

# Using a Smartphone Spectroscope to Observe Electronic Transitions of Atoms<sup>1</sup>

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## Learning Outcomes:

1. Observe emission lines from a fluorescent bulb, and calibrate smartphone spectroscope.
2. Observe emission lines of hydrogen.
3. Determine the initial electronic state responsible for the observed emission line.
4. Describe the quantization of energy at the atomic level.
  - a. Describe the relationship between energy and light (wavelength and frequency)
  - b. Explain absorption and emission in terms of transitions between energy levels
  - c. Recognize spectral lines result from transitions of electrons between energy level

## Introduction:

The periodic table organizes elements based on their electronic structure (*i.e.* the arrangement of electrons around the nucleus). How was electronic structure determined? This mystery was unraveled through spectroscopy, a branch of science which investigates how matter interacts with light. This exercise illustrates how electrons in atoms interact with light, and how spectroscopic information is used to determine the organization of electrons in an element.

Our current understanding of electronic structure is the outcome of contributions by many scientists who measured and analyzed spectroscopic information. For example, Max Planck (1858 – 1947) was one of the first scientists to conclude that atoms interact with light through the absorption and emission of energy. He proposed the quantitative relationship between energy ( $E$ ) and frequency of light ( $\nu$ ) as:

$$E = h\nu \quad (1),$$

where  $h$  is a constant named in his honor (Planck's constant =  $6.626 \times 10^{-34}$  J·s).

Neils Bohr (1885 – 1962), a contemporary of Planck, used this relationship between light and energy to explain line emission spectra observed for hydrogen. He postulated that electrons orbit the nucleus like a planet orbiting the sun. The energy of an electron in an orbit about the nucleus has a defined energy. When electrons change orbit, either by absorbing or emitting energy, line spectra are produced. The wavelength of hydrogen line spectra is reproduced by the following equation:

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<sup>1</sup> Meets Florida teaching standard **SC.912.P.10.9**. *Explain that when electrons transition to higher energy levels they absorb energy, and when they transition to lower energy levels they emit energy. Recognize that spectral lines are the result of transitions of electrons between energy levels that correspond to photons of light with an energy and frequency related to the energy spacing between levels (Planck's relationship  $E = h\nu$ ).*

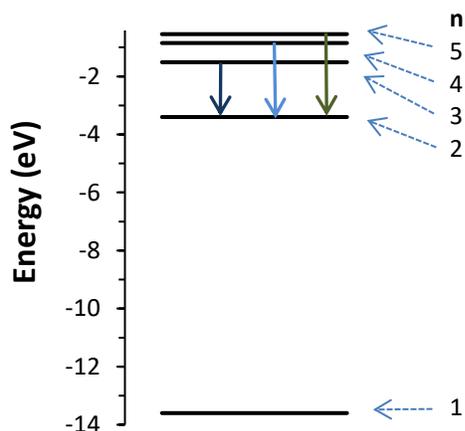
$$\frac{1}{\lambda} = \frac{R_h}{hc} \left( \frac{1}{n_f^2} - \frac{1}{n_i^2} \right) \quad (2),$$

where  $R_h$  the Rydberg constant ( $2.18 \times 10^{-18}$  J),  $c$  is the speed of light, and  $n$  is an integer representing the electronic state (the subscripts represent the initial and final electronic state). Inserting the appropriate values for the constants (and  $n_f = 2$ ), the wavelengths of line emission spectra observed for hydrogen are reproduced by the following equation:

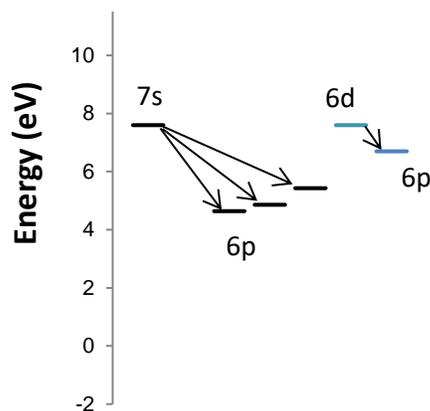
$$\frac{1}{\lambda} = 0.011 \left( 0.25 - \frac{1}{n_i^2} \right) nm^{-1} \quad (3).$$

