

Investigating and identifying elements through spectroscopic emissions after exposure to thermal energy

Introduction:

The purposes of this lab were to I) understand the Bohr and quantum mechanical models of the atom and their relation to atomic emission spectrums, II) perform spectroscopic analysis on several samples to develop proficiency with a spectroscope and a spectrometer, III) develop an understanding of the role that emission lines and spectroscopy have played in the refinement of atomic models, and IV) develop an understanding of the role that emission lines and spectroscopy have played in the discovery of new substances. Elements and compounds were exposed to thermal energy in order to observe the Bohr model of the atom in conjunction with atomic photon emissions, both qualitatively (Part I: flame color) and quantitatively (Parts II & III: emission spectrums).

Niels Bohr, a Danish physicist, developed an extension of Thomas Rutherford's model of the atom in 1915. Bohr theorized, like Rutherford, that the atom includes a nucleus, which is denser than the rest of the atom, as well as negatively charged particles (later labeled electrons). However, Bohr further refined the Rutherford model by stating that electrons orbit the nucleus in a circular pattern. Bohr stipulated that electrons absorb energy from light photons, causing them to "jump" away from the nucleus. As they "fall" back into orbit, [Figure 1] they release the absorbed energy in order to stabilize. This released energy takes the form of light photons, cast off from the electrons in specific color signatures. The frequency and wavelength of the light photon determine how much energy is transferred to the electron; frequency of a light wavelength determines the color of the light. The lower the wavelength of the light particle (nanometers), the higher the frequency (hertz); thus, higher frequency/lower wavelength light photons transfer more energy as they come into contact with electrons. Electrons "fall" in a way with no discernable pattern. The color of the light photon emitted from the falling electron is determined by the energy required for the electron to jump back into place. For example, the energy required to jump from energy level 6 to energy level 5 is different than the jump from energy level 2 to energy level 1, because the gap between the latter is more significant, and thus requires more energy to achieve (Figure 2). Bohr's model has since been improved, but the quantum mechanics remain relevant.

The quantum mechanical model of atomic structure, which succeeded Bohr's model, employs the "uncertainty principle," which states that the position and momentum of an electron are impossible to determine. Instead, the quantum mechanical model illustrates electrons as cloud-like formations, basing electron placement within an atom on likely probability of presence (Figure 3).

The atomic emission spectrum of a material is the collection of electromagnetic radiation of electrons as they transition from a high-energy state to a low-energy state, as Bohr theorized. Each element has a unique emission spectrum because each element has varying distances of gaps between energy levels, making it possible for scientists to identify an unknown substance based on photon emission signatures. Spectroscopy is the study of the relation of matter to emitted energy. Spectroscopy originated through the study of visible light as separated through a prism. A spectrometer is the instrument used to separate the atomic photon emissions into the visible light spectrum based on wavelength. As new elements have been discovered, their individual emission spectrums have been recorded, making it easy for scientists to both confirm a discovery as well as identify unknown materials. In 1868, Pierre Jensen, a French astronomer, observed a solar eclipse through a spectroscope. Jensen “found two spectra, composed of five or six very bright lines:” red, yellow, and blue. Comparing his results to known spectral signatures of elements, Jensen matched the red and blue lines to hydrogen. Because Jensen was unable to identify the element associated with the yellow spectral signature, he concluded that he had identified a new element. However, it wasn’t until 1895 that his discovery of a new element was confirmed. Norman Lockyer, a British astronomer, repeated Jensen’s experiment a few months later in 1868 and observed the sun through a spectroscope and arrived at the same conclusion: Hydrogen was present, but there was an unknown element that was responsible for the yellow spectral emissions. Lockyer called the new element helium after *helios*, the Greek word for sun. 30 years later, Lockyer was finally able to confirm the discovery of helium after he collected a material with the same emission spectrum of that of the sun from the radioactive decay of a uranium material.

