Experiment 10 - Evidence for Quantum Mechanics

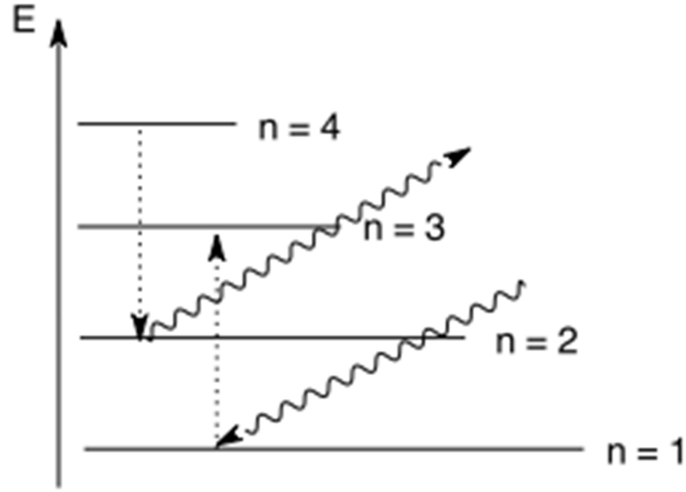

Figure 1: Energy level diagram. Dotted arrows show transitions between energy levels. Wavy arrows show photons being emitted (as in the transition from 4 to 2) or absorbed (as in the transition from 1 to 3).

Introduction

According to quantum mechanics, electrons exist in atoms in distinct, clearly separated **energy levels** or **states**, like stairs, instead of a smooth ramp of energies. The energy of the states is like the height of the step. When atoms absorb energy (by being heated or irradiated with high energy light) the electrons in that atom are excited to higher energy levels (higher steps). These electrons are unstable in their **excited state**, and they release their excess energy by falling back down to lower energy levels (lower steps), giving off light in the process. Electrons can jump or fall many steps at a time, but must start and end *on* a step, not *in between* steps. The wavelengths of light given off correspond to the energy changes that the electrons undergo (the height difference of initial and final steps). The change in energy of the electron is given by

$$\Delta E_{electron} = E_f - E_i \qquad \text{Eq 1}$$

where E_i is the energy of the electron's initial state and E_f is the energy of the electron's final state. Lower energy states are more stable (electrons tend to "fall" to lower energy states, just like objects tend to fall to lower positions); the lowest energy state is called the **ground state**. The zero of energy is usually chosen to be the energy of a "free electron" that isn't bound to any atom. Electrons in atoms or molecules are more stable than free electrons, and thus have negative energies. Thus, **the energy of an energy level is the energy required to remove an electron in that level from the atom**. When electrons change states, they often either emit or absorb a **photon** (light particle) whose energy is the change in energy of the electron (absorbing a photon excites electrons to higher energy states, while relaxing to lower energy states releases a photon). The energy of the photon is the absolute value of the electron's change in energy:

$$|\Delta E_{electron}| = E_{photon} \qquad \text{Eq 2}$$

Recall that there is a simple relationship between the energy of a photon and its frequency and wavelength:

$$E_{photon} = h\upsilon = hc/\lambda \qquad \text{Eq 3}$$

Where h is Planck's constant (6.626×10^{-34} Js), υ is the frequency of the photon, c is the speed of light (3.0×10^8 m/s) and λ is the wavelength of the photon.

Each element has a different emission spectrum, because each element's electrons have different energy levels and different spacings of those energy levels. These energy levels depend

largely on electrostatic interactions between electrons and nuclei. The nuclear charge, number of electrons, and distance between the electrons and nucleus help determine 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 objectives of this lab are generally to explore the experimental evidence for quantum mechanics, and more specifically:
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 various light sources with a spectroscope, which works much like a prism works to separate white light into a rainbow of colors. In particular, we'll look at the emission spectra of gaseous elements in discharge tubes, in which a pure elemental gas is heated very hot by a high voltage, until it emits light.

