Bohr Model of the Atom – Line Spectra

PURPOSE

The goal of this experiment is to gain a familiarity with the emission and absorption spectra of elements.

THEORY

When an atom is excited or absorbs energy (in the form of heat, light, or electricity) it often re-emits that energy which includes visible light. Visible light refers to electromagnetic radiation perceptible to the human eye. Light is a form of electromagnetic radiation in which energy is transmitted in the form of a wave. A wave is characterized by both the wavelength(λ) and frequency(υ). If you were to sit on a beach and watch the ocean waves as they came into the beach, the wavelength would refer to the distance between the peaks of each wave and the frequency refers to the number of waves that pass a given point per unit time.

Atoms of each element emit specific wavelengths(λ) and thus a characteristic color. The physicist Neils Bohr developed a model of the atom to explain why excited atoms emit specific wavelengths of light. In his model electrons travel around the nucleus in well-defined fixed orbits similar to the planets around the sun. See figure 1. Bohr's orbits existed at specific distances from the nucleus and had a fixed amount of energy.

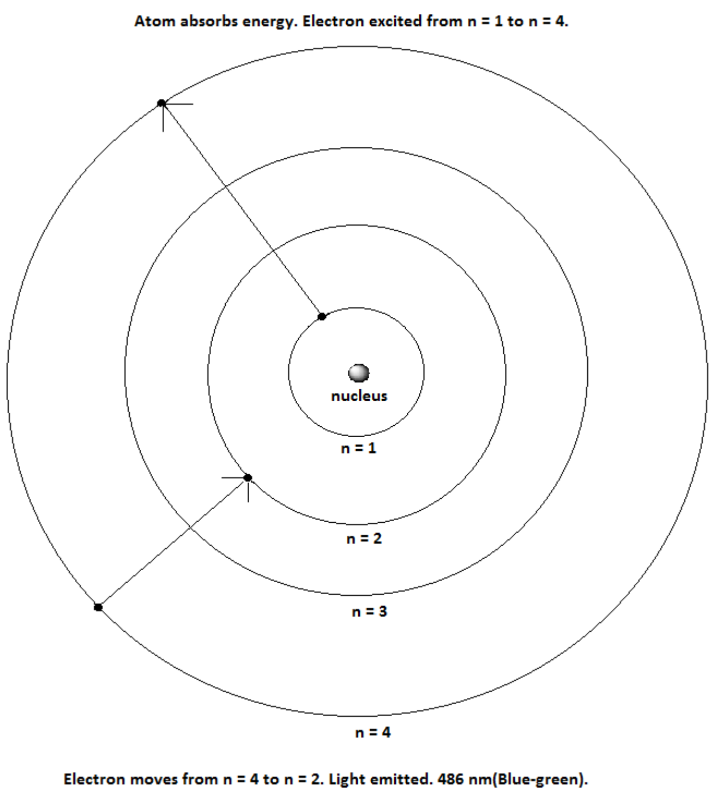

Fig. 1. Bohr model of the Hydrogen atom.

Electrons in orbits closest to the nucleus are the lowest in energy. For an electron to be excited from a low orbit to a high orbit it must absorb a specific amount of energy. This specific amount of energy is the difference in energy between the two orbits(ΔE). When the electron goes from a high orbit to a lower orbit a specific amount of energy(ΔE) is given off.

In Bohr's model each orbit is assigned a n value which has integer values(1 ,2 ,3, etc) and represents the distance from the nucleus. The orbit closest to the nucleus is assigned a n value of one(1). The amount of energy absorbed($+\Delta E$) or emitted($-\Delta E$) can be determined from the following empirical equation

$$\Delta E = R_H\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) \qquad (1)$$

where R_H is known as the Rydberg Constant(2.179×10^{-18}J).

As electrons move between the energy levels of an atom they emit or absorb energy sometimes in the form of light. In this experiment, we will only focus on the transitions involving light. If the electron is going to be promoted, i.e. move from a lower energy level to a higher energy level, the atom must absorb energy by absorbing light. If the electron falls from a higher energy lever to a lower energy level it will release energy by emitting light. The energy of the light emitted or absorbed must exactly match the difference in energy between the starting atomic energy level and the final atomic energy level. The wavelength (color) of light is related to the energy of light by the equation:

$$\lambda = \frac{hc}{\Delta E} \qquad (2)$$

where ΔE is the absolute value of the change in energy in joules, h is Planck's constant (6.626×10^{-34} J·s), c is the speed of light(3.00×10^{8} m/s), and λ is the wavelength in centimeters. Wavelengths of light are ordinarily given in nm; where 1 nm equals 1×10^{-9} m. The change in energy, ΔE, of the atom or molecule is positive if light is absorbed and negative if it is emitted.

Because of the need for the energy of light to exactly match the energy difference between the initial and final energy level, a given type of atom will not absorb all kinds of light. It will only absorb those colors of light that have the correct amount of energy. Since every kind of atom has a slightly different electronic configuration, the wavelengths of light absorbed or emitted by an element are unique to that type of element, like an atomic fingerprint. The fingerprint recorded by looking at which wavelengths of light an atom emits is called the Emission Spectrum.

