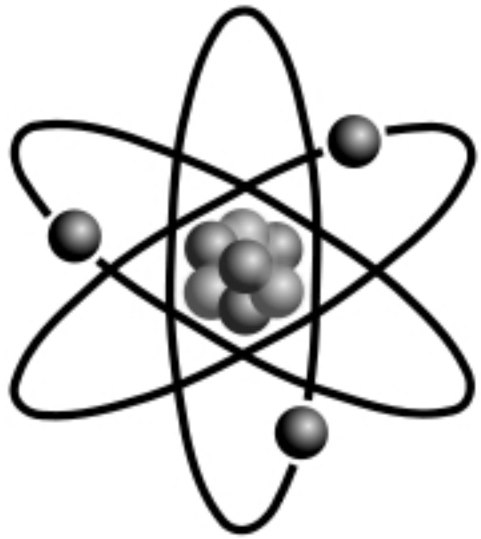
Historical Development of the Quantum-Mechanical Model

Rutherford visualized the atom as a miniature solar system, with electrons circling the nucleus in much the same way that planets orbit the sun.

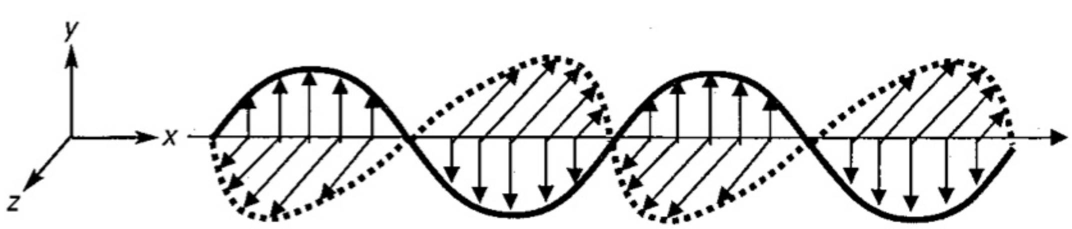


While this model is simple to understand, it is also inaccurate: electrons do not orbit the nucleus in well-defined paths. In this unit, we will learn how scientists believe the electrons are arranged in an atom. We will also examine the historical development of this modern model of the atom.

Electromagnetic Radiation

Much of our understanding of how electrons behave in atoms comes from the studies of how light interacts with matter. Thus, in order to understand the electrons, we must first consider some aspects of light.

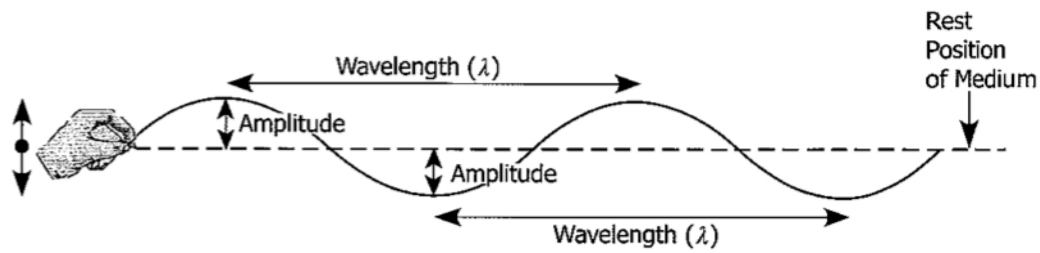
Light travels in waves, similar to the waves caused by a moving boat. Light waves, however, are electromagnetic waves, and light is a form of **electromagnetic radiation**. An electromagnetic wave consists of electric and magnetic fields oscillating at right angles to each other and to the direction of motion of the wave.



X-rays, gamma rays, and radio waves are other forms of electromagnetic radiation. All waves, whether they are water waves or electromagnetic waves, can be described in terms of four characteristics: amplitude, wavelength, frequency, and speed.

