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The Bohr Model of the Atom

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Bewick **Parsons**

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Authors
Sharon Bewick, Richard Parsons

Editor
Shonna Robinson

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Chapter 1

The Bohr Model of the Atom

1.1 The Nature of Light

Lesson Objectives

The student will:

- perform calculations involving the relationship between the wavelength and frequency of electromagnetic radiation, $v = \lambda f$.
- perform calculations involving the relationship between the energy and the frequency of electromagnetic radiation, $E = hf$.
- state the velocity of electromagnetic radiation in a vacuum.
- name at least three different areas of the electromagnetic spectrum.
- when given two comparative colors or areas in the electromagnetic spectrum, identify which area has the higher wavelength, the higher frequency, and the higher energy.

Vocabulary

- amplitude
- crest
- electromagnetic spectrum
- frequency (f)
- hertz (Hz)
- trough
- velocity (v)
- wavelength (λ)

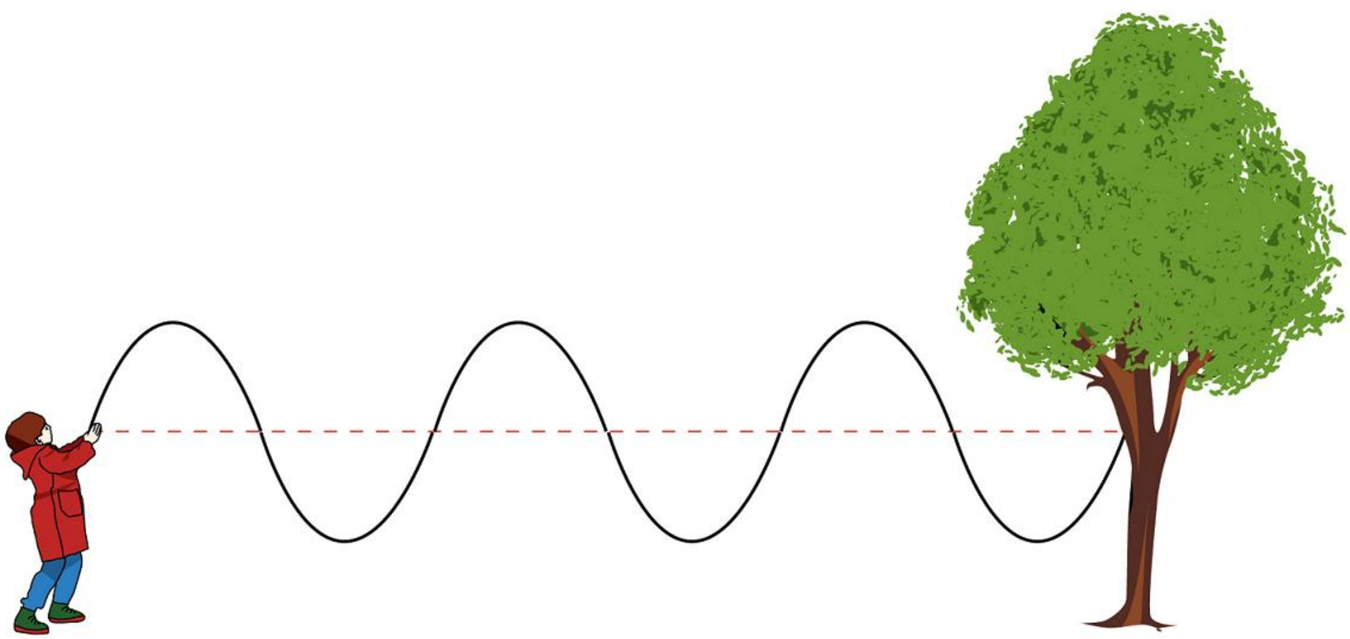
Introduction

In order to understand how Rutherfod's model of the atom evolved to the current atomic model, we need to understand some basic properties of light. During the 1600s, there was a debate about how light travels. Isaac Newton, the English physicist, hypothesized that light consisted of tiny particles and that a beam of light would therefore be a stream of particles. Around the same time, Christian Huygens, a Dutch physicist, suggested that light traveled as a waveform in the same way energy travels in water.

Neither hypothesis became the dominant idea until 200 years later, when the Scottish physicist James Clerk Maxwell proposed a wave model of light in 1864 that gained widespread support. Maxwell's equations related electricity, magnetism, and light so comprehensively that several physicists suggested students should major in other sciences because everything in physics had been discovered. Scientists thought that Maxwell's work permanently settled the "wave versus particle" debate over the nature of light. Fortunately, quite a few students did not take their advice. Sixty years later, German physicist Max Planck would raise the issue again and renew the debate over the nature of light.