Figure 1: An illustration of the Bohr model of the atom



<http://library.thinkquest.org/C006669/data/Chem/atomic/bohr.html>

Figure 2: The energy levels of electron orbit in an atom

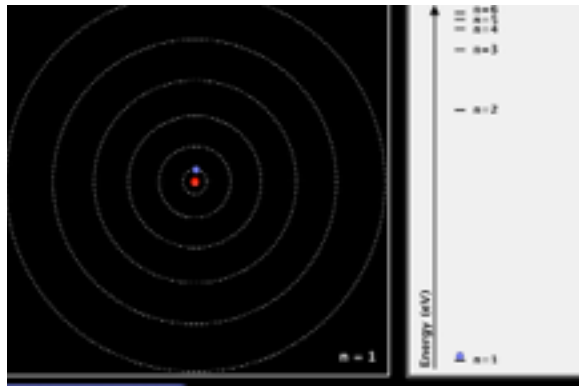
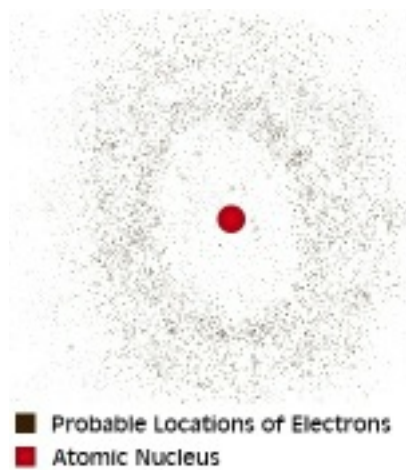


Figure 3: The quantum mechanical atomic model including the illustration of electron clouds



<http://www.buzzle.com/articles/electron-cloud-model.html>

Results: Part I Qualitative

Figure 4: Visual color emissions of materials when exposed to thermal energy

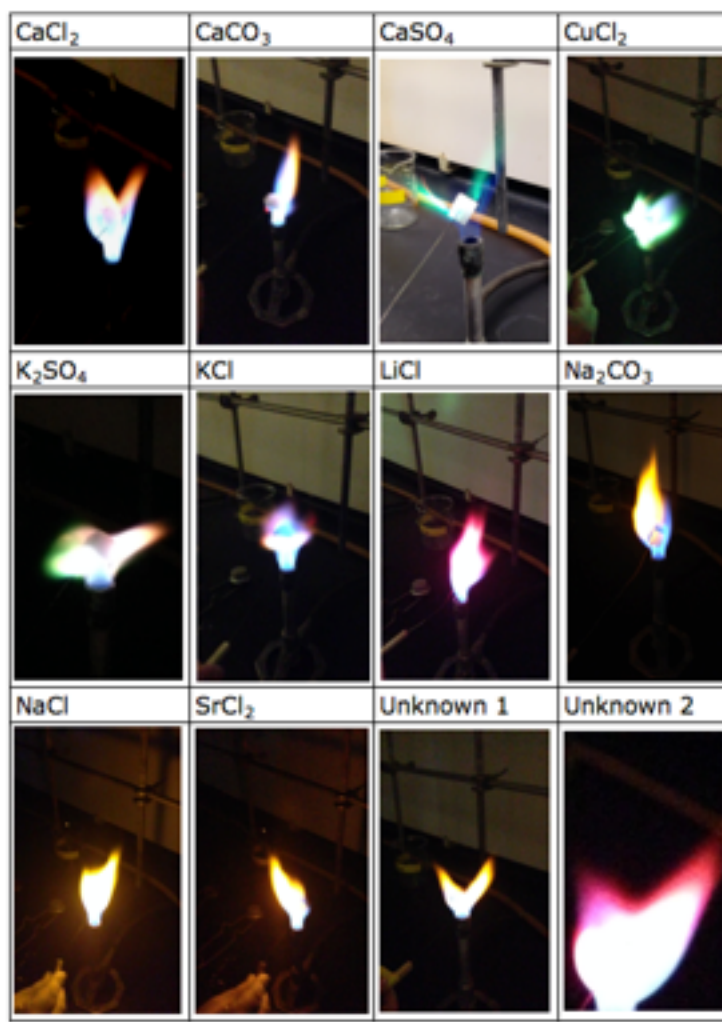


Table 1: Observed color of solution and corresponding flame color

Solution	Flame Color	Liquid Color
NaCl	bright orange	transparent
CuCl ₂	light green	turquoise
LiCl	magenta/red/orange/purple	transparent
KCl	light pink/orange	transparent
CaCl ₂	light orange	transparent
SrCl ₂	bright red/light orange/yellow	transparent
CaCO ₃	light orange	opaque(milky)
Na ₂ CO ₃	bright orange	transparent
K ₂ SO ₄	red/purple/pink/green/blue/orange	transparent
CaSO ₄	blue-green/turquoise	opaque(milky)