In this experiment we will be looking at the Emission Spectra of several different elements. These elements are present in sealed glass tubes which contain a pure gas of the given element, like H_2, He, Na, etc. The tubes are placed in a special holder that passes an electrical current from one end of the tube to the other, like in a neon light. The electricity passing through the gas excited the electrons in the gas molecules and promotes them to higher energy levels. When they return to their ground state configurations they emit a characteristic wavelengths of light, i.e. an emission spectrum. Within every atom there are many different electronic transitions possible. Therefore there are many different wavelengths of light emitted. In order to separate the different wavelengths of light we will use a hand held spectroscope. These spectroscopes are similar to prisms in that they separate a mixture of light into its component colors. The list of different wavelengths of light emitted by an element is the Emission Spectrum of that element.

In this experiment you will measure the wavelengths of three lines in the hydrogen atomic spectrum, and be asked to explain the origin of each line in terms of the energy levels of the atom.

There are several ways in which one might analyze an atomic spectrum, given the energy levels of the atom, but a simple and powerful one is to calculate the wavelengths of some of the lines that are theoretically allowed and to see if they match those which are observed. We shall use this method in our experiment.

Procedure

SAFETY AND WASTE DISPOSAL

Wear approved eye protection at all times in the laboratory. Some reagents may be dissolved in strong acids(nitric acid) and while the amounts are small caution is needed when using in open flame.

Part A:

You will be given a spectroscope with which to view a hydrogen lamp and another lamp(mercury, nitrogen and/or a helium). You will be able to see the prominent lines in each.

WARNING: DO NOT LOOK AT THESE LAMPS WITH THE NAKED EYE!!

Record the four prominent lines of mercury, the three prominent lines of hydrogen, the three prominent lines of nitrogen and the four prominent lines of helium. Fill into your data sheet.

Part B:

When metal ions in solutions are evaporated in a flame, each kind of metal atom is excited and gives off light at a few discrete wavelengths. The color of the light depends on the type transitions and the type of ion.

Color Absorbed	Color Observed	Wavelenght Range Absorbed(nm)
Violet	Yellow-Green	400-435
Blue	Yellow	435-480
Green-Blue	Orange	480-490
Blue-Green	Red	490-500
Green	Purple	500-560
Yellow-Green	Violet	560-580
Yellow	Blue	580-595
Orange	Green-Blue	595-605
Red	Blue-Green	605-750

Observe the emission spectra of the cations in various solutions, as they are aspirated into the flame of a Bunsen burner. Record the color of the aspirated flame and the corresponding wavelength range of the emission spectrum into your data sheet using the provided table.

Part C:

Observe the spectrum of the light from a fluorescent lamp with your spectroscope, and note your observations. Compare to the spectrum for sunlight if possible.

CALCULATIONS:

Part D: Calculation of the Wavelength of the Lines in the Hydrogen Spectrum NOTE: KNOW FOR QUIZ.

The lines in the hydrogen spectrum all arise from jumps made by the electrons from one energy level to another. The wavelengths in nm of these lines can be calculated by Equation 1 and 2. For example, to find the wavelength of the spectral line associated with the transition from the n = 2 level to the n = 1 level, calculate the difference, ΔEN, between the energies of those two levels. Using equation 2, the wavelength in nm of the spectral line will be determined. To convert the wavelength to nm, use the conversion factor 1 nm = 1×10^{-9} m.

Part E: Assignment of Observed Lines in the Hydrogen Spectrum

From Part A take the wavelength of the hydrogen lines. Write these in Table 5 of your data sheet. **Note**: the second column will be filled in after appropriate calculations have been performed in Table 4.

Compare the wavelengths you have calculated (Table 4) with those which you measured (Table 5). If you have been careful, your measured wavelengths should match several of those which you calculated. On the line opposite each wavelength in Table 5, write the quantum numbers of the upper and lower states for each line whose origin you can recognize by comparison of your calculated values with the observed values.

REVIEW QUESTIONS – KNOW FOR QUIZ:

1. Briefly explain the Bohr Model of the Hydrogen Atom.

2. What is meant by an Atomic Spectra or Line Spectra?

3. Know how to calculate the energy(J), the wavelength(nm), and the energy in kJ/mole when an electron moved from one orbit to another. Understand the importance of a calculated value that is positive or negative.

4. The hydrogen lines you have observed belong to the Balmer series. When Balmer found his famous series for hydrogen in 1886, he was limited experimentally to wavelengths in the visible and near ultra-violet regions from 250.0 nm to 700.0 nm. From the entries in Table 5, what common characteristic do the lines in the Balmer series have?

5. In a normal hydrogen atom, the electron is in the lowest energy state. How much energy in does it take to ionize a H atom? Hint: Electron is excited to an orbit an infinite distance away from the nucleus.

6. Describe the spectrum of the fluorescent light and explain what you observe.