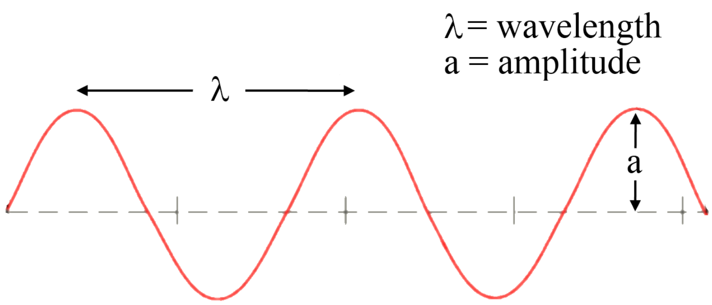
The Wave Form of Energy

The wave model of electromagnetic radiation is somewhat similar to waves in a rope. Suppose we tie one end of a rope to a tree and hold the other end at a distance from the tree so that the rope is fully extended. If we then jerk the end of the rope up and down in a rhythmic way, the rope will go up and down. When the end of the rope we are holding goes up and down, it pulls on the neighboring part of the rope, which will also go up and down. The up and down motion will be passed along to each succeeding part of the rope, and after a short time, the entire rope will contain a wave like the one shown in the image below.



The red line in the diagram shows the undisturbed position of the rope before the wave motion was initiated. The **crest** is the highest point of the wave above the undisturbed position, while the **trough** is the lowest point of a wave below the undisturbed position. It is important for you to recognize that the individual particles of rope *do not move horizontally.* Each point on the rope only moves up and down. If the wave is allowed to dissipate, every point of the rope will be in the exact same position it was in before the wave started. The wave in the rope moves horizontally from the person to the tree, but no parts of rope actually move horizontally. The notion that parts of the rope are moving horizontally is a visual illusion. Like the wave, the energy that is put into the rope by jerking it up and down also moves horizontally from the person to the tree.

If we jerk the rope up and down with a different rhythm, the wave in the rope will change its appearance in terms of crest height, distance between crests, and so forth, but the general shape of the wave will remain the same. We can characterize the wave in the rope with a few measurements. The image below shows an instantaneous snapshot of the rope so that we can indicate some characteristic values.



The distance from one crest to the next crest is called the **wavelength** of the wave. You could also determine the wavelength by measuring the distance from one trough to the next or between any two identical positions on successive waves. The symbol used for wavelength is the Greek letter lambda, λ. The distance from the maximum height of a crest to the undisturbed position is called the **amplitude** of the wave. The **velocity** of a wave is the distance traveled by the wave in one second. We can obtain the velocity of the wave by measuring how far a crest travels horizontally in a unit of time. The SI unit for velocity is meters/second.

Another important characteristic of waves is called frequency. The **frequency** of a wave is the number of cycles that pass a given point per unit of time. If we choose an exact position along the path of the wave and count how many crests pass the position per unit time, we would get a value for frequency. Based on this description, the unit for frequency would be cycles per second or waves per second. In science, however, frequency is often denoted by 1/s or s^{-1}, with "cycles" being implied rather than explicitly written out. This unit is called a **hertz** (abbreviated Hz), but it means cycles per second and is written out in calculations as 1/s or s^{-1}. The symbol used for frequency is the Greek letter nu, ν. Unfortunately, this Greek letter looks a very great deal like the italicized letter v. You must be very careful when reading equations to see whether the symbol is representing velocity (v) or frequency (ν). To avoid this problem, this text will use a lower case letter f as the symbol for frequency.

The velocity, wavelength, and frequency of a wave are all related, as indicated by the formula: $v = \lambda f$. If the wavelength is expressed in meters and the frequency is expressed in 1/second (s^{-1}), then multiplying the wavelength times the frequency will yield meters/second, which is the unit for velocity.

