

# PHY221 LAB 07-1: Absorption and Emission of Photons

## Bohr model of hydrogen

The Bohr model is a semi-classical model of the hydrogen atom. It is wrong! However, it had success in predicting the ground state (i.e. lowest) energy of atomic hydrogen, and it's an easy model to understand.

In the Bohr model, just as planets attracted to a star orbit the star, an electron attracted to a proton orbits the proton. However, unlike planet-star systems, the electron-proton system is only allowed to have certain orbits, i.e. certain radii and certain energies. These allowed orbits are labeled  $n = 1$ ,  $n = 2$ , etc., with  $n = 1$  being the lowest radius (and lowest energy) orbit.

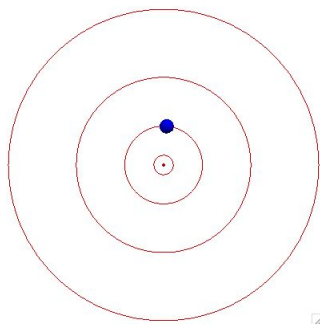


Figure 1: A planetary model for a hydrogen atom.

What is  $n$  for the electron shown in the image above? Note that the electron is NOT drawn to scale.

A better model of a hydrogen atom is the “electron cloud” model. An electron can be found, with a certain probability, anywhere inside a certain cloud. (The cloud for  $n=1$  is shown in Figure 2.) The shape, size, and configuration of this cloud depends on the energy of the electron. Though this model is more accurate, our simple minds prefer simple models based on things we already understand. Thus, we'll often appeal to the Bohr model as we think about energy states of a hydrogen atom.



Figure 2: The “electron cloud model” for a hydrogen atom.

In the Bohr model, an electron orbits the proton in a circular orbit at a certain distance away. Only certain orbits are allowed. Those orbits are the ones where the sum of the kinetic energy of the electron, the kinetic energy of the proton, and the coulomb potential energy of the electron-proton interaction is

$$E = K + PE = \frac{-13.6 \text{ eV}}{n^2} \quad n = 1, 2, 3, \dots \quad (1)$$

$n$  is an integer that represents the energy state (i.e. the particular orbit in the Bohr model) of the atom. The lowest energy of the atom is called the *ground state*,  $n = 1$ .

When an electron moves from a lower energy to a higher energy, the electron-proton system (i.e. atom) gains energy. When an electron moves from a higher energy to a lower energy, the electron-proton system loses energy.

Make a table and calculate the energies for the first 7 orbits.

What is the change in the energy of a hydrogen atom,  $\Delta E$ , if the electron makes the following transitions?

$n = 1$  to  $n = 2$  :

$n = 1$  to  $n = 3$  :

$n = 2$  to  $n = 3$  :

$n = 2$  to  $n = 4$  :

$n = 2$  to  $n = 5$  :

$n = 2$  to  $n = 6$  :

$n = 3$  to  $n = 4$  :

$n = 3$  to  $n = 5$  :

$n = 3$  to  $n = 6$  :

If the atom transitions from a higher state to a lower state in each of the cases above (such as from  $n = 2$  to  $n = 1$ ), how would it change your answers?

## Photon Energy

Light can be modeled as a particle called the *photon*. A photon of light has no mass, thus it's unlike many other particles that we are familiar with. However, it still has energy and momentum. Because its mass is zero, a photon has zero rest energy; therefore, all of its energy is kinetic energy. Its energy is

$$E = hf \quad (2)$$

where  $f$  is the frequency of the photon and  $h$  is Planck's constant,  $6.63 \times 10^{-34} \text{ m}^2\text{kg/s}$  or  $4.136 \times 10^{-15} \text{ eV s}$ . In the wave model, light is an electromagnetic wave whose electric field and magnetic field oscillate with a frequency  $f$ . The wavelength of the wave and its frequency are related by

$$\lambda f = c \quad (3)$$

where  $c$  is the speed of light in a vacuum,  $3 \times 10^8 \text{ m/s}$ . The wavelength of light has units of distance. Visible light wavelengths are generally given in units of nm. The prefix "nano" means "one billionth". Thus, there are  $1 \times 10^9 \text{ nm}$  in 1 m.

## Photon Emission and Absorption by Hydrogen

If a continuous spectrum of light (meaning that the light contains photons of all energies) shines on atomic hydrogen gas (the atoms are not bound as molecules), then the hydrogen atoms will absorb certain photons and will increase in energy to a new energy level. The photons absorbed are the ones that correspond exactly to a change in energy between allowed energy levels.

$$|\Delta E_{\text{H atom}}| = E_{\text{photon}}$$

If the hydrogen absorbs a photon, then  $\Delta E_{\text{H atom}}$  is positive. If it emits a photon, then  $\Delta E_{\text{H atom}}$  is negative.

