (the Pythagorean theorem) (5)

The diffracted angle θ is related to the wavelength (λ) of the line by:

 x/L = sin θ = λ/d (6)

where d is the spacing between 2 adjacent lines on the diffraction grating. The spacings are usually given in lines per inch. The grating in our laboratory has 13400 lines/inch. Calculate sine θ from x and L; calculate d in cm/line. Then:

 λ = d sin θ (7)

(What is the unit for sin θ? If d is converted to cm/line, what will be the initial unit for your calculated λ? What is the usual unit for reporting wavelength?)

# **E8-ATOMIC EMISSION SPECTRA**

Section\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ Name\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

 **Report Sheet**

1. Helium Spectrum (you may not actually see nine lines)

**Wavelength (nm)**

 **Line Color Experimental Corresponding**

 **Value Value from Table**

 1 \_\_\_\_\_\_\_ \_\_\_\_\_\_\_ \_\_\_\_\_\_\_

 2 \_\_\_\_\_\_\_ \_\_\_\_\_\_\_ \_\_\_\_\_\_\_

 3 \_\_\_\_\_\_\_ \_\_\_\_\_\_\_ \_\_\_\_\_\_\_

 4 \_\_\_\_\_\_\_ \_\_\_\_\_\_\_ \_\_\_\_\_\_\_

 5 \_\_\_\_\_\_\_ \_\_\_\_\_\_\_ \_\_\_\_\_\_\_

 6 \_\_\_\_\_\_\_ \_\_\_\_\_\_\_ \_\_\_\_\_\_\_

 7 \_\_\_\_\_\_\_ \_\_\_\_\_\_\_ \_\_\_\_\_\_\_

 8 \_\_\_\_\_\_\_ \_\_\_\_\_\_\_ \_\_\_\_\_\_\_

 9 \_\_\_\_\_\_\_ \_\_\_\_\_\_\_ \_\_\_\_\_\_\_

2. Hydrogen Spectrum

 **Wavelength (nm)**

**Line Color Experimental Value from Pre-Lab**

 **Value**

 1 \_\_\_\_\_\_\_ \_\_\_\_\_\_\_ \_\_\_\_\_\_\_

 2 \_\_\_\_\_\_\_ \_\_\_\_\_\_\_ \_\_\_\_\_\_\_

 3 \_\_\_\_\_\_\_ \_\_\_\_\_\_\_ \_\_\_\_\_\_\_

 4 \_\_\_\_\_\_\_ \_\_\_\_\_\_\_ \_\_\_\_\_\_\_

3. Diffraction Grating: Line measurement and calculation

 **Line**

 **Element Number Color Wavelength**

Look up the visible line spectra for mercury, used in trial 3, and compare to your experimental values. Attach a separate sheet with your calculations. Compare the data collected using the Chemistry 60 Spectroscope to your “personally built” spectroscope for H2 and He. Comment on your construction.

# **E8-ATOMIC EMISSION SPECTRA**

 **Post- Laboratory Questions**

1. If the lowest energy emission in the Lyman Series could be observed, would it lie in the visible range? Calculate its wavelength.

2. The Paschen Series of lines arises from electronic transitions to energy level n=3. Calculate the frequency in Hz of the light emitted by an electron falling from n=4 to n=3.

3. Explain why the emissions of the Paschen Series are lower energy overall than those in the Balmer Series.

# **E8-ATOMIC EMISSION SPECTRA**

Section\_\_\_\_\_\_\_\_\_\_\_\_\_\_ Name\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

 **Pre-Laboratory Assignment**

1. Calculate the wavelengths of the lines in the Balmer Series from the values of ninitial and nfinal using equations (4) and (2). Give the wavelength in nanometers. Show a sample calculation here, and enter the values in the table below.

 **Electronic Transitions in the Spectrum of Hydrogen**

**Line ni nf 1/ni2-1/nf2 Wavelength, nm**

1 \_\_\_ \_\_\_ \_\_\_\_\_\_\_ \_\_\_\_\_\_\_

2 \_\_\_ \_\_\_ \_\_\_\_\_\_\_ \_\_\_\_\_\_\_

3 \_\_\_ \_\_\_ \_\_\_\_\_\_\_ \_\_\_\_\_\_\_

4 \_\_\_ \_\_\_ \_\_\_\_\_\_\_ \_\_\_\_\_\_\_

2a. For hydrogen, ionization would correspond to moving the electron from n=1 to ∞. Using equation (4), calculate the ionization energy of one atom of hydrogen in joules.

b. Calculate the ionization energy of atomic hydrogen in kJ/mol.