Unknown 1	dark red-orange	transparent
Unknown 2	magenta/orange	transparent

Discussion: Part I

The purpose of this section of the lab was to collect qualitative data relating to the flame color of different chemical compounds. Based on the collected data, there is no correlation between the color of the initial liquid and the color of the corresponding flame. For example, the two opaque liquids had distinctly different color signatures (light orange and blue-green). Likewise, although the majority of the solutions were transparent, the data illustrated the presence of color signatures from all areas of the visible light spectrum. Our data qualitatively illustrates the differences in spectroscopic emissions between different elements on the periodic table.

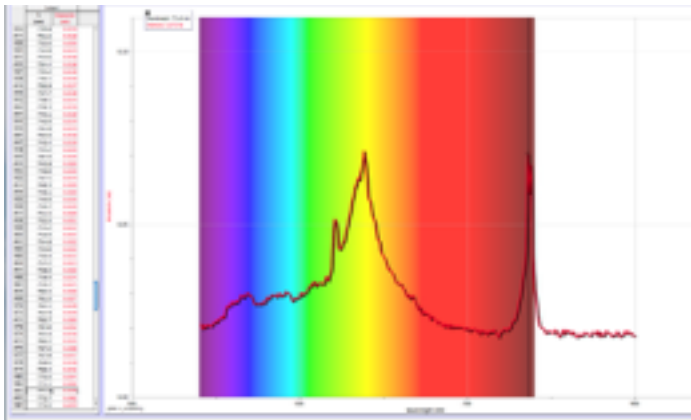
The cations, or metals (Ca, Cu, K, Li, Na, Sr), in the solution are the determiners of the color of the flame. Both NaCl and Na₂CO₃ have an identical flame color of bright orange; this is also the case of CaCl₂ and CaCO₃ that have an identical flame color of light orange (Table 1). CuCl₂, the only solution with the presence of copper as well as the only solution with a significant green hue to its flame, provides further evidence to this statement.

Based on the data that was collected during the experiment, Unknown 1 contains Strontium (Sr), as SrCl₂ was the only solution to display significant red hues in its flame. The presence of orange in the flame could be indicative of the presence of Cl, Cl₂, CO₃, or SO₄, as several solutions that include these elements/compounds burn with an orange hue. Unknown 2 likely contains Lithium (Li), as LiCl was the only solution to burn with a magenta hue. As with Unknown 1, the presence of orange in the flame could be explained in a variety of ways.

I am confident with these results because the colors of the samples' respective flames were visually documented for later reference. However, there is always the risk of contamination of a sample, should the melamine foam have been dipped in multiple solutions. This contamination would cause changes in our data, and ultimately render our findings useless.

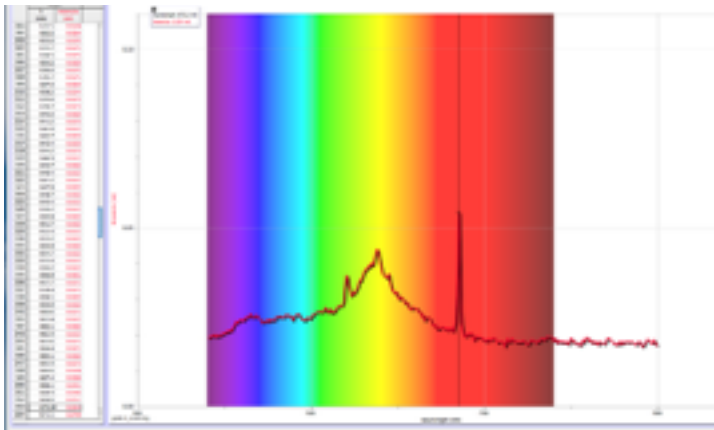
Results: Part II Spectrometer

Figure 5: KCl spectroscopic wavelength graph (intensity vs. wavelength)



