Experiment 16 - Line Emission Spectra and Flame Tests

When elements are heated, the electrons in that element are excited to higher energy levels. These electrons are unstable in their excited state, and they release their excess energy by falling back down to lower energy levels, giving off light in the process. The wavelengths of light given off correspond to the energy changes that the electrons undergo.

$$\Delta E_{electron} = E_{photon} = h\upsilon = hc/\lambda$$

Each element has a different emission spectrum, because each element's electrons have different energy levels and different spacings of those energy levels. In this experiment the spectra of a number of elements will be generated in two ways: by heating in a burner flame, and by electric discharge in a sealed glass tube. The objects of this lab are:

1. To observe the color of light emitted when a sample of a chemical is heated, and to use this color as a means of identifying a few positive ions;
2. To observe the emission spectra of gaseous elements in discharge tubes. In the emission spectrum, the wavelengths of light are separated from each other using a spectroscope, which works much like a prism works to separate white light into a rainbow of colors.

Part 1 - Flame Tests

In this part of the experiment, different solutions containing metal ions are heated in a burner flame. The flame excites the electrons in these metal ions to higher energy levels. As the excited electrons fall back down to lower energy levels, they give off light of specific wavelengths. A number of different wavelengths (each with a different color) are emitted in each case, but without a spectroscope to separate the wavelengths, the eye merely sees them as a single color. Sometimes this color is distinctive enough that it can be used to identify the element.

Part 2 - Spectroscopic Examination

Most natural light is a mixture of different wavelengths, each a different color. The spectroscope is an instrument that separates any beam of light into its constituent wavelengths, and spreads these separated beams out, so they can be seen. (The action of a spectroscope is similar to the action of a prism.)

If the light that is shone into a spectroscope consists of a very large number of wavelengths, each only very slightly different from the ones next to it, the separated light makes what is called a continuous spectrum. It looks like a rainbow of colors, each one merging into the next. An actual rainbow is a continuous spectrum that you see when the atmosphere acts like a giant spectroscope.

Discrete spectra or line spectra are those that consist of just a very few wavelengths of light, all of which are sufficiently different so that when separated, they fall at widely-spaced intervals, with large dark gaps in between them. This type of spectrum looks like separate lines of colored light that have spaces between them. Such spectra occur when elements in the form of very thin gases are heated. It was the study of these line emission spectra that led to the modern theory of electronic energy levels in the atom. The fact that only certain wavelengths of light were produced was evidence that the energy levels of electrons in atoms were *quantized* (only certain energies were possible).

Safety Precautions:

- Wear your safety goggles.
- Use caution when handling the HCl.

Waste disposal:

- Used solutions of metal ions should go in the INORGANIC WASTE bottles (which have a blue label) in one of the fume hoods after the experiment.

Procedure

Part 1 - Flame Tests

1. Obtain a piece of nickel/chrome wire about 15 cm long. Collect a few milliliters of 6 M HCl in a test tube for use as a cleaning solution.
2. Bend one end of the wire into the smallest possible loop, about 1 mm in diameter. This loop will hold a drop of solution when you are doing the flame tests.
3. Clean the wire by first dipping it into 6 M HCl, then holding it in the hottest part of the flame. Repeat until you no longer see any significant color to the flame that comes off of the wire. *Repeat this cleaning process between each sample tested.*
4. To test a solution, let a drop of the solution fall onto the loop of the wire. (Do not let the dropper itself touch the wire; this could cause contamination of the solution in the dropper bottle.)
5. Hold the drop of solution in the burner flame, recording the color you see. **Note:** Na^+ ion gives an especially strong and persistent color. This ion is present as an impurity in most solutions, since glass bottles contain Na^+ which contaminates the solution. The strong color of Na^+ in the flame can obscure other colors such as the pale lavender of K^+. To avoid this difficulty, try looking at a K^+ flame test through a piece of blue glass, which should remove the sodium flame color. Colors that are not orange should come through. It is also a good idea to test the sodium solutions last.
6. Test the known solutions containing K^+, Ba^{2+}, Ca^{2+}, Cu^{2+}, Sr^{2+}, and Na^+. In each case, record the color of the flame produced. Then do flame tests on various miscellaneous items, as follows:
   1. purified water, from a squeeze bottle
   2. tap water
   3. a drop of your saliva
   4. urine from a dropper bottle provided (optional!!)
   5. a raisin (it will be necessary to heat it for several minutes to drive off moisture before the ion's flame color is visible)
   6. pink liquid soap

   In each case, record the color of the flame and decide which positive ion(s) are present.
7. Test an unknown liquid to find out which positive ion is present. (It will be one of the six ions you already tested.) Report the flame color observed and your conclusion as to the identity of the ion. **IMPORTANT:** each person should test his or her own unknown. You and your lab partner must use different unknowns.

Part 2 - Spectroscopic Examination

8. Look through a spectroscope at daylight or a regular tungsten light bulb and at the overhead fluorescent light bulb. All of these sources give "continuous" rather than "discrete line" spectra; that is they give the whole rainbow of colors, each merging into the next.
9. Without moving any of the spectroscopes, look at the displayed elements in gas discharge tubes, and record the spectrum produced by each. Use crayons to record

the colors seen by making vertical lines on your lab report to represent the spectra obtained.

Questions

(Note: to answer these questions, refer to your textbook and your lecture notes.)

1. Fill in the following table regarding the hydrogen spectrum. Refer to your textbook for the initial and final values of n.

color	known λ (nm)	$n_{initial}$	n_{final}
red	656.4		
turquoise	486.3		
purple	434.2		
purple	410.3		

2. Show the calculation of any two of the above wavelengths using:

$$E = -2.18 \times 10^{-18} J\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right) \qquad \text{and} \qquad E = \frac{hc}{\lambda}$$

(Plug in values of n_i and n_f to calculate E, then solve for λ.)

3. Show the calculation for the conversion of one of the above energies to the corresponding photon energy in units of kJ/mole. (Hint: E in the above equations if the energy per photon.)
4. What causes light to be emitted from an atom?
5. Why are only certain wavelengths emitted from an atom? (In other words, why do the spectra show lines of light instead of a continuous rainbow of light?)
6. Why does hydrogen emit different wavelengths of light than mercury?