

Chemistry 513

Scarsdale High School

Atomic Spectroscopy Lab

Objective:

Observe the atomic emission spectra of a number of elements.

Calculate the energy of the “red” line of hydrogen and compare it to the know value.

Relate the concepts and principles of the wave mechanical model to the observations made in this lab.

Use the following file name format to share your lab report via Google Docs:

File Name: **SHS_HC_LB_sec#_Spectroscopy_Lname**

Lab Report:

- Introduction
- Results, Tables and Observations
- Conclusion
- References (MLA Format)

Due Date: One week after the completion of the lab.

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Atomic Spectroscopy

Introduction

According to the Bohr atomic model, electrons orbit the nucleus within specific energy levels. These levels are defined by unique amounts of energy – that is, the energy of the levels is quantized. Electrons possessing the lowest energy are found in the levels closest to the nucleus. Electrons of higher energy are located in progressively more distant energy levels.

If an electron absorbs sufficient energy to bridge the gap between energy levels, the electron may jump to a higher level. Since this change results in a vacant orbital, the configuration is unstable. The excited electron releases its newly acquired energy and falls back to a lower energy level and ultimately to its ground state. Often, the excited electrons return to their ground state, several distinct energy emissions occur. The energy that electrons absorb is often of a thermal (heat) or electrical nature, but the energy that electrons emit when returning to the ground state is always in the form of electromagnetic radiation (light).

In this experiment we will observe the wavelengths of various spectra and compute the frequencies of the waves using the equation:

$$c = \lambda f$$

where c = speed of light = 2.99×10^8 m/s

λ = wavelength (meters)

f = frequency (cycles per second = $1 / s$ = Hz)

In 1900 Max Planck studied visible emissions from hot glowing solids. He proposed that light was emitted in packets of energy called quanta, and that the energy of each packet was proportional to the frequency of the light wave. According to Einstein and Planck, the energy of the packet could be expressed as the product of the frequency (f) of emitted light and Planck's constant, h . From the frequencies determined for the wavelengths of light you observe, you can then compute the energy by the equation:

$$E = hf$$

Where h = Planck's Constant = 6.63×10^{-34} J·s

E = energy (Joules = J)

f = frequency (Hz)

Long ago, Newton demonstrated that if white light passes through a prism or diffraction grating, its component wavelengths are bent at different angles (diffraction). This process produces a rainbow of distinct colors known as a continuous spectrum. If, however, the light emitted from heated gases or an energized ion is viewed in a similar manner, isolated lines of color are observed

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against a black background. These lines form characteristic patterns that are unique to a each element. The lines represent the characteristic wavelengths of light emitted when excited electrons of the gas fall to the ground state level. If, on the other hand, white light passes through a gas, certain wavelengths of light will be absorbed by the gas molecules while allowing the other wavelengths to pass through. The wavelengths absorbed are those have the exact energy needed to excite electrons in the gas. The resulting spectrum would appear as a continuous spectrum (rainbow) with dark lines. The dark lines represent the wavelengths of light absorbed. Such a spectrum is called an absorption spectrum. The dark lines in an absorption spectrum of an element would thus match up exactly with the bright lines in its emission spectrum. This is because the waves producing these lines are identical in energy.

By analyzing the emission spectrum of hydrogen gas, Bohr was able to calculate the energy content of the major electron levels – shells. Although the electron structure as suggested by his atomic model has been modified according to modern quantum theory, his description and analysis of spectral lines are still valid.

In addition to the fundamental role spectroscopy played in the development of today's atomic model, this technique can also be used in the identification of elements. Since the atoms of each element contain unique arrangements of electrons, they will produce a unique pattern of emission lines that can be used as spectral finger prints in identifying the element.