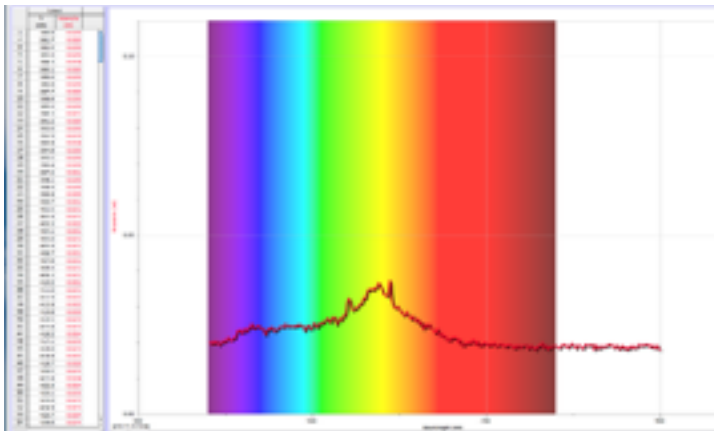
Peak wavelengths and corresponding intensity: 541.3 nm (.052 rel), 576.3 nm (.071 rel), 771.9 (.0721 rel)

Figure 6: LiCl spectroscopic wavelength graph (intensity vs. wavelength)



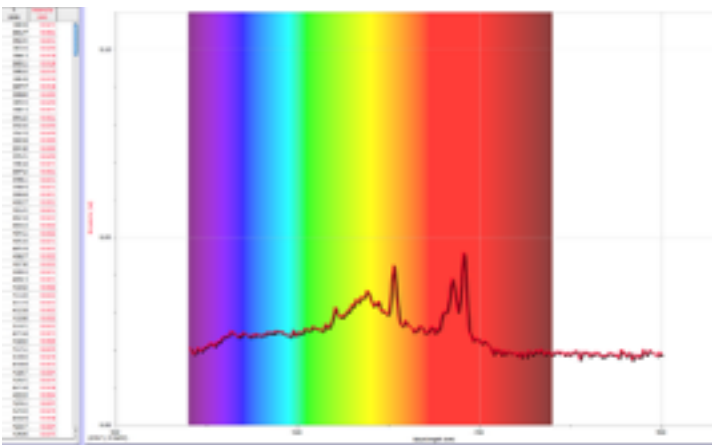
Peak wavelengths and corresponding intensity: 540.9 nm (.0367 rel), 576.3 nm (.0417 rel), 670.3 nm (0.054 rel)

Figure 7: NaCl spectroscopic wavelength graph (intensity vs. wavelength)



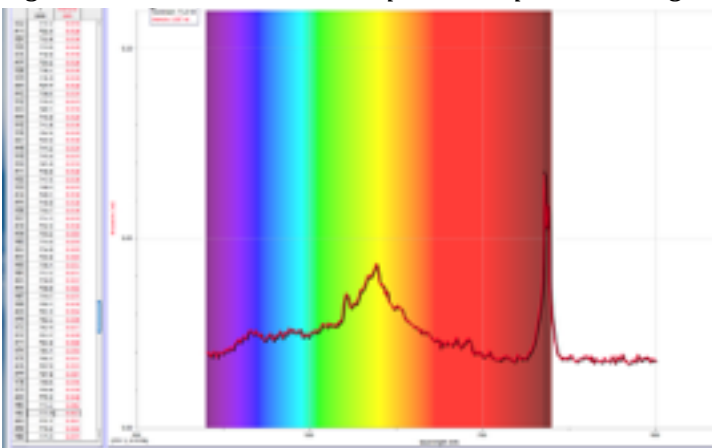
Peak wavelengths and corresponding intensity: ~542 nm (.034 rel), ~576 nm (.037 rel), ~581nm (.038 rel)

Figure 8: SrCl₂ spectroscopic wavelength graph (intensity vs. wavelength)



Peak wavelengths and corresponding intensity: ~541 nm (.031 rel), ~578 nm (.037 rel), ~605 nm (.043 rel), ~672 nm (.0395 rel), ~684 nm (.0455 rel)

Figure 9: Unknown solution spectroscopic wavelength graph (intensity vs. wavelength)



Peak wavelengths and corresponding intensity: 542.5 nm (.035 rel), 576.3 nm (.043 rel), 771.9 nm (.067 rel)