Energy diagrams, as illustrated in Figure 1, are commonly used by chemists to illustrate different energy states of an electron ( $n$ , solid line), and energy transitions (solid arrows). In this exercise, you will record wavelengths of hydrogen energy transitions (Figure 1A) based on a calibrated spectroscope using wavelengths associated with mercury transitions (Figure 1B). Based on your observations, you will determine what energy transitions (e.g.  $n = 5 \rightarrow 2$ ) are associated with each hydrogen emission line, as well as the value of the Rydberg constant in Eq. 2 and 3.

### A. Hydrogen



### B. Mercury



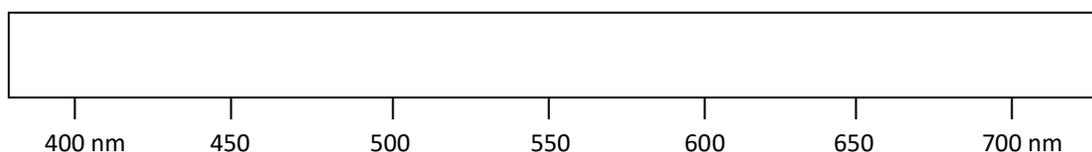
**Figure 1. Energy diagram illustrating electronic transitions of hydrogen (A) and mercury (B).**

Unfortunately, the Bohr model described above only works for atoms containing one electron (i.e. hydrogen). It is much more difficult to explain the line emission spectra of mercury. The field of quantum mechanics developed to describe the energy states of an electron in multi-electron systems. Since electrons interact with each other, the energy levels within an electronic state ( $n$ ) are different, and they are described by additional electronic sublevels (e.g.  $s$ ,  $p$ ,  $d$ , and  $f$ ). For example, mercury energy transitions, such as  $7s \rightarrow 6p$ , are illustrated in Figure 1B.

Name: \_\_\_\_\_

**Procedure:**

1. Follow the instructions for the smartphone spectroscope found at [www.kissinstruments.com](http://www.kissinstruments.com). Download app from *iTunes* or *Googleplay* as appropriate for your smartphone, and watch instructional video for calibrating spectroscope. Calibrate the scope using a fluorescent light source and record the wavelengths for hydrogen line spectra.
2. Identify colors associated with the wavelengths in the following diagram. Indicate the relationship between wavelength and energy by drawing an arrow pointing to higher energy.



3. Fill in the following table. Use the observed wavelength and the following equation:

$$\frac{1}{\lambda} = 0.011 \left( 0.25 - \frac{1}{n_i^2} \right) nm^{-1}$$

to determine the initial electronic state ( $n_i$ ) responsible for the spectral line. Note that the value of  $n$  is an integer, so round off to one significant figure.

Wavelength of Spectral line (nm)	Color	$n_i$
410 nm	Violet	6

- \_\_\_\_\_ 1. Which color of light is the **most** energetic?  
A. red      B. yellow      C. green      D. violet
- \_\_\_\_\_ 2. With respect to the Bohr model of the hydrogen atom, which electronic transition would result in emission of visible radiation?  
A. Level 5 to level 1      C. Level 5 to level 3  
B. Level 5 to level 2      D. Level 5 to level 4
- \_\_\_\_\_ 3. With respect to the Bohr model of the hydrogen atom, what color of light is emitted when an electron moves from the third energy level to the second energy level?  
A. red      B. yellow      C. green      D. violet
- \_\_\_\_\_ 4. Why does the Bohr model fail for atoms other than hydrogen?  
A. proton/proton interactions.  
B. proton/electron interactions.  
C. electron/electron interactions.  
D. none of the above
- \_\_\_\_\_ 5. What wavelength range is associated of visible light  
A. 300-900 nm  
B. 400-700 nm  
C. 200-1000 nm  
D. 800-1000 nm