Bohr's Model of the Atom

If you have seen a fireworks display, then you have experienced a fantastic example of chemistry in action. The different colours that appear in fireworks arise when electrons in the atoms of various chemicals become excited by absorbing electrical or thermal energy, and then emit that energy at various wavelengths. This seemingly simple explanation comes from decades of work on the structure of the atom by some of the world's most talented scientists.

Limits of the Rutherford Model of the Atom

In the previous section you read about some of the experiments that led to the discovery of the electron, proton, and neutron, and to Ernest Rutherford's model of the atom. The model of the atom proposed by Rutherford predicted that electrons move around the nucleus of the atom, much like planets orbit the Sun. This idea seemed reasonable because even though the Sun's gravity pulls planets toward it, this pull is counteracted by the planets' movement. It seemed reasonable that electrons orbiting an atomic nucleus would behave in the same way.

However, it became apparent that there was a problem with this idea. A body that is moving in an orbit is constantly changing direction, and a body that is changing direction or speed is accelerating. Physicists had demonstrated that when a charged particle accelerates, it continuously produces electromagnetic radiation (emitted as photons). According to classical light theory, an electron travelling in an orbit emits energy as photons and, therefore, loses energy. If an electron loses energy as it orbits, it should spiral in toward the positively charged nucleus (**Figure 1**). Since the electron is negatively charged and opposite charges attract, the atom would eventually collapse. However, this prediction is not supported by evidence. Generally, most atoms are stable and do not appear to be collapsing. This suggests that, although electrons are constantly moving, they do not lose energy. Rutherford's model, therefore, is not able to explain the stability of atoms.

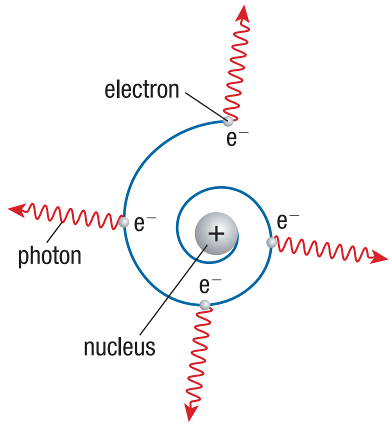


Figure 1 An electron accelerating around the nucleus would continuously emit electromagnetic radiation and lose energy. Therefore, it would eventually fall into the nucleus and the atom would collapse. However, this is not consistent with real-world observations.

Atomic Spectra

Spectroscopy is the scientific study of spectra (plural of spectrum) in order to determine properties of the source of the spectra. Spectrometers and spectrophotometers measure the intensity of light at different wavelengths in similar ways. Light first passes through a sample, and then is dispersed by a prism or, more commonly, a diffraction grating. The dispersed light forms a spectrum. A detector in the instrument then scans the spectrum and calculates the amount of light absorbed or transmitted at each wavelength. **Figure 2** shows an early spectroscope and a more modern spectrophotometer.

spectroscopy the analysis of spectra to determine properties of their source

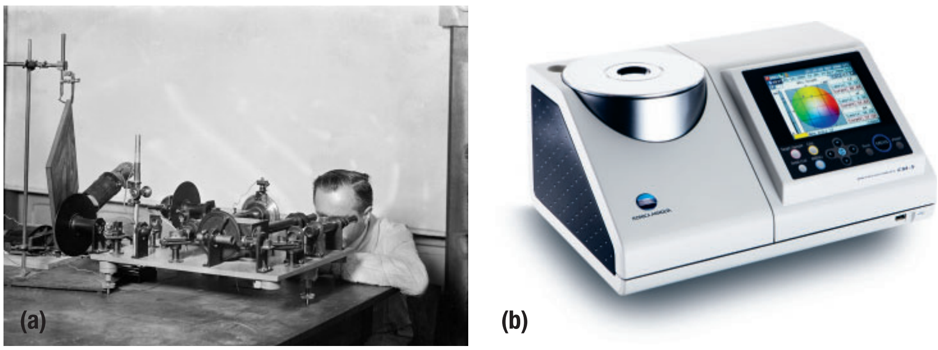


Figure 2 (a) Early spectroscopes used a candle and a gas lamp as light sources and focused light using a prism. (b) Modern spectrophotometers have a sealed area for the sample that does not allow interfering light to enter the unit, and can be adjusted to allow analysis at different wavelengths.

