Experiment 01

Development of Bohr model due to atomic emission spectra of some s-block elements

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1.) Hypothesis
The imagination of the atomic structure has a long and steadily developing history. From the philosophical beginnings of Democritus over Dalton’s theorems to Rutherford’s orbit theory, more and more improvements were added to the model. However, all atomic models until this point saw the atom as a particle-alike object. In 1913 Niels Bohr introduced a new approach to the structural composition of an atom, in which he assumed that the electrons revolve around the nucleus on distinct radii, each of those radii due to their distance from the nucleus represent an energy level. These energy levels were not continuously, but existed only at discrete values. Each of these energy packs was called a quantum. The idea of quantized radiation was first introduce by Max Planck in 1900 and could explain many phenomena which physicists could not account for with common methods to that time. Bohr’s model was the first step to the wave model which will be discussed in chapter 5.

Bohr based his atomic model on the observation that atoms can emit light. This experiment shows how Bohr developed his atomic model from this property of atoms. To understand his conclusions made from this property, it is necessary to be familiar with electromagnetic (EM) radiation, because “light is a form of electromagnetic radiation”1.

Electromagnetic radiation is an electrical field and a magnetic field, which are travelling in wave form. A wave is characterised by its frequency (f) and its wavelength (λ). The smaller the wavelength of a wave the higher is its frequency. Electromagnetic waves can obtain a very broad range of frequencies. It reaches from frequencies around $10^{18}$ to frequencies of approximately $10^{20}$. The whole band divided into its frequencies is called the electromagnetic spectrum. Visible light is also part of the electromagnetic spectrum. It obtains wavelengths from 700 – 400 nm. The different colours of visible light are thereby related to specific

![Figure 1: The electromagnetic spectrum divided into categories due to frequency](image-url)
With a diffracting grating, white light, this contains every colour of visible light, can be split up into its constituent colours; each colour represents a specific frequency. The splitting up happens because the diffraction grating bends the waves of different wavelengths to different extents. "The lower the wavelength, the greater will be the angle of deviation."\(^2\) In view of the electromagnetic spectrum in figure 1 the angle of deviation of red light is therefore the lowest and that of violet light is the greatest. Figure 2 visualises the principle of light diffraction on a simple glass prism.

Electromagnetic radiation is a form of energy. The amount of energy a wave represents depends on the wavelength respectively frequency of the wave. "The lower the wavelength (and hence the higher the frequency), the greater is the energy"\(^3\) of the electromagnetic wave. Therefore, light, which is a form of electromagnetic radiation, is also a form of energy. Red light \((\lambda \approx 700\,\text{nm})\) has the lowest energy and violet light \((\lambda \approx 400\,\text{nm})\) has the highest energy of visible light.

As mentioned above it is known that atoms are able to emit light. As light is only a form of electromagnetic radiation, in which turn is a form of energy, atoms are able to emit energy.

Due to Newton’s law of energy conservation it is of course not possible that atoms can emit energy without energy being absorbed by them. So atoms only emit energy which they have obtained from an external source. According to Bohr, external energy is absorbed by the electrons around the nucleus. At the same time he assumed that “the electrons within the ... atom will not absorb or radiate energy so long as it stays in one of a” distinct “number of circular orbits”\(^6\). Each of those orbits represents an energy level. If an external source tries to transmit energy on an electron the electron can only absorb that energy if there is enough energy to “climb” onto a higher energy level. Because there is not a continuous amount of energy levels, that energy must have distinct values to move the electron onto a higher energy level. Because all atoms always try to be in the lowest possible energetic state, the electron falls back onto the lower energy level immediately. The energy which it has absorbed from
the external energy source is thereby emitted as electromagnetic radiation. If the energy of the emitted electromagnetic radiation corresponds to the energy obtained by frequencies of visible light, the human eye is able to detect that emission. If the electrons can only obtain discrete energy levels, the spectrum of the radiation emitted by an atom cannot be continuously but must only show distinct wavelengths: those wavelengths which correspond to the difference between two of those energy levels the electron can obtain.

2.) Diagram

![Diagram of the setup of the experiment to determine the atomic emission spectra of the elements Li, Na, Sr, Ba](image_url)

*Figure 3: Set up of the experiment to determine the atomic emission spectra of the elements Li, Na, Sr, Ba*

3.) Method

The experiment was split up into the following two steps:

Firstly, a hand-held spectrograph was built. Secondly, ionic compounds were burnt under light emission whereby the spectrograph was used to see the spectrum of the emitted light of the atoms.

3.1) Building of a hand-held spectrograph

The principle construction of a hand-held spectrograph is shown in figure 2. At one end of the cardboard tube a diffracting grating and on the other end a paper with a narrow slit (1 inch long and 0.5 mm wide) was fixed with tape. Thereby the paper with the narrow slit had to
cover the whole opening of this end of the tube, so that it was only possible to look into the tube through the slit. The long axis of the slit was aligned with the short axis of the grating.