Task 1:

Using the Spectroscopes (black triangular plastic box), observe the following light sources:

1. For each light source, measure the smallest wavelength of light observable in the spectroscope
 - a. White light
 - i. fluorescent lights or look out the window in the direction of the sunlight
 - b. Incandescent light
 - c. Geisler tubes
 - i. Hydrogen
 - ii. Helium
 - iii. Mercury
2. Calculate the frequency for each wavelength measured above
3. Calculate the energy for each wavelength measured above
4. Record this information in a properly constructed table

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Task 2:

Following the instructions listed in “Experiment 47: Flame Tests”

Test the following materials:

| Compound | Observation |
|-------------------|-------------|
| NaCl | |
| KCl | |
| LiCl | |
| SrCl ₂ | |
| BaCl ₂ | |
| MgCl ₂ | |
| CaCl ₂ | |
| CuCl ₂ | |
| Unknown | |

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Experiment 47



Flame Tests

Problem

Can we identify an unknown mixture by using a flame test?

Introduction

Flame tests provide a way to qualitatively test for the presence of specific metallic ions. The heat of the flame excites the electrons in the metal ion, and this energy is released as the electrons "fall back" to their ground states. The color we see is a combination of the visible wavelengths of light emitted by the ion.

In this lab you will perform flame tests on seven different metal ions. You will use your observations to identify two unknown solutions.

Prelaboratory Assignment

- ✓ Read the entire experiment before you begin.
- ✓ Answer the Prelaboratory Questions.
 1. Why must the nichrome wire be cleaned thoroughly between each flame test?
 2. Why do we see colors in the flame tests, and why are there different colors for different metal ions?

Materials

Safety goggles

Lab apron

Nichrome wire loop

Well plate with solutions

Bunsen burner

Wash bottle with deionized water

6.0 M HCl

Seven watch glasses (or petri dishes)

Solutions of the following salts:

barium ~~nitrate~~ chloride

calcium ~~nitrate~~ chloride

copper(II) ~~nitrate~~ chloride

lithium ~~nitrate~~ chloride

potassium ~~nitrate~~ chloride

sodium chloride

strontium ~~nitrate~~ chloride

Safety



1. Safety goggles and a lab apron must be worn at all times in the laboratory.
2. Many of these salts are toxic. If you come in contact with any solution, wash the contacted area thoroughly.
3. The 6.0 M HCl is corrosive. Handle it with extreme care.

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Procedure

1. Clean the nichrome wire. First, rinse with deionized water. Next, dip the loop into the 6.0 M HCl solution. Place the loop into the flame of the Bunsen burner for about a minute. Pay attention to the color of the clean nichrome wire in the flame.
2. Place a small amount of each solution in separate watch glasses.
3. Perform a flame test on each solution by first heating the loop of the nichrome wire in the Bunsen burner. ~~Hold the watch glass with the solution to be tested next to the intake of the Bunsen burner and~~ place the hot loop into the solution. Make careful observations of the flame of the Bunsen burner and record your observations.
4. Clean the wire as described in step 1, and test each of the remaining six solutions separately.
5. Obtain ~~two~~^{one} unknown solutions from your teacher and perform flame tests on each (cleaning the wire between unknowns). Record all observations.

Cleaning Up



1. Clean up all materials (make sure the nichrome wire is clean).
2. Dispose of all chemicals as instructed by your teacher.
3. Wash your hands thoroughly before leaving the laboratory.

Analysis and Conclusions

Complete the **Analysis and Conclusions** section for this experiment either on your Report Sheet or in your lab report as directed by your teacher.

1. How does the flame test provide support for quantized energy levels? Explain your answer.
2. List the metal ions present in your two unknown solutions and provide reasons.

Something Extra

Does the anion in a salt affect the color observed in a flame test? Design an experiment to answer this question, discuss it with your teacher, and try it.

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Cleanup:

PROPERLY dispose of all excess chemicals by putting them into the white waste container in the fume hood. DO NOT DUMP them down the sink!

Return all materials to the same location from where they were taken.

Clean and rinse all glassware and put it back on the appropriate shelf in the cabinets.

Wipe down your lab station and throw away any stray papers or rubbish.