## Experiment

In this experiment, you will study the visible photons emitted by hydrogen by viewing the emission spectrum of hydrogen, and you will compare the energies of the visible photons to the change in energy of the hydrogen atom.

1. Calibrate the hand-held spectrometer. Point the slit of the spectrometer at the fluorescent lights. Then, look through the diffraction grating at the scale. You should see a series of lines on the scale, superimposed on top of a continuous spectrum.
2. The green line has a wavelength of 547 nm, and therefore should be at 547 nm on the scale. If it is not, then stick the tip of a pencil in the small hole on the scale and move it side to side until the green line is at 547 nm.
3. The scale is now calibrated.
4. Point the slit of the spectrometer at a hydrogen source and view the spectrum.
5. Record the wavelengths, in nm, of the spectral lines that you see.

Make a data table indicating the color and wavelength of the visible lines.

6. Each color line represents an emitted photon of a certain frequency. Thus, if you see four lines, then there are only four different frequency photons emitted in the visible region by hydrogen. (I say emitted in the visible region because our eyes only detect photons with frequencies in the visible region. Hydrogen emits photons with frequencies in the infrared and ultraviolet regions, but our eyes cannot detect them.)

It's wonderful to see the spectral lines; however, scientists typically use more precise instruments to measure the intensity of light as a function of wavelength. Then, they graph intensity as a function of wavelength (or sometimes frequency or sometimes wavenumber). The peaks on the graph correspond to the emission spectral lines that you would see.

1. An Ocean Optics spectrometer has been set up for you at a workstation. Go to that workstation to complete your experiment.
2. View Logger Pro. It should already be set up and configured.

According to the programmers of Logger Pro who placed the continuous visible spectrum on the graph, what is the wavelength region for the visible spectrum? (i.e. What is the minimum and maximum wavelength for visible light?)

Does light of wavelength 300 nm not exist or can you simply not see it with your eyes?

According to the programmers of Logger Pro and yourself who is viewing the colors, what are the regions of wavelength for the colors: Violet, Blue, Green, Yellow, Orange, Red? Make a table below showing the wavelength regions for these colors? Also, record the frequency range and energy range for each color, using Equations 1 and 2 above to calculate frequency and energy for a given wavelength.

3. By default, Logger Pro plots absorbance vs. wavelength. But we wish to measure intensity. Logger Pro may already be set up to plot intensity vs. wavelength. If it is not, then go to Experiment→Change Units→Spectrometer: 1→Intensity to change what is being measured by the spectrometer.

4. Point the fiber optic cable at the light source, holding it a small distance (a few cm) away from the source.
5. Click .
6. After getting a reasonable spectrum, click .
7. Measure the wavelengths of the peaks found in the spectrum. To do this, click the  icon in the toolbar. This will give you an inspector that you can use to read the data points on the graph.
8. Make a data table listing the wavelengths of the lines measured with your cheap hand-held spectrometer and the wavelengths of the peaks measured with the spectrometer. Do the peaks correspond to the spectral lines that you saw in the hand-held spectrometer? What wavelength is each color line for hydrogen?

## Analysis

Using your measured wavelengths, calculate the frequency and energy (in eV) for the visible photons emitted by hydrogen. Make a table showing the measured wavelength, measured frequency, and measured energy of the visible photons emitted by hydrogen.

Compare the energies for the peaks to the allowed energies (i.e. energy states) for the hydrogen atom. Do the energies of the photons emitted by hydrogen (i.e. energies of the peaks) equal any of the allowed energies of the hydrogen atom?

By examining the change in energy of a hydrogen atom when it transitions from one state to another state, compare the change in the energy of a hydrogen atom with the energy of the photons emitted by hydrogen. Which transitions correspond to the emission of visible photons?

Do the other transitions not emit photons or can we just not see them?

Based on your findings, write an equation that relates the energy of a photon emitted by hydrogen to the energy of a hydrogen atom.

## Electromagnetic Spectrum

Photons can have various frequencies. As a result, scientists have grouped and named certain ranges of photon frequencies (or wavelengths). This way of categorizing photons is called the *electromagnetic spectrum*. The electromagnetic spectrum can be defined by wavelength, frequency, or energy, since all of these are related. For the purpose of relating atomic transitions to the energy of photons absorbed or emitted, it helps to know the energies of the various regions of the electromagnetic spectrum.

The following definitions of the regions of the electromagnetic spectrum were taken from *HyperPhysics* by Carl R. Nave at Georgia State University.