Example:

What is the velocity of a rope wave if its wavelength is 0.50 m and its frequency is 12 s^{-1}?

$$v = \lambda f = (0.50 \text{ m}) \cdot (12 \text{ s}^{-1}) = 6.0 \text{ m/s}$$

Example:

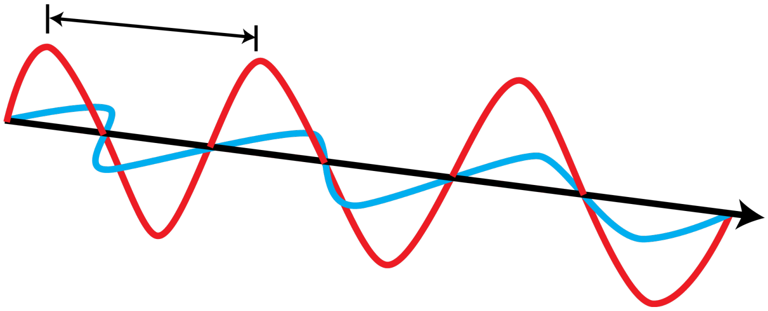
What is the wavelength of a water wave if its velocity is 5.0 m/s and its frequency is 2.0 s^{-1}?

$$\lambda = \frac{v}{f} = \frac{5.0 \text{ m/s}}{2.0 \text{ s}^{-1}} = 2.5 \text{ meters}$$

Electromagnetic Waves

Electromagnetic radiation is a form of energy that consists of electric and magnetic fields traveling at the speed of light. Electromagnetic waves carry this energy from one place to another and are somewhat like waves in a rope. Unlike the wave in a rope, however, electromagnetic waves are not required to travel through a medium. For example, light waves are electromagnetic waves capable of traveling from the sun to Earth through outer space, which is considered a vacuum.

The energy of an electromagnetic wave travels in a straight line along the path of the wave, just like the energy in the rope wave did. The moving light wave has associated with it an oscillating electric field and an oscillating magnetic field. Scientists often represent the electromagnetic wave with the image below. Along the straight-line path of the wave, there exists a positive electric field that will reach a maximum positive charge, slowly collapse to zero charge, and then expand to a maximum negative charge. Along the path of the electromagnetic wave, this changing electric field repeats its oscillating charge over and over again. There is also a changing magnetic field that oscillates from maximum north pole field to maximum south pole field. Do not confuse the oscillating electric and magnetic fields with the way light travels. Light does not travel in this weaving wave pattern. The light travels along the black line that represents the undisturbed position. For an electromagnetic wave, the crests and troughs represent the oscillating fields, not the path of the light.



Although light waves look different from the wave in the rope, we still characterize light waves by their wavelength, frequency, and velocity. We can measure along the path of the wave the distance the wave travels between one crest and the succeeding crest. This will be the wavelength of the electromagnetic radiation. The frequency of electromagnetic waves is still the number of full cycles of waves that pass a point in a unit of time, just like how frequency is defined for rope waves. The velocity for all electromagnetic waves traveling through a vacuum is the same. Although technically the velocity of electromagnetic waves traveling through air is slightly less than the velocity in a vacuum, the two velocities are so close that we will use the same value for the velocity. In a vacuum, every electromagnetic wave has a velocity of 3.00×10^8 m/s, which is symbolized by the lower case *c*. The relationship, then, for the velocity, wavelength, and frequency of electromagnetic waves is: $c = \lambda f$.

Example:

What is the wavelength of an electromagnetic wave traveling in air whose frequency is $1.00 \times 10^{14}\ s^{-1}$?