Discussion: Part II

The purpose of this section of the lab was to observe the qualitative spectroscopic emission spectra of substances, as shown through intensity vs. wavelength graphs, and to identify an unknown based on the intensity vs. wavelength graphs observed for known compounds. Based on the spectroscopic results, the unknown solution contains potassium, as it shares a peak at 771.9 nm with the KCl sample. The sample also likely contains Chlorine, as it shares a peak at 576 nm with both the KCl and NaCl samples. The unknown solution shows high levels of activity in the yellow, particularly at around 600 nm. The NaCl sample has a peak similar to the unknown sample at roughly 600 nm, leading me to believe that the unknown sample is KCl and NaCl. This data illustrates the correlation between flame color and intensity vs. wavelength graphs, as KCl was observed to burn with a color of pink, which would be on the red end of the visible spectrum, as is observed in the spike at 771 nm, and NaCl burned with a distinct orange color, confirmed by the wavelength peak at 581 nm in the orange section of the color spectrum.

Intensity vs. Wavelength graphs can be used along with atomic emission spectra to identify an element. On an intensity vs. wavelength graph, the wavelength represents the color of the line present on the atomic emission spectrum, and the intensity corresponds to how distinct the color line is. For example, for the KCl sample, the atomic emission spectrum would show very distinct lines in red and yellow, with a less distinct line in green at roughly 540 nm, 576 nm, and 772 nm, based on the spectroscopic data represented on the graph.

Though I am confident in my results, there are possible sources of error that should be recognized; the first is the amount of oxygen in the flame of the Bunsen burner. Flame needs oxygen in order to burn, but there is always the possibility that the oxygen in the flame corrupted our data. Oxygen burns with a bright yellow-orange flame, and, looking at the data, it is very possible that the activity in the yellow region of the spectrums could be the burning of oxygen along with our intended sample.

Results: Part III Spectrum Tube Analysis

Figure 10: Atomic Emission Spectrum of Unknown D

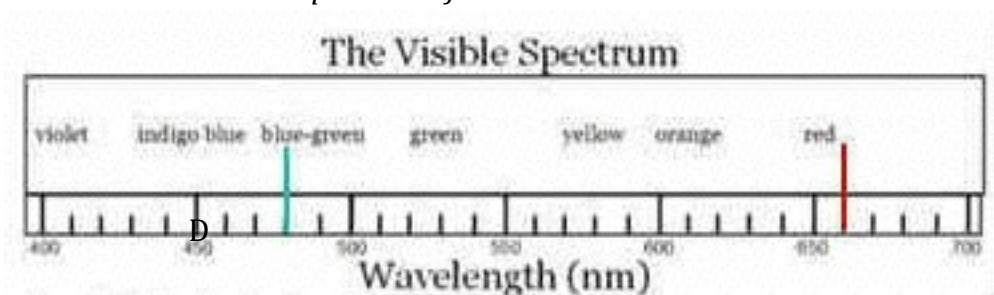


Figure 11: Atomic Emission Spectrum of Unknown E

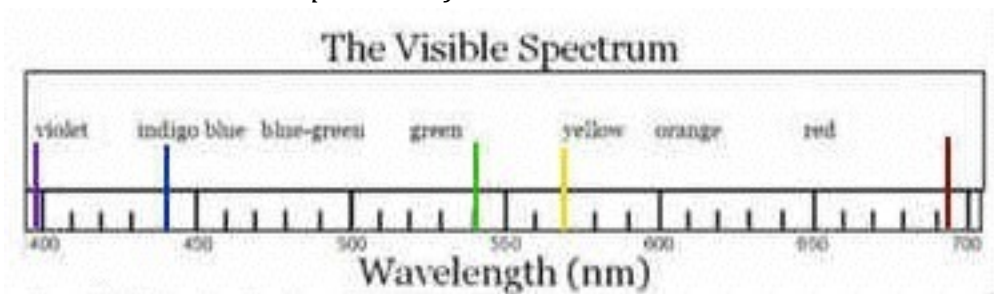
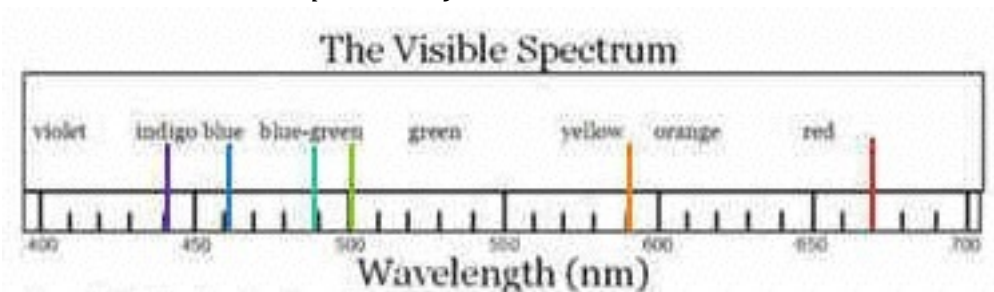