The earliest analytical instrument invented expressly for spectroscopy was a spectroscope similar to the one shown in Figure 2(a). Robert Bunsen and Gustav Kirchhoff invented the spectroscope to use in the first spectroscopy investigations, which they conducted in 1859. They viewed and analyzed the spectra produced by emission of energy by various substances, especially elements. As with most fields of scientific study, advances in spectroscopy dovetailed with advances in technology. Investigations of light emitted from excited substances led to further developments in atomic theory.

The Atomic Spectrum of the Hydrogen Atom

The atomic spectrum of the hydrogen atom played an important role in advancing atomic theory. Hydrogen gas, $H_2(g)$, is a molecular element. When a high-energy spark is applied to a sample of hydrogen gas, the hydrogen molecules absorb energy, which breaks some of the H–H bonds. The resulting hydrogen atoms are excited: they contain excess energy. The excited hydrogen atoms release this excess energy by emitting light of various wavelengths. When this light is passed through a spectroscope, it forms an emission spectrum. An **emission spectrum** is the spectrum (or pattern of bright lines) seen when the electromagnetic radiation of a substance is passed through a spectrometer.

emission spectrum the spectrum of electromagnetic radiation emitted by an atom; results when an atom is returned to a lower energy state from a higher energy state

continuous spectrum an emission spectrum that contains all the wavelengths in a specific region of the electromagnetic spectrum

line spectrum an emission spectrum that contains only those wavelengths characteristic of the element being studied

Two types of emission spectra can be produced, depending on the nature of the source. A **continuous spectrum** contains every wavelength in a particular region of the electromagenetic spectrum. For example, when white light passes through a prism, a continuous spectrum appears (**Figure 3(a)**) containing all the wavelengths of visible light. In contrast, a **line spectrum** contains only particular wavelengths, and arises when excited electrons emit energy. **Figure 3(b)** shows the line spectrum of the hydrogen atom. Each coloured band corresponds to a discrete wavelength.

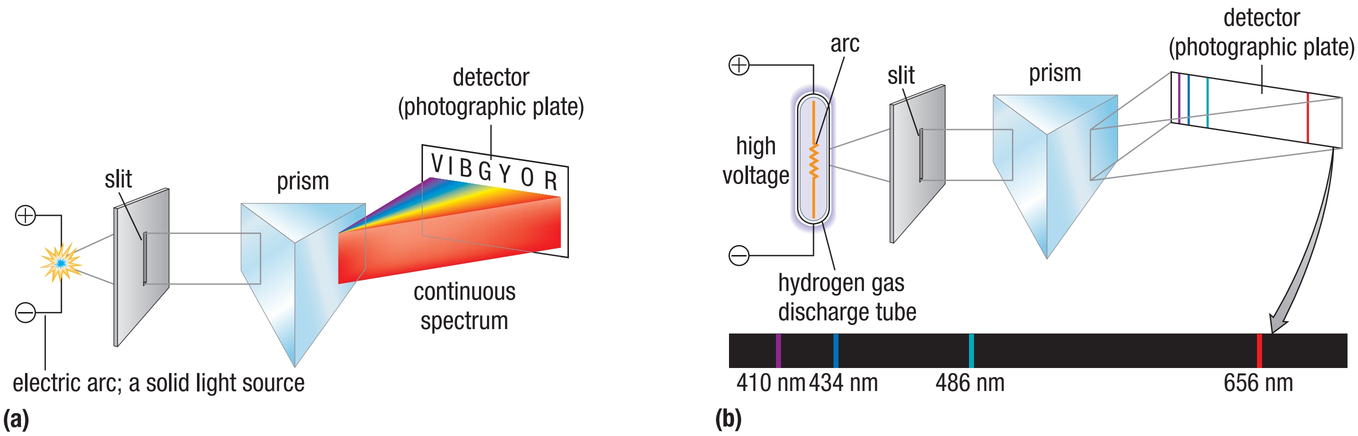


Figure 3 (a) A continuous spectrum contains all wavelengths of visible light (indicated by the initial letters of the colours of the rainbow). (b) The line spectrum for the hydrogen atom contains only a few discrete wavelengths.

Investigation 3.2.1

Bright-Line Spectra (page 180)
All atoms absorb and emit electromagnetic radiation. In this investigation, you will observe the visible spectra of various substances.