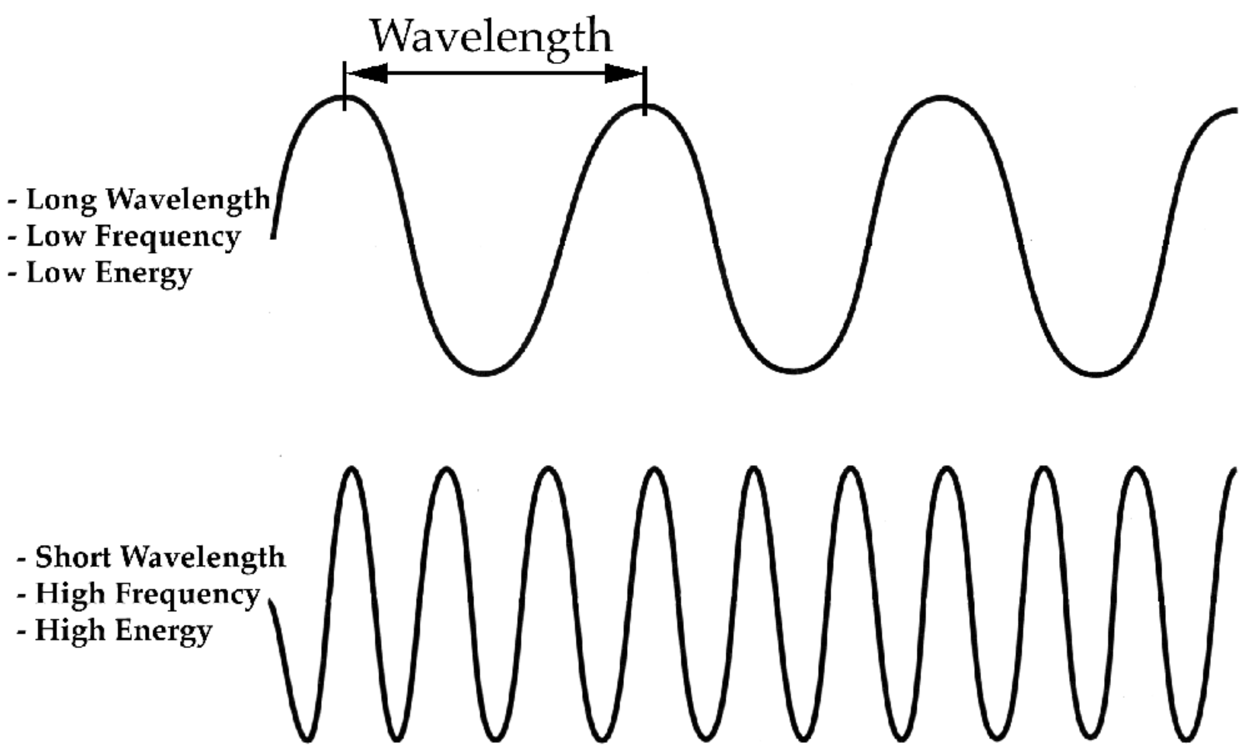
The **amplitude** of a wave is the height of the wave measured from the origin to its crest. The brightness, or intensity, of light depends on the amplitude of the light wave.

The **wavelength** is the distance between successive crests of the wave. The light that your eyes can see, called visible light, has wavelengths in the range of 400 to 750 nanometers.



The **frequency** of a wave is the number of times the wave completes a cycle of upward and downward motion in one second. The frequency of visible light varies from about 4×10^{14} to 7×10^{14} cycles per second.

Light, regardless of its wavelength, moves through space at a constant speed of $3.0\times10^{8}\ m/s$, which is the **speed of light**. Because light moves at a constant speed, there is a relationship between its wavelength and its frequency. As shown below, the shorter the wavelength, the greater the frequency.

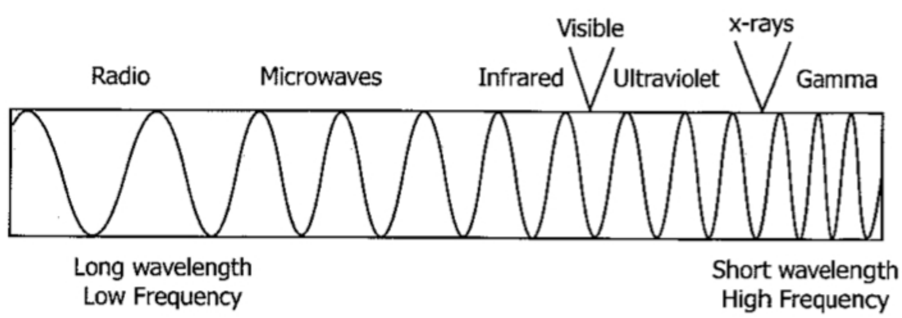


Electromagnetic Spectrum

When ordinary sunlight passes through a glass prism (or a raindrop), the light "spreads out" and is separated into the colors of the rainbow (ROYGBIV). This array of colors is called the **visible spectrum**. The visible spectrum is an example of a continuous spectrum because one color fades gradually into the next color.