3.2) Determination of the emission spectra of Lithium, Sodium, Strontium and Barium

3.2.1) Apparatus
Dilute hydrochloric acid, lithium salt (s), sodium chloride (s), barium chloride (s), strontium salt (s), Bunsen burner, tongs, nichrome wire

3.2.2) Procedure
The Bunsen burner was lit, so that the air hole was fully opened (nearly transparent flame). The wire was held into the flame to clean it. Once the wire had glowed, it was put into hydrochloric acid, so that it became totally clean. Then, the clean wire was dipped into the solid sodium chloride, so that some of the salt stuck to it. Afterwards the wire with the sodium chloride on its tip was held into the flame. Consequently the flame was coloured yellow-reddish until all the sodium chloride was used. Meanwhile the flame colouration was observed by looking through the spectrograph. This procedure was repeated with lithium salt, barium chloride and strontium salt instead of sodium chloride.

4.) Results
The spectrograph used for the experiment did unfortunately not work properly, so that the flame could only be observed by the naked eye. Those observations are shown in table 1.

<table>
<thead>
<tr>
<th>Element</th>
<th>Observed flame colour</th>
</tr>
</thead>
<tbody>
<tr>
<td>Lithium</td>
<td>Red</td>
</tr>
<tr>
<td>Sodium</td>
<td>Orange</td>
</tr>
<tr>
<td>Strontium</td>
<td>Red</td>
</tr>
<tr>
<td>Barium</td>
<td>Green</td>
</tr>
</tbody>
</table>

Table 1 shows the observed flame colourations for the four tested elements. The observations were made by the naked eye.

Figures 4 to 7 show therefore the observed atomic emission spectra of lithium, sodium, barium and strontium observed in external experiments.

*Figure 4: Atomic emission spectrum of lithium (Li)*

*Figure 5: Atomic emission spectrum of sodium (Na)*
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5.) Conclusion
The measured atomic emission spectra of lithium, sodium, strontium and barium do show not a continuous spectrum, but lines which correspond to a specific frequencies. Those observations confirm Bohr’s atomic model. Each line in the spectrum represents the energy released by an electron falling down from a high to a low energy level. However, Bohr used emission spectra of hydrogen atoms to develop his theory. The atomic emission spectrum of hydrogen looks much simpler than those of the four tested elements in this experiment (hydrogen emission spectra is shown in figure 8).

According to Bohr, the lines correspond to the energy released when an excited electron (an electron lifted by an external energy one or more energy levels) falls down to a lower energy level again. This can be expressed by the following formula:

\[ E_{\text{radiation}} = E_{\text{higher}} - E_{\text{lower}} \]

Where \( E_{\text{higher}} \) is the energy the electron obtains when being excited.

He labelled each energy level which the electrons could obtain with reference to the hydrogen emission spectrum with a number, the so called principle quantum number (\( n \)). Those principle quantum numbers are integers which increase with the distance of the energy level from the nucleus. The lowest energy level, which is nearest the nucleus, has therefore the principle quantum number 1; the second lowest energy level has the principle quantum number 2 and so. Each of those numbers represents a radius on which the electron orbits the nucleus. Those orbits are called shells. The energy levels described by the principle quantum number are not be confused with the lines in the spectrum. The actual energy value of the energy levels described by the principle quantum number is not the same to the energy corresponding to the lines in the spectrum; the latter are the product of an electron falling down from one energy level to another. The principle quantum number describes the actual energy levels the electrons obtain; the energy levels to which the electrons are lifted respectively fall down. \( E_{\text{higher}} \) and \( E_{\text{lower}} \) in the formula shown above can therefore be described by a principle quantum number, whereas \( E_{\text{radiation}} \) can not.
While this nomenclature worked very well for the hydrogen atom, it did not work as properly for the sodium emission spectrum for example which is a little bit more complicated (see figure 5). More lines are seen in the spectrum, which means there are more energy levels available in the sodium atom than in the hydrogen atom. Notwithstanding, there seemed to be an accumulation of energy levels around certain wavelengths. This led to the assumption that the points of accumulation represent a region of the shells described by the principle quantum number from the hydrogen atom, but that those shells described by the principle quantum number are divided into a distinct number of “sub-energy levels”, so called subshells. The numerical energetic value of the energy level described by the principal quantum number is not equal for every atom; it varies from element to element due to their proton numbers. It turns out that there are different numbers of subshells for different principle quantum numbers:

- “For the shell n = 1, there is only one subshell, labelled 1s.
- For the shell n= 2, there are two subshells, labelled 2s and 2p. The 2p subshell is at a higher energy than the 2s subshell.
- For the shell n = 3, there are three subshells, labelled 3s, 3p and 3d. These subshells increase in energy in the order $3s < 3p < 3d$.”

The existence of subshells could not be recognised with the hydrogen atom because all the subshells of a shell in a hydrogen atom have the same amount of energy; the hydrogen atom is therefore said to be degenerated.

The Bohr model was a very convincing way of explaining the atomic emission spectra of the elements. However, it still assumed the electrons to be particle-like objects. The Bohr model does not take into account that an accelerating charged particle radiates respectively steadily emits energy. That would lead into an energy loss of the electron and it would fall into the nucleus.