The investigations of Bunsen, Kirchhoff, and other scientists in the late nineteenth century revealed that each element has its own unique line spectrum. The spectra of the known elements were quickly catalogued. The line spectrum is like a fingerprint of a specific element. If a new spectrum was found, it provided evidence of a new element. In fact, the elements cesium and rubidium were discovered within a year of the invention of spectroscopy. There are many applications of line spectra. For example, astronomers use line spectra to identify the composition of stars.

The unique line spectrum of the hydrogen atom is significant to atomic theory because it indicates that the electron of the hydrogen atom can exist only at discrete energy levels. In other words, the energy of the electron in the hydrogen atom is quantized. This observation is consistent with Planck's quantum theory. The particular wavelengths of light emitted by the electrons of hydrogen atoms are produced by changes in energy. When excited electrons in hydrogen atoms move to a lower energy level, they emit a photon of light. This is true of excited electrons in other atoms as well.

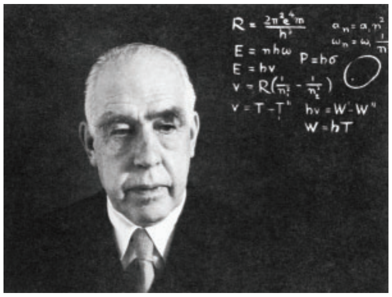
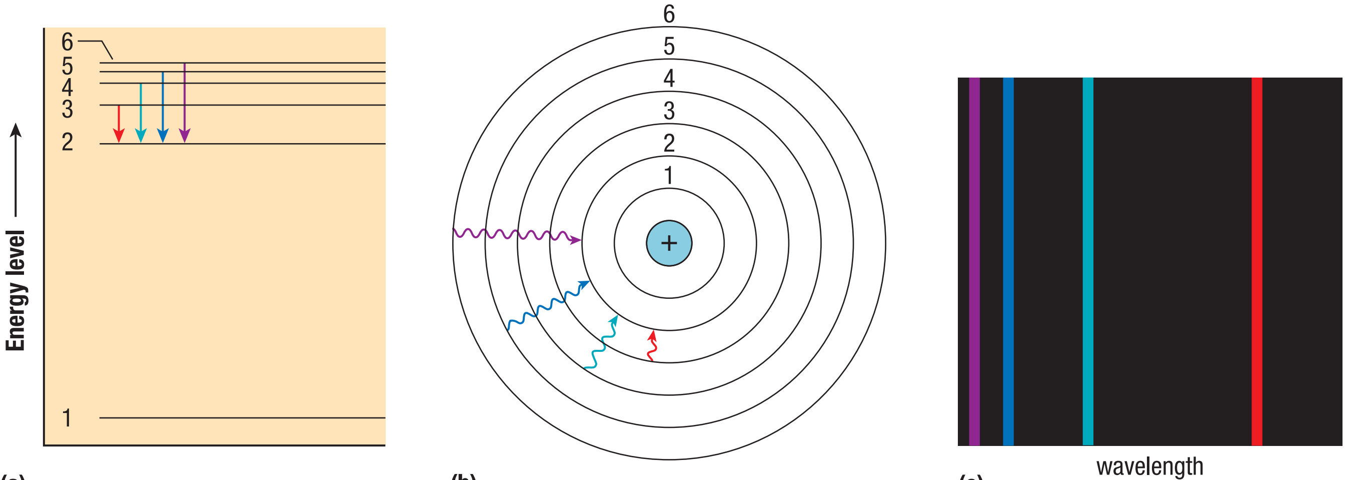


Figure 4 Niels Bohr (1885–1962) developed a quantum model for the hydrogen atom and, even though his model was later proved to be incorrect, Bohr was awarded the Nobel Prize in Physics in 1922.

The Bohr Model of the Atom

Niels Bohr was a Danish physicist who studied under J.J. Thomson at Cambridge University in the United Kingdom (**Figure 4**). In 1913, Bohr used the emission spectrum of the hydrogen atom to develop a quantum model for the hydrogen atom. He knew that his model had to account for the experimental evidence provided by spectroscopy: that electrons could have only particular discrete energy levels. Bohr accounted for this data by proposing that electrons could move only in specific orbits around the nucleus. He assigned each orbit a specific energy level, and postulated that the energy level of an orbit increased with its distance from the nucleus. When an electron gained more energy (for example, became excited), it could move into an orbit farther from the nucleus. Although Bohr's atomic model did not explain *why* electrons behaved this way, it was consistent with the observed line spectrum of the hydrogen atom. **Figures 5(a)** and **5(b)** show electron transitions in the Bohr model for the hydrogen atom. Compare these to the line spectrum of the hydrogen atom, shown in **Figure 5(c)**.



