Flame Testing

There is a short list of elements that, when their electrons are given some extra quanta of energy, the energy can be "seen" by the naked human eye. A common, visually interesting, activity in the chemistry lab is to demonstrate this phenomenon by placing one of these elements in a lab burner flame. A bright color will flash momentarily, signaling the energy exchange happening in the energy levels of the element. This activity, for one fairly obvious reason, is known as a flame test. It can be pulled off in many ways, from drizzling a few crystals of a salt onto a flame to evaporating a drop or two of salt solution over a flame.

Thus, when you heat a drop of sodium chloride solution in a lab burner flame, for example, energy is added to the electrons of the sodium atom. This added energy is emitted when the excited electrons in the sodium atoms give off light and fall back to lower shells. The light emitted possesses wavelengths and colors, depending on the amount of energy the atoms absorbed. You'd expect that each excited sodium atom would emit one type of light. But it doesn't quite work that way because there are billions and billions of atoms being excited and they produce billions of excitations and emissions. And not all of the atoms receive the same amount of energy, thus the excited electrons will travel to different energy levels.

This means that the color produced by the heated NaCl solution is actually a series of discrete emission lines. Each emission describes one quantum path that an electron can take after it absorbs energy from the flame. Each element has a unique set of emission lines because each element's electrons have a unique quantized distribution around the nucleus.

The energy and wavelength of the light is described by the equation $E = hc/\lambda$, where λ is the wavelength, h is Planck's constant (6.63×10^{-34} J sec), and c is the speed of light (3.00×10^{8} m/sec). If you capture the image of light emitted by an element heated by a flame, the color your eyes see can be broken down into a set of emission lines, and the emission lines are mathematically described by the equation shown above.

In this experiment, you will use a spectrometer to measure the light emission produced by the salts of several elements as they are heated in a lab burner flame. You will use this information to draw comparisons between the various substances that you test.

OBJECTIVES

In this experiment, you will

- Measure the emission spectrum of various salt solutions.
- Characterize the emission spectra of produced by various elements.

MATERIALS

Vernier Spectrometer, w/o light source/ cuvette holder attachment
or Vernier SpectroVis
computer
Logger *Pro* 3 software
fiber optic accessory
lab burner
dropper bottles of salt solutions, which may include:
LiCl, NaCl, KCl, $SrCl_2$, $CuCl_2$, $BaCl_2$, $CaCl_2$
platinum or nichrome wire flame loop
distilled water
dropper bottle of 1 M HCl solution

PROCEDURE

1. Use a USB cable to connect a spectrometer to your computer. Connect a fiber optic cable to the spectrometer.

2. Start Logger *Pro* 3 (version 3.6 or newer) on your computer.

3. In your data table, record the substances that you will be testing.

4. Prepare the spectrometer to measure light emission.
   a. Open the Experiment menu and select Change Units ► Spectrometer: 1 ► Intensity.
   b. Open the Experiment menu again and select Set Up Sensors ► Spectrometer:1.
   c. Change the Sample Time to **90** ms. Change the Samples to Average to **1**.
   d. Click the red box in the upper right hand corner to close the dialog box.

5. Ignite the lab burner and adjust the flame so that it is mostly a blue color. The flame need not be vigorous.

6. Clean the flame loop by placing one drop of 1 M HCl on the loop itself. Hold the loop in the burner flame for about ten seconds. Repeat this process twice.

7. Measure the emission spectrum of a salt solution. **Note**: This step is easier to do with two people. One person should hold the fiber optic cable and a second person should work with the flame loop.
   a. Click ▶ Collect.
   b. Hold the fiber optic cable with your fingers very nearly at the tip of the cable. Aim the tip of the cable toward the flame, and no close than 5-6 cm from the flame. If your fingertips are warm, the cable is probably too close to the flame.
   c. Select a dropper bottle of salt solution and place one drop of solution on the flame loop.
   d. Hold the loop in the burner flame.
   e. If the graph does not show a series of peaks, take the loop out of the flame and carefully cool it with a few drops of distilled water. Place another drop of the salt solution on the loop and place the loop back in the flame.
   f. When you achieve a satisfactory graph, click ■ Stop. Write down your observations of the emission spectrum in your data table.

8. To store the data, select Store Latest Run from the Experiment menu.

9. Repeat Steps 6 – 8 with the remaining salt solutions.

10. (Optional) Save and/or print a copy of your test results.

DATA TABLE – Flame Test Observations

Salt	Metal	Color of flame	Wavelengths of emission lines

DATA TABLE – Energy Calculations

Salt	Wavelength of an emission line, m	Frequency (Hz, s^{-1})	Energy (J)

Calculation Guide

a. To convert wavelength, 1 nm = 1×10^{-9} m
b. To calculate the frequency, use the equation: $c = \lambda\nu$. ***c*** is the speed of light (3×10^{8} m/s), **λ** is the wavelength in meters, and **ν** is frequency.
c. To calculate the energy of an emission line, use the equation: $E = h\nu$. The value of ***h***, the Planck constant, is 6.63×10^{-34} J sec.

DATA ANALYSIS

1. Describe, in detail, the emission spectrum of each salt. Emphasize the features of each spectrum that distinguishes it from the other salts that you tested.

2. Identify the wavelengths of every peak in the graph of each salt.

3. Choose at least one emission line from each salt and calculate its energy. If possible, use a wide range of wavelengths. Compare the energies of the different emission lines. Which part of the visible light range possesses the greatest energy?