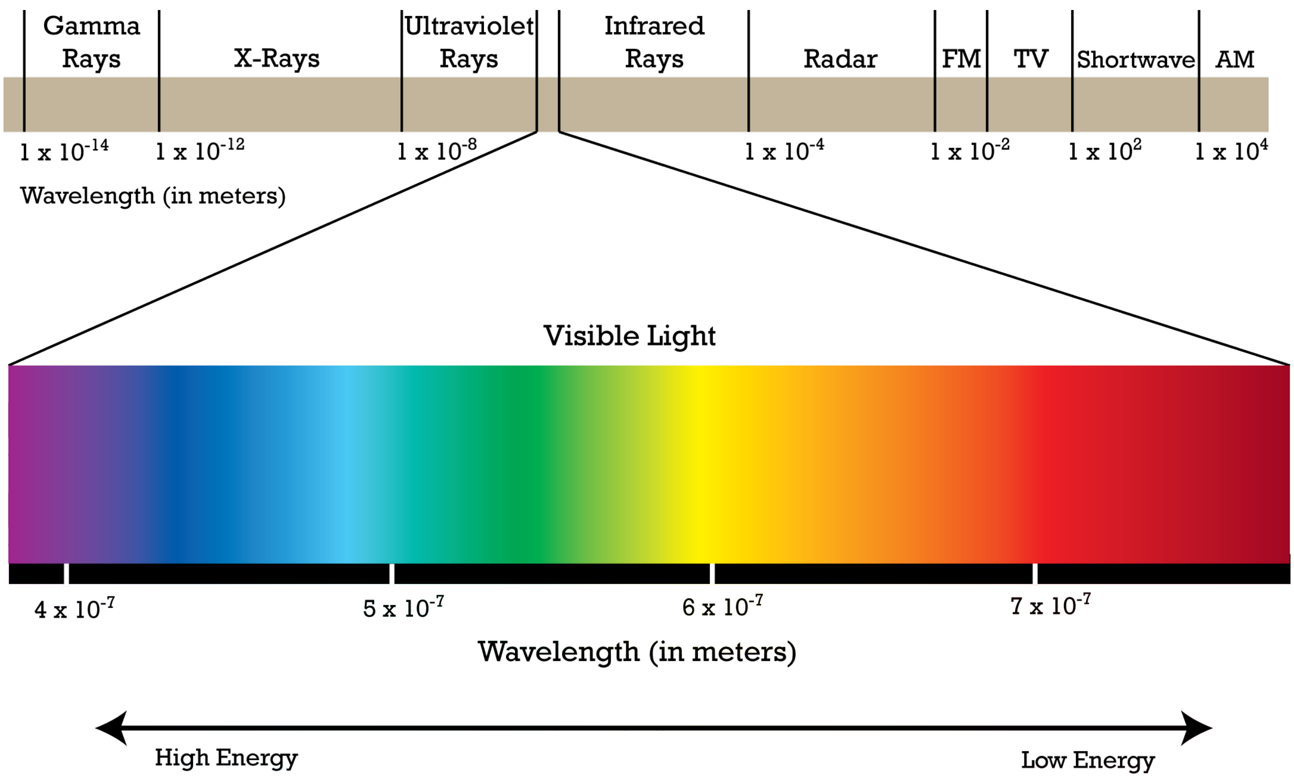
$$\lambda = \frac{c}{f} = \frac{3.00 \times 10^8\ \text{m/s}}{1.00 \times 10^{14}\ \text{s}^{-1}} = 3.00 \times 10^{-6}\ \text{m}$$

The Electromagnetic Spectrum

In rope waves and water waves, the amount of energy possessed by the wave is related to the amplitude of the wave; there is more energy in the rope if the end of the rope is jerked higher and lower. In electromagnetic radiation, however, the amount of energy possessed by the wave is only related to the frequency of the wave. In fact, the frequency of an electromagnetic wave can be converted directly to energy (measured in joules) by multiplying the frequency with a conversion factor. The conversion factor is called Planck's constant and is equal to 6.6×10^{-34} joule $\cdot$ seconds. Sometimes, Planck's constant is given in units of joules/hertz, but you can show that the units are the same. The equation for the conversion of frequency to energy is $E = hf$, where E is the energy in joules (symbolized by J), h is Planck's constant in joules $\cdot$ second, and f is the frequency in s^{-1}.

Electromagnetic waves have an extremely wide range of wavelengths, frequencies, and energies. The **electromagnetic spectrum** is the range of all possible frequencies of electromagnetic radiation. The highest energy form of electromagnetic waves is gamma rays and the lowest energy form (that we have named) is radio waves.

In the image below, the electromagnetic waves on the far left have the highest energy. These waves are called gamma rays, and they can cause significant damage to living systems. The next lowest energy form of electromagnetic waves is called X-rays. Most of you are familiar with the penetration abilities of these waves. Although they can be helpful in imaging bones, they can also be quite dangerous to humans. For this reason, humans are advised to try to limit as much as possible the number of medical X-rays they have per year. After X-rays, ultraviolet rays are the next lowest in energy. These rays are a part of sunlight, and rays on the upper end of the ultraviolet range can cause sunburn and eventually skin cancer. The next tiny section in the spectrum is the visible range of light. The band referred to as visible light has been expanded and extended below the full spectrum. These are the frequencies (energies) of the electromagnetic spectrum to which the human eye responds. Lower in the spectrum are infrared rays and radio waves.



The light energies that are in the visible range are electromagnetic waves that cause the human eye to respond when they enter the eye. The eye then sends signals to the brain, and the individual "sees" various colors. The waves in the visible region with the highest energy are interpreted by the brain as violet. As the energy of the waves decreases, the colors change to blue, green, yellow, orange, and red. When the energy of the wave is above or below the visible range, the eye does not respond to them. When the eye receives several different frequencies at the same time, the colors are "blended" by the brain. If all frequencies of visible light enter the eye together, the brain sees white, and if no visible light enters the eye, the brain sees black.