(a) **(b)** **(c)**

Figure 5 Electron transitions in the Bohr model for the hydrogen atom. An energy-level diagram (a) and an orbit-transition diagram (b), each showing electron transitions in the Bohr model for the hydrogen atom. Both of these account for the observed line spectrum of the hydrogen atom (c). The orbits are not drawn to scale. The lines in the visible region of the spectrum correspond to transitions from higher levels to level 2.

To help you envision how the orbits in Bohr's model relate to the line spectrum of the hydrogen atom, imagine a ball sitting on a staircase. Since the ball can only be positioned on a stair, it can only ever be found at specific distances from the ground. Applying Bohr's theory to this analogy, the higher up the staircase the ball is, the more potential energy it has. If the ball moves up the staircase (that is, to a higher energy level), it gains potential energy. If it moves down the staircase (that is, to a lower energy level), it loses potential energy. The ball in **Figure 6** is moving down the staircase, so it is losing potential energy.

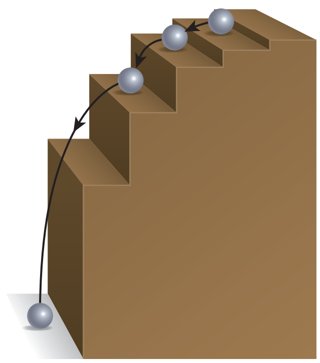


Figure 6 The position of the ball on the stairs determines its quantity of potential energy.

In the Bohr model of the atom, the electron is analogous to the ball in Figure 6 and the orbits are analogous to the different stairs. As with the ball on the stairs, electrons can only be at specific positions (energy levels or orbits) in relation to the nucleus of the atom. In **Figure 7**, the radius, r_x, of each orbit is analogous to the height of a stair from the floor in Figure 6. The electron gains or loses potential energy by moving from one position (orbit) to another.

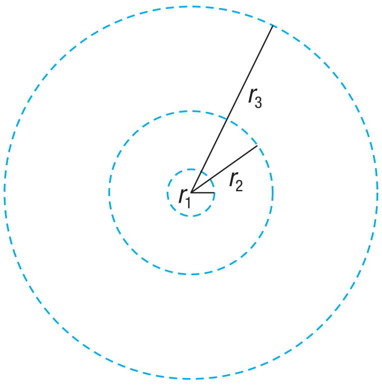


Figure 7 The position of an electron relative to the nucleus of an atom determines its quantity of potential energy.

transition the movement of an electron from one energy level to another

ground state the lowest energy state for an atom

The movement of an electron from one energy level to another is called a **transition**. During a transition to a higher energy level, an electron absorbs a specific quantity of energy, such as when it is struck by a photon. During a transition to a lower energy level, an electron emits a photon of a particular quantity of energy. The lowest possible energy state for an atom is called the **ground state**. There are no excited electrons in the ground state.

Successes and Failures of the Bohr Model

Recall that in a Bohr-Rutherford diagram, the numbers of protons, p^+, and neutrons, n^0, are noted in the nucleus. The concentric circles represent the different energy levels of electrons, and each contains a specific number of electrons. In an attempt to be consistent with observations related to the quantization of energy in atoms, Bohr's model assumes that each energy level can hold a maximum number of electrons. For the first 18 elements in the periodic table, the Bohr model predicts that the first, second, and third orbits can contain a maximum of 2, 8, and 18 electrons, respectively, and that the lower energy levels must fill first. The corresponding Bohr–Rutherford diagrams are especially useful for the first 20 elements of the periodic table, in which atoms of all the elements are arranged according to the number of protons and electrons in a neutral atom. Beyond the first 20 elements, however, Bohr–Rutherford diagrams become too cumbersome to be useful.