The first step to the development of the modern view of the atomic structure was done in 1801 by Thomas Young. With the so called “double-slit experiment” he could prove that light has wave-like properties. Thereby he sent light through two parallel, very small holes. Due to the very tiny diameters of the two holes, they acted like two “point sources” of light radiation. Therefore the light waves travel away from the holes in concentric circles. The experiment was arranged so that the waves which were emitted from the place of the holes had the same frequency and amplitude and were in phase with each other. On a screen behind the holes a picture of alternating bright (high intensity) and dark (low intensity) stripes was projected. Young concluded that this picture was a result of superposition and interference,
properties which only waves obtain. If light were a particle, the picture had shown only two regions of high particle density exactly on the points on the screen in line with the holes.

The same principle of wave proof was also carried out for electrons. The only problem for a “double-slit experiment” with electrons was that the holes needed to be much smaller than those for light. The diameter of the holes must be very small compared to the wavelength of the radiation to cause the wave to move in concentric circles from the holes. Only if this condition is fulfilled, a regularly interference pattern is visible on the screen. This is true for both the light and the electron “double-slit experiment”, but electrons have a much smaller wavelength than visible light. Consequently the holes must be much smaller than those used for light; that small indeed that only the distance that the space between the atoms in a crystal lattice was small enough to diffract the electrons in a sufficient way. Therefore an electron beam was sent through a crystal lattice and similar interference patterns to those of light were found. It was concluded that electrons as well must obtain wave properties.

However, it was observed that light not only obtained wave-like properties but also particle-like properties (wave-particle duality). One of the famoust observation in this context is the photoelectric effect. The term photoelectric effect describes the phenomenon that light can hit electrons out of solid metals. This fact alone is not very surprising, because it is known that a wave is a form of energy transport and therefore could be able to do that. However, in a classical wave only its intensity (i.e. the amplitude) of the wave determines the energy of the wave: the energy is not correlated to the frequency of the wave. During experiments with the photoelectric effect, it was found that not only the intensity of the wave but also the frequency determines how many electrons were hit out of the metal. It was even found that when the frequency of the radiation with which the metal was radiated sunk under a certain frequency, no electron was hit out of the metal at all, even if the intensity of the wave was extremely big. This totally contradicts the behaviour of a classical wave. All the energy of the wave is transmitted to a single electron, as if a particle would collide with the electron. The quasi “particles” of electromagnetic radiation are called photons.

There are other experiments with light that show that it has particle properties. A variation of the double slit experiment for example. If the intensity of the radiation is reduced to a point where only one photon at a time is released, small bright spots are discovered on the screen. If the points where those spots are observed are jotted down over a period of time, they form the same black-white pattern as in the original Young-experiment; it is the pattern of wave
interference. Figure 9 visualises this observation. (The same experiment was carried out with the same results with an electron beam.)

This observation means, where there is a maximum in the interference pattern observed in the double-slit experiment, most photons will be located; where there is a minimum in the interference pattern observed in the double-slit experiment no photon is located. This leads to the assumption that the probability of finding a photon at a place on the screen is related to the intensity of the wave at that place. However, it is not possible to predict the location of a single photon; there is only a probability of finding a photon at a place due to the intensity of the wave at this place. Where there is a high intensity of the wave, there is a higher chance of finding a photon than at a place where there is a lower intensity of a wave. Where the intensity of the wave is zero, there is no chance of finding the photon. The “wave-like aspects of a photon (or electron) is represented by a wave function, \( \phi \). \( \phi^2 \) gives the probability of finding a photon at a specific place.

In view of the wave-particle duality of light is also be true for electrons (“electron” itself is quasi a term which corresponds to the “photon” of light) Erwin Schrödinger developed a “mathematical model of the atom”. It is commonly known as the Schrödinger equation. With it it is possible to determine different values of the wave function. “Solutions of the Schrödinger equation account for the quantization of electronic energy levels”.

Solutions to the Schrödinger equation also let assume that there are regions around the nucleus where the chance of finding the electron of a specific energy level is higher than in other regions around the atom. The probabilities of finding an electron of a specific energy level can be added at points around the nucleus to derive a three-dimensional body which is structured into layers of the probability of finding an electron. The overall probability of finding an electron of a specific energy level within this three dimensional body is 90%. Those bodies are called atomic orbitals.

6.) Bibliography

6.1) Quotes
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Note 4, 5: http://134.91.165.4/Lehre/Material/PMAll/vorproben.pdf


6.2) Figures
Figure 1: http://en.wikipedia.org/wiki/Image:Electromagnetic-Spectrum.png
Figure 2: http://134.91.165.4/Lehre/Material/PMAll/vorproben.pdf

Figure 4, 5, 6, 7, 8: http://chemistry.bd.psu.edu/jircitano/periodic4.html

Figure 9: http://upload.wikimedia.org/wikipedia/commons/7/7e/Double-slit_experiment_results_Tanamura_2.jpg

6.3) Reference

2.) http://www.iap.uni-bonn.de/P2K/