All the objects that you see around you are light absorbers – that is, the chemicals on the surface of the objects absorb certain frequencies and not others. Your eyes will then detect the frequencies that strike them. Therefore, if your friend is wearing a red shirt, it means that the dye in that shirt reflected the red

and absorbed all the other frequencies. When the red frequency from the shirt arrives at your eye, your visual system sees red, and you would say the shirt is red. If your only light source was one exact frequency of blue light and you shined it on a shirt that absorbed every frequency of light except for one frequency of red, then the shirt would look black to you because no light would be reflected to your eye.

Lesson Summary

- The wave form of energy is characterized by velocity, wavelength, and frequency.
- The velocity, wavelength, and frequency of a wave are related by the expression: $v = \lambda f$.
- Electromagnetic radiation comes in a wide spectrum that includes low energy radio waves and very high energy gamma rays.
- The frequency and energy of electromagnetic radiation are related by the expression: $E = hf$.

Further Reading / Supplemental Links

This website provides more information about the properties of electromagetic waves and includes an animation showing the relationship between wavelength and color.

- http://micro.magnet.fsu.edu/primer/java/wavebasics/index.html

Review Questions

1. Name at least three different areas in the spectrum of electromagnetic radiation.
2. Which color of visible light has the longer wavelength, red or blue?
3. What is the velocity of all forms of electromagnetic radiation traveling in a vacuum?
4. How can you determine the frequency of a wave when the wavelength is known?
5. If the velocity of a water wave is 9.0 m/s and the wave has a wavelength of 3.0 m, what is the frequency of the wave?
6. If a sound wave has a frequency of 256 Hz and a wavelength of 1.34 m, what is its velocity?
7. What is the relationship between the energy of electromagnetic radiation and the frequency of that radiation?
8. What is the energy, in joules, of a light wave whose frequency is 5.66×10^8 Hz?

1.2 Atoms and Electromagnetic Spectra

Lesson Objectives

The student will:

- describe the appearance of an atomic emission spectrum.
- explain why an element can be identified by its emission spectrum.

Vocabulary

- emission spectrum

Introduction

Electric light bulbs contain a very thin wire that emits light upon heating. The wire is called a filament. The particular wire used in light bulbs is made of tungsten. A wire made of any metal would emit light under these circumstances, but one of the reasons that tungsten is used is because the light it emits contains virtually every frequency, making the emitted light appear white. Every element emits light when energized, either by heating the element or by passing electric current through it. Elements in solid form begin to glow when they are sufficiently heated, while elements in gaseous form emit light when electricity passes through them. This is the source of light emitted by neon signs (see **Figure** 1.1) and is also the source of light in a fire. You may have seen special logs created for fireplaces that give off bright red and green colors as they burn. These logs were created by introducing certain elements into them in order to produce those colors when heated.

Figure 1.1: The light emitted by the sign containing neon gas (on the left) is different from the light emitted by the sign containing argon gas (on the right).

Each Element Has a Unique Spectrum

Several physicists, including Anders J. Angstrom in 1868 and Johann J. Balmer in 1875, passed the light from energized atoms through glass prisms in such a way that the light was spread out and the individual frequencies making up the light could be seen.

Figure 1.2: This is the unique emission spectrum for hydrogen.

In **Figure** 1.2, we see the emission spectrum for hydrogen gas. The **emission spectrum** of a chemical element is the pattern of frequencies obtained when the element is subjected to a specific excitation. When hydrogen gas is placed into a tube and electric current passed through it, the color of emitted light is pink. But when the light is separated into individual colors, we see that the hydrogen spectrum is composed of four individual frequencies. The pink color of the tube is the result of our eyes blending the four colors.

Every atom has its own characteristic spectrum; no two atomic spectra are alike. Because each element has a unique emission spectrum, elements can be identified by using them. **Figure** 1.3 shows the emission spectrum of iron.

You may have heard or read about scientists discussing what elements are present in the sun or some more distant star. How could scientists know what elements are present if they have never been to these faraway places? Scientists determine what elements are present in distant stars by analyzing the light that comes from those stars and using the atomic spectrum to identify the elements emitting that light.