Figure 12: Atomic Emission Spectrum of Unknown F



Discussion: Part III

The purpose of this section of the lab was to observe the atomic emission spectrum of three unknowns, and be able to identify them through comparison to the known atomic emission spectrums of elements on the periodic table. Based on the atomic emission spectrum of Sample D, I believe its identity is that of hydrogen. The known atomic spectrum of hydrogen shows distinct and clear lines at 486 nm and 656 nm (Figure 13). This spectrum lines up exactly with the atomic emission spectrum found for Sample D (Figure 10). Sample E can be identified as magnesium. The known atomic spectrum of magnesium shows a red line at roughly 690 nm, a yellow line at 570 nm, a green line at 550 nm, a blue line at 440 nm, and a violet line at 400 nm (Figure 14). The results match with the data collected from Sample E, with the exception of two orange lines and two green lines that were not visible in our sample when viewed with a spectroscope (Figure 11). Sample F appears to be lithium. Our sample revealed a distinct red line at roughly 670 nm, a yellow line at roughly 590 nm, and a grouping of blue-green lines at 440 nm, 460 nm, 480 nm, and 500 nm, respectively (Figure 12). This data corresponds to the grouping and placement of color lines in the known spectrum of lithium (Figure 15). This data shows the credibility of using known spectroscopic emissions of elements on the periodic table to identify unknown elements. By comparing the unknowns to the known spectrums, it becomes possible to identify similarities between the two spectrums, and ultimately make a determination as to the identity of the unknowns.

In 1860, a spectrophotometer was used to observe the elements in Durkheim mineral water. When the atomic emission spectrum of the water was observed, the identifying lines of potassium, lithium, sodium, calcium, and strontium were found. When these elements were chemically removed from the water, their lines on the spectrum disappeared, as expected. However, after all of the known elements were removed from the water, two unexpected blue lines remained on the spectrum that did not match the spectrum of any known element. The discovery of unknown spectral lines indicated the discovery of a previously unknown element. The element name "cesium" came from the Latin word "*caesius*" meaning sky blue, to correspond with the colors of the lines on the element's emission spectrum.

I am reasonably confident with my identification of the unknowns in part 3 of this experiment. In the case of Unknown D, I feel very strongly that it is hydrogen, as hydrogen has a very simple, easily distinguishable emission spectrum. The other two unknowns, however, I am less sure of. A variety of variables could play a role in inaccurate data, such as marking the incorrect wavelength during initial observation of the samples atomic emission spectrum. When we first observed Unknown F, the plasma tube had become overheated, giving us an atomic emission spectrum resembling a continuous spectrum, or color gradient. When the plasma tube was exchanged, the emission spectrum was very weak, which could easily have led to a slight degree of error in the documentation of the data. Not only this, but emission spectrums can become very complex, making it difficult to identify one specific spectrum out of the entire periodic table. With Unknown E, identified as magnesium, the absent lines on the atomic emission spectrum we collected could be explained by

a couple different errors. Looking at the known spectrum of magnesium, it is easy to see that several of the lines are grouped closely together. An error could have occurred during observation, mistaking a pair of lines for a single line on the spectrum. Thus, instead of two yellow lines and two green lines, we observed one of each. There is also always the possibility of misidentification, as the observed atomic emission spectrum of the Unknown was only compared to a fraction of the known spectral emissions of elements on the periodic table.

Figure 13: Known atomic emission spectrum of Hydrogen



Figure 14:

Known atomic emission spectrum of Magnesium

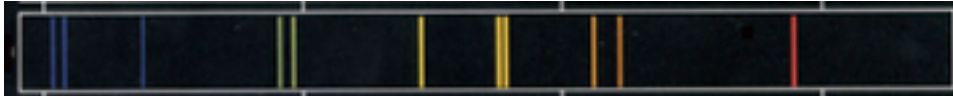


Figure 15: Known atomic emission spectrum of Lithium