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

You can do the parts in either order. The hydrogen emission tube can only run for about 30 minutes or it will stop giving good data. Take turns observing it while it is running, and work on other parts of the lab while you wait for a turn.

Part 1 - 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).

1. Look through a spectroscope at daylight or a regular tungsten light bulb. Point the small window directly at the light, then look through the narrow end, and off to the side where the numbers are. Is this a continuous or line spectrum?

2. Look through the spectroscope at the overhead fluorescent light bulb. (This works best if you stand directly under the light and point the spectroscope opening directly at the tube.) Which type of spectrum is this? How is it different from the regular light bulb?
3. If you want, take a spectroscope outside and look at the streetlights with it. What colors do you see?
4. Use the spectroscope to look through the cobalt glass at the regular lightbulb. What changes compared to the spectrum without the glass? What colors are transmitted, and which are absorbed by the glass?
5. Without moving any of the spectroscopes, look at the displayed elements in gas discharge tubes, and record the spectrum produced by each. Sketch the spectrum of each discharge tube, using colored pencils if you wish.
6. Carefully record the wavelengths (in nanometers) of the lines in the hydrogen emission spectrum from the hydrogen discharge tube. Also record the energies in eV (the other scale on the spectrometer.)
7. Carefully record the wavelengths and energies of the lines from the helium discharge tube.

Calculations

1. Calculate the frequency (in Hz) and energy (in J/photon and kJ/mol) of the photons emitted by the hydrogen emission tube using the wavelengths you recorded. You'll also need the energies in eV, so if you didn't read them off the spectrometer you'll need to calculate them.
2. Using the Bohr formula for the energies of stable orbits (energy levels),

$$E_n = -Z^2(13.6 \text{ eV})/n^2$$

where n is the quantum number which identifies the energy level, Z is the nuclear charge (atomic number), and E_n is the energy of the n^{th} level, find the energies of the first six levels (n = 1 - 6).
3. Draw an energy level diagram (similar to Figure 1) for the hydrogen atom showing the relative spacing of the energy levels, and the transitions you observed. Be careful to draw a real energy axis, to scale, and show the actual differences in the energy levels.
4. Your energies for the hydrogen emission lines should be equal to the difference between some of these energy levels. Which transitions do your lines match? Give the initial and final n for each line and show how it fits Eq 2.

Part 2 - Flame Tests

In this part of the experiment, different solutions or powders containing metal ions are heated in a Bunsen burner flame. The flame excites the electrons 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 atom or molecule that produces it.

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. Small short test tubes are best. You may also use several test tubes with HCl for a two-step rinse.
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, some people find it helpful to look at the flame 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; check the colors with and without cobalt glass. 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
7. Test an unknown liquid to find out which positive ion is present. (It will be one of the six ions you already tested in solution.) 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.
8. For more spectacular results you will test some solids as well. To test solids, get a bit of the solid or powder stuck to your loop by dipping in distilled water, then in powder, then hold the loop in the flame. Use very little of the solid or it might be very difficult to clean the loop. As before, clean the loop with HCl in between each measurement, and you may use the cobalt glass to eliminate the sodium flame color if you wish. Solid powders provided might include: Fe^{2+}, and Fe^{3+} salts, $CuCl_2$, $CuSO_4$, boric acid, and cream of tartar.

Name: Section:

Experiment 10 Pre-Lab Sheet

1. (2 pts) What is the difference between a continuous and a discrete spectrum?

2. (2 pts) Why do atomic gases produce line spectra instead of continuous spectra?

3. (2 pts) The energy levels in the hydrogen atom have energies $E_n = -C/n^2$. Why are the energies negative? Hint: how is zero energy defined?

4. (2 pts) What is the frequency of a photon with a wavelength of 315 nm? Show the calculation with all the units.

5. (2 pts) What is the energy in kJ/mol of photons with a wavelength of 315 nm? Show the calculation with all the units.