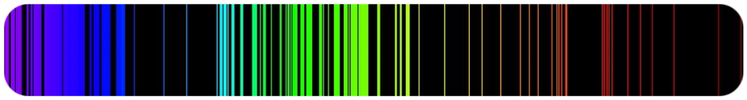


Figure 1.3: This is the unique emission spectrum of iron.

Lesson Summary

- Atoms have the ability to absorb and emit electromagnetic radiation.
- Each element has a unique emission spectrum.

Further Reading / Supplemental Links

This website "Spectral Lines" has a short discussion of atomic spectra. It also has the emission spectra of several elements.

- http://www.colorado.edu/physics/2000/quantumzone/index.html

Review Questions

1. The emission spectrum for an element shows bright lines for the light frequencies that are emitted. The absorption spectrum of that same element shows dark lines within the complete spectrum for the light frequencies that are absorbed. How can you explain that the bright lines in the emission spectrum of an element exactly correspond to the dark lines in the absorption spectrum for that same element?

1.3 The Bohr Model of the Atom

Lesson Objectives

The student will:

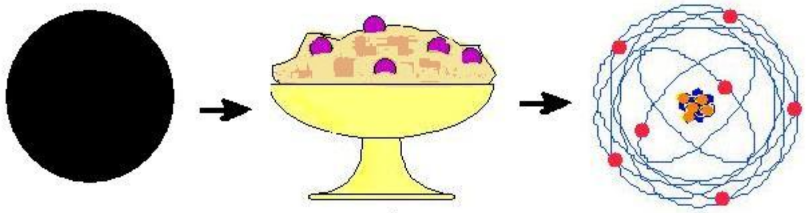
- describe an electron cloud containing Bohr's energy levels.
- describe how the Bohr model of the atom explains the existence of atomic spectra.
- explain the limitations of the Bohr model and why it had to be replaced.

Vocabulary

- energy level

Introduction

By 1913, our concept of the atom had evolved from Dalton's idea of indivisible spheres to Thomson's plum-pudding model and then to Rutherford's nuclear atom theory.



Rutherford, in addition to carrying out the experiment that demonstrated the presence of the atomic nucleus, proposed that the electrons circled the nucleus in a planetary-like motion. The planetary model of the atom was attractive to scientists because it was similar to something with which they were already familiar, namely the solar system. Unfortunately, there was a serious flaw in the planetary model. At that time, it was already known that when a charged particle moves in a curved path, the particle emits some form of light or radio waves and loses energy in doing so. If the electron circling the nucleus in an atom loses energy, it would necessarily have to move closer to the nucleus (because of the loss of potential energy) and would eventually crash into the nucleus. Scientists, however, saw no evidence that electrons were constantly emitting energy or crashing into the nucleus. These difficulties cast a shadow on the planetary model and indicated that it would eventually be replaced.

The replacement model came in 1913 when the Danish physicist Niels Bohr (pictured in **Figure** 1.4) proposed an electron cloud model where the electrons orbit the nucleus but did not have to lose energy.

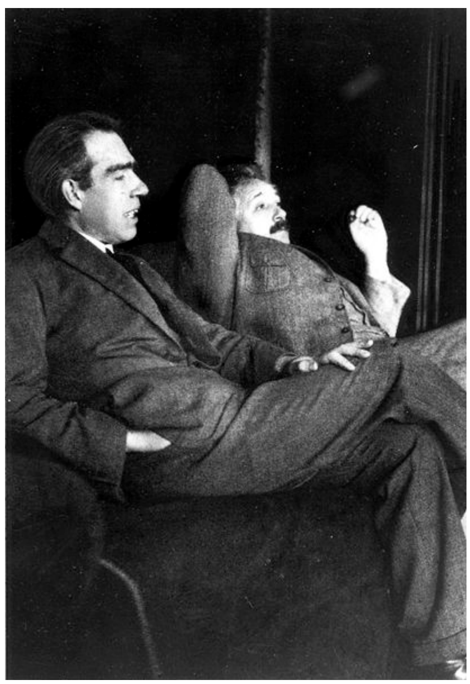


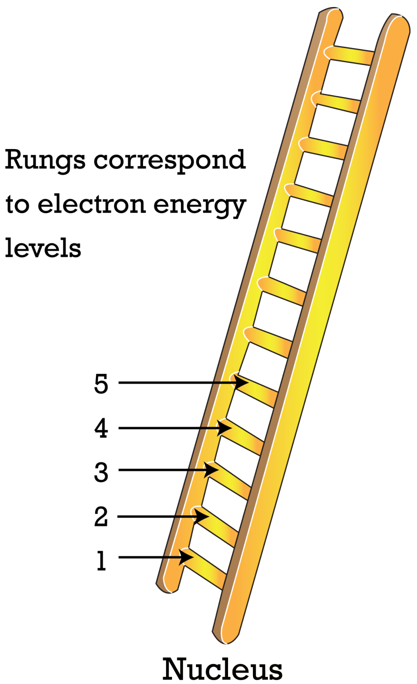
Figure 1.4: Niels Bohr and Albert Einstein in 1925.