Bohr's model of the atom initially appeared to be very promising for understanding the behaviour of atoms because it appeared to be consistent with observed chemical and physical properties. For example, the energy levels Bohr calculated for the electron in the hydrogen atom were very similar to values obtained from the hydrogen atom's emission spectrum by spectroscopy. However, the electron energies predicted by Bohr's model were not consistent with observed data for atoms with more than one electron (**Figure 8**). Scientists eventually concluded that Bohr's model did not fully describe the structure of an atom. Still, the Bohr model is of great historic importance because it included the quantization of energy in atoms and paved the way for later theories. Bohr–Rutherford diagrams are so widely recognized, however, that it can be easy to forget that according to current theories of the atom, electrons do not actually orbit the nucleus.

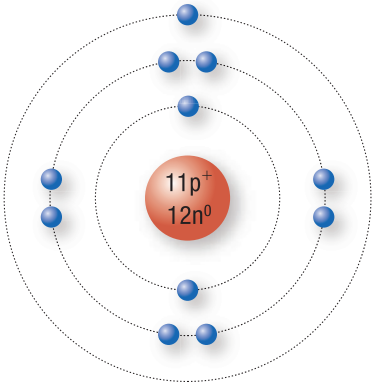


Figure 8 Electron energies for a neutral sodium atom, as predicted by the Bohr–Rutherford model. Current atomic theories do not support this arrangement of electron energies.

3.2 Review

Summary

- Spectroscopy is the study of light emitted by excited sources.
- Spectroscopy of excited gaseous elements led to the discovery of line spectra, which are unique to specific atoms and elements.
- Line spectra are consistent with Planck's quantum theory.
- Niels Bohr proposed a model of the atom that was consistent with experimental observations of the line spectrum of the hydrogen atom.
- In the Bohr model of the atom, electrons travel in circular orbits of quantized energy around the atomic nucleus.

Questions

1. Explain the main weakness with the Rutherford model of the atom and how Bohr addressed it. K/U
2. Describe what happens when atoms or molecules absorb light. K/U
3. Scientists use emission spectra to confirm the presence of an element in materials. Explain why this is possible. K/U
4. Using a series of diagrams, show what happens to the electrons of an atom when they are excited and how they can produce spectra that can be viewed in a spectroscope. K/U C
5. Explain why the emission spectrum of an atom or molecule depends on its arrangement of electrons. K/U
6. The emission spectrum of an element is unique. K/U A
   (a) Explain why the emission spectrum is sometimes referred to as an element's fingerprint.
   (b) Give a real-life example of how the emission spectrum could be used to help determine the nature of an unknown chemical.
   (c) Would using an emission spectrum be considered qualitative or quantitative analysis? Explain your answer.
7. In both ground-state sodium and magnesium atoms, the electrons are found in the first, second, and third energy levels. These electrons will jump to higher energy levels when energy is applied, and then fall back down, releasing their energy and giving off a spectrum. Why do you think the spectra for sodium and magnesium are not the same? Why might you think they would be the same? K/U T/I
8. (a) What is spectroscopy?
   (b) Discuss how spectroscopy was useful to the development of early atomic theory. K/U
9. (a) Describe the Bohr model of the atom, including quantization and emission spectra.
   (b) Discuss the successes and failures of the Bohr model. K/U
10. Why is the work of Bohr and Rutherford on atomic theory sometimes referred to collectively as the Bohr–Rutherford model? K/U
11. (a) What part of the original Bohr model still seems to be well supported by experimental evidence?
    (b) Identify one weakness in Bohr's atomic theory. K/U
12. When drawing the energy levels of an atom with one electron, your friend draws the diagram shown in **Figure 9**. Describe the line spectra that will be produced by an atom with this arrangement of energy levels. Provide evidence that this is not an accurate representation of the energy levels in an atom. K/U T/I

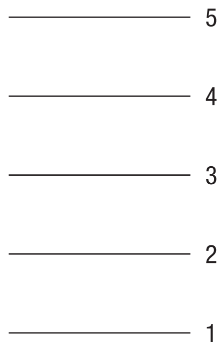


Figure 9

13. Using a graphic organizer or a chart,
    (a) compare and contrast Thomson's and Bohr's models of the atom
    (b) compare and contrast Bohr's and Rutherford's models of the atom K/U C
14. Summarize the evolution of atomic theory, starting with Thomson and ending with Bohr. Use a labelled diagram or a flow chart. (You may want to add this information to your answer to Question 6 in Section 3.1 to create a complete flow chart showing the development of atomic theory.) K/U C