Region	Wavelength	Frequency	Energy
Gamma Rays	< 0.001 nm	$> 10^{20}$ Hz	$> 1$ MeV
X-rays	10 nm – 0.001 nm	$3 \times 10^{16} - 1 \times 10^{20}$ Hz	124 eV – 1 MeV
Ultraviolet	400 nm – 10 nm	$7.5 \times 10^{14} - 3 \times 10^{16}$	3.1–124 eV
Visible	750 nm – 400 nm	$4 \times 10^{14} - 7.5 \times 10^{14}$ Hz	1.65 – 3.1 eV
Infrared	1 mm – 750 nm	$0.003 \times 10^{14} - 4 \times 10^{14}$	0.0012 – 1.65 eV
Millimeter Waves, Telemetry	10 mm – 1 mm	30 – 300 GHz	$1.2 \times 10^{-4} - 1.2 \times 10^{-3}$ eV
Microwaves, Radar	187 mm – 10 mm	1.6 – 30 GHz	$6.6 \times 10^{-6} - 1.2 \times 10^{-3}$ eV
TV and FM Radio	5.55 m – 0.187 m	54 – 1600 MHz	$2.2 \times 10^{-7} - 6.6 \times 10^{-5}$ eV
Short Wave	187 m – 5.55 m	1.605 – 54 MHz	$6.6 \times 10^{-9} - 2.2 \times 10^{-7}$ eV
AM Radio	600 m – 200 m	500 – 1500 kHz	$2 \times 10^{-9} - 6.6 \times 10^{-9}$ eV

Table 1: Electromagnetic Spectrum

## Visible Spectrum

Defining where one color ends and another begins requires completely artificial definitions. But at least the middle of these ranges are well defined. The following designation is from *University Physics* by Young and Freedman.

Region	Wavelength	Frequency	Energy
Violet	440 nm – 400 nm	$6.8 \times 10^{14} - 7.5 \times 10^{14}$ Hz	2.8 – 3.1 eV
Blue	480 nm – 440 nm	$6.3 \times 10^{14} - 6.8 \times 10^{14}$ Hz	2.6 – 2.8 eV
Green	560 nm – 480 nm	$5.4 \times 10^{14} - 6.3 \times 10^{14}$ Hz	2.2 – 2.6 eV
Yellow	590 nm – 560 nm	$5.1 \times 10^{14} - 5.4 \times 10^{14}$ Hz	2.1 – 2.2 eV
Orange	630 nm – 590 nm	$4.8 \times 10^{14} - 5.1 \times 10^{14}$ Hz	2.0 – 2.1 eV
Red	700 nm – 630 nm	$4.3 \times 10^{14} - 4.8 \times 10^{14}$ Hz	1.8 – 2.0 eV

Table 2: Visible Region of the Electromagnetic Spectrum

## Emission and Absorption of a Photon by a Hydrogen Atom

If light “collides” with a hydrogen atom, the atom will absorb a photon *if and only if* the energy of the photon is equal to the energy of a transition between orbits. In this case, the energy lost by the photon (because it is absorbed and is no more) is equal to the energy gained by the hydrogen atom.

Likewise, if a hydrogen atom is in a higher energy state (i.e. higher orbit), and the electron makes a transition to a lower energy orbit, the atom will emit a photon of energy equal to the energy loss of the atom. In this case, the energy lost by the hydrogen atom is equal to the energy of the photon created and emitted.

Conservation of energy gives

$$|\Delta E|_{atom} = E_{photon} \quad (4)$$

If the photon is absorbed, then the electron makes a transition to a higher energy orbit. If the photon is emitted, then the electron makes a transition to a lower energy orbit. The difference of the energies of the orbits is  $\Delta E_{atom}$ .

Suppose a photon is absorbed by a hydrogen atom causing the electron to be "excited" from the  $n=1$  state to the  $n=5$  state. What "kind" (i.e. x-ray, ultraviolet, visible, infrared, etc.) of photon was this?

What color photon would be emitted if an electron made a transition from  $n=3$  to  $n=2$ ?

A beam of photons, each with energy 12.09 eV, is incident on a container of atomic hydrogen where all of the atoms are in the ground state. To what state will a particular atom be excited?

A beam of photons, each with energy 11 eV, is incident on a container of atomic hydrogen where all of the atoms are in the ground state. To what state will a particular atom be excited?

What type of photons are emitted when hydrogen falls from a state  $n > 1$  to  $n = 1$ ?

What about  $n > 2$  to  $n = 2$ ?

What about  $n > 3$  to  $n = 3$ ?