Bohr's Energy Levels

The key idea in the Bohr model of the atom is that electrons occupy definite orbits that require the electron to have a specific amount of energy. In order for an electron to be in the electron cloud of an atom, it must be in one of the allowable orbits and have the precise energy required for that orbit. Orbits closer to the nucleus would require the electrons to have a smaller amount of energy, and orbits farther from the

nucleus would require the electrons to have a greater amount of energy. The possible orbits are known as **energy levels**.

Bohr hypothesized that the only way electrons could gain or lose energy would be to move from one energy level to another, thus gaining or losing precise amounts of energy. It would be like a ladder that had rungs at certain heights (see image below). The only way you can be on that ladder is to be on one of the rungs, and the only way you could move up or down is to move to one of the other rungs. Other rules for the ladder are that only one person can be on a given rung and that the ladder occupants must be on the lowest rung available. Suppose we had such a ladder with 10 rungs. If the ladder had five people on it, they would be on the lowest five rungs. In this situation, no person could move down because all the lower rungs are full. Bohr worked out the rules for the maximum number of electrons that could be in each energy level in his model. In its normal state (ground state), this would require the atom to have all of its electrons in the lowest energy levels available. Under these circumstances, no electron could lose energy because no electron could move down to a lower energy level. In this way, the Bohr model explained why electrons circling the nucleus did not emit energy and spiral into the nucleus.



Bohr Model and Atomic Spectra

The evidence used to support the Bohr model came from the atomic spectra. Bohr suggested that an atomic spectrum is created when the electrons in an atom move between energy levels. The electrons typically have the lowest energy possible, but upon absorbing energy, the electrons would jump to a higher energy level, producing an excited and unstable state. The electrons would then immediately fall back to a lower energy level and re-emit the absorbed energy. The energy emitted during these electron "step downs" would be emitted as light and would correspond with a specific line in the atomic emission spectrum. Bohr was able to mathematically produce a set of energy levels for the hydrogen atom. In his calculations, the differences between the energy levels were the exact same energies of the frequencies of light emitted in the hydrogen spectrum. One of the most convincing aspects of the Bohr model was that it predicted that the hydrogen atom would emit some electromagnetic radiation outside the visible range. When scientists looked for these emissions in the infrared region, they were able to find them at the exact frequencies predicted by the Bohr model. Bohr's theory was rapidly accepted and he received the Nobel Prize for

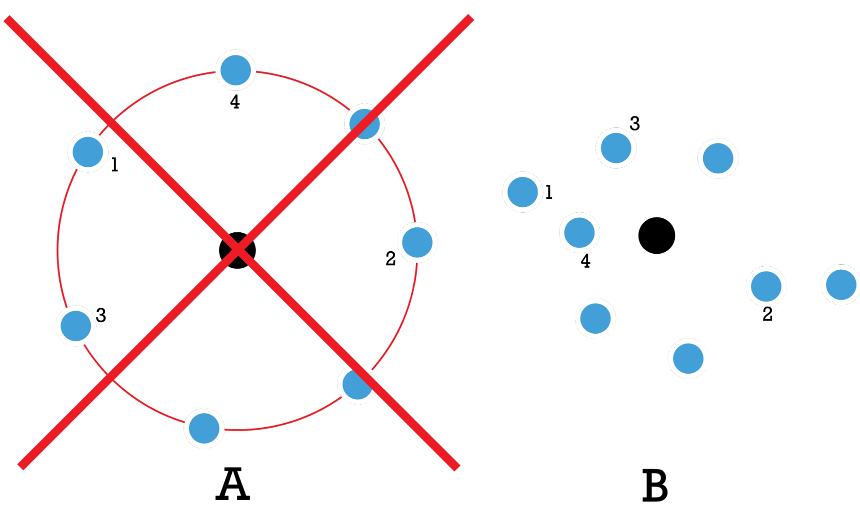
physics in 1922.

Shortcomings of the Bohr Model

The development of the Bohr model is a good example of applying the scientific method. It shows how the observations of the atomic spectra led to the creation of a hypothesis about the nature of electron clouds. The hypothesis also made predictions about emissions that had not yet been observed (the infrared light emissions). Predicted observations such as these provide an opportunity to test the hypothesis through experimentation. When these predictions were found to be correct, they provided evidence in support of the theory. Of course, further observations can also provide insupportable evidence that will cause the theory to be rejected or modified. In the case of the Bohr model of the atom, it was determined that the energy levels in atoms with more than one electron could not be successfully calculated. Bohr's system was only successful for atoms that have a single electron, which meant that the Bohr model did not accurately reflect the behaviors of most atoms.

Another problem with Bohr's theory was that the Bohr model did not explain why certain energy levels existed. As mentioned earlier in this lesson, at the time it was already known that charged particles emit some form of light or radio waves when moving in a curved path. Scientists have used this principle to create radio signals since 1895. This was the serious flaw in Rutherford's planetary model of the atom, which Bohr attempted to deal with by suggesting his electron cloud model. Although his calculated energy levels for the hydrogen were supported by hydrogen's emission spectrum, Bohr did not, however, explain why only the exact energy levels he calculated were present.

Yet another problem with the Bohr model was the predicted positions of the electrons in the electron cloud. If Bohr's model were correct, the electron in the hydrogen atom in ground state would always be the same distance from the nucleus. Although the actual path that the electron followed could not be determined, scientists were able to determine the positions of the electron at various times. If the electron circled the nucleus as suggested by the Bohr model, the electron positions would always be the same distance from the nucleus. In reality, the electron is found at many different distances from the nucleus. In the figure below, the left side of the image (labeled as A) shows the random positions an electron would occupy as predicted by the Bohr model, while the right side (labeled as B) shows some actual positions of an electron.



The Bohr model was not, however, a complete failure. It provided valuable insights that triggered the next step in the development of the modern concept of the atom.

Lesson Summary

- The Bohr model suggests each atom has a set of unchangeable energy levels, and electrons in the electron cloud of that atom must be in one of those energy levels.
- The Bohr model suggests that the atomic spectra of atoms is produced by electrons gaining energy from some source, jumping up to a higher energy level, then immediately dropping back to a lower energy level and emitting the energy difference between the two energy levels.
- The existence of the atomic spectra is support for the Bohr model of the atom.
- The Bohr model was only successful in calculating energy levels for the hydrogen atom.

This video provides a summary of the Bohr atomic model and how the Bohr model improved upon Rutheford's model **(1i; 1g I&E, 1k I&E)**: `http://www.youtube.com/watch?v=bDUxygs7Za8` (9:08).

Figure 1.5: (Watch Youtube Video)
http://www.ck12.org/flexbook/embed/view/362

This video describes the important contributions of many scientists to the modern model of the atom. It also explains Rutherford's gold foil experiment **(1g I&E)**: `http://www.youtube.com/watch?v=6773jO6fMnM` (9:08).

Figure 1.6: (Watch Youtube Video)
http://www.ck12.org/flexbook/embed/view/363

Further Reading / Supplemental Links

These various videos examine the components of the Bohr model of the atom.

- `http://www.youtube.com/watch?v=QI50GBUJ48s&feature=related`
- `http://www.youtube.com/watch?v=hpKhjKrBn9s`
- `http://www.youtube.com/watch?v=-YYBCNQnYNM&feature=related`

Review Questions

1. What is the key concept in the Bohr model of the atom?

2. What is the general relationship between the amount of energy of an electron energy level and its distance from the nucleus?
3. According to Bohr's theory, how can an electron gain or lose energy?
4. What happens when an electron in an excited atom returns to its ground level?
5. What concept in Bohr's theory makes it impossible for an electron in the ground state to give up energy?
6. Use the Bohr model to explain how an atom emits a specific set of frequencies of light when it is heated or has electric current passed through it.
7. How do scientists know that the sun contains helium atoms when no one has even taken a sample of material from the sun?

Image Sources

(1) *Bohr and Einstein*. Public domain.

(2) Images of neon (http://en.wikipedia.org/wiki/File:NeTube.jpg) and argon (http://en.wikipedia.org/wiki/File:ArTube.jpg) signs created by Pslawinski, created into a composite by Richard Parsons. *Neon and argon gas signs*. CC-BY-SA 2.5.

(3) *Iron emission spectrum*. Public domain.

(4) *Hydrogen Emission Spectrum*. Public domain.