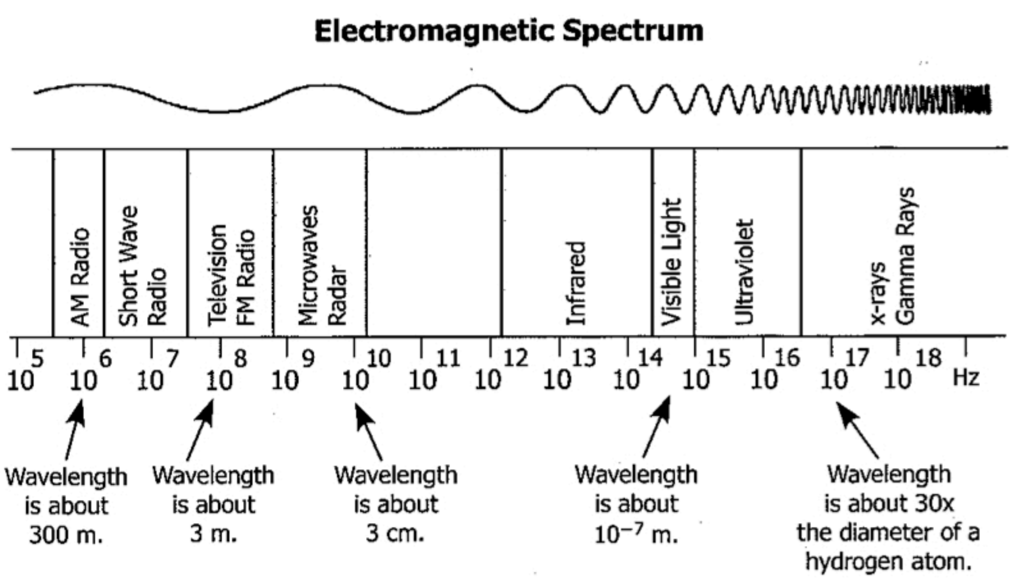
Visible light constitutes a very small portion of the total electromagnetic spectrum. The rest of the electromagnetic spectrum is invisible to the human eye. Microwaves, radio waves, and X-rays are all part of the electromagnetic spectrum.

The diagram below shows how the electromagnetic spectrum is arranged, with low-frequency waves on the left and high-frequency waves on the right.



Electromagnetic waves of different frequencies (and different wavelengths) have different names. This is because electromagnetic waves tend to have special uses or properties, depending on their frequencies.

The graphic below gives a more detailed look at the electromagnetic spectrum.



Within the visible portion of the spectrum, the different colors have different wavelengths (and therefore also different frequencies). Violet light has the shortest wavelength (and the highest frequency). Red light has the longest wavelength (and the lowest frequency).

Quantum Theory

By the start of the 20th century, the wave model of light was almost universally accepted by scientists. There were, however, some observations that they couldn't explain with the physics of the time.

One of these was the way that electromagnetic radiation is emitted by hot objects. For example, when a metal is heated, it first emits invisible infrared radiation (heat), but no visible light. Upon further heating, the metal begins to glow red, then yellow, and finally becomes "white hot." Scientists wondered why a different range of wavelengths is emitted at different temperatures.

Some other questions that scientists in those days could not answer included how elements such as barium and strontium gave rise to green and red colors when heated, and what caused neon lights to give off characteristic colors of light.

Planck's Theory

In 1900, Max Planck was able to predict accurately how the spectrum of radiation emitted by an object changes with its temperature. To do so, however, he had to make a radical suggestion. Planck proposed that the energy emitted or absorbed by an object was limited to certain specific amounts, what he called a **quantum** of energy. The word quantum means a fixed amount, and the plural of quantum is quanta.

In the case of electromagnetic radiation, there is a relationship between the frequency of a particular radiation and the energy with which it is associated. In other words, each color of light corresponds to a specific quantity of energy. In addition, higher frequency light has more energy than lower frequency light.

The Photoelectric Effect

Although Planck's discovery of energy quanta was a major turning point in science, it did not receive much attention at first. Even Planck was uncomfortable with the concept he had proposed. It was not until 1905, when Einstein used Planck's theories to explain another puzzling phenomenon, the photoelectric effect, that scientists really took notice.

In the photoelectric effect, electrons are ejected from the surface of a metal when light shines on the metal. For each type of metal, a minimum frequency of light is needed to release electrons. For example, red light is incapable of releasing electrons from sodium metal, even if the light is extremely bright. Violet light, on the other hand, releases electrons easily, even if it is very faint.

Einstein realized that Planck's idea of energy quanta was the key to understanding the photoelectric effect. He proposed that light consists of quanta of energy that behave like tiny particles of light.

He called these energy quanta **photons**. Einstein proposed that each photon carried a specific amount of energy that was determined by the frequency of the light.

Einstein stated that when a photon strikes the surface of a metal, it transfers its energy to an electron in a metal atom. He reasoned that the electron cannot just use part of the energy from the photon, and it cannot collect energy from several photons. Thus, an individual photon either had enough energy to release the electron, or it did not.

The important part of this idea is that the energy of the photon (based on its frequency) and not the number of photons (the intensity of the light) determines whether an electron is ejected or not. So why does violet light free electrons from sodium metal but red light does not? Violet light has a greater frequency and, as a result, a greater amount of energy per photon.

Dual Nature of Light

The idea that light consists of tiny particles, or photons, was conclusively proven in 1923 when Arthur Compton demonstrated that a photon could collide with an electron much as two tiny balls might collide. So a photon behaves like a particle, but a very special particle that always travels at the speed of light and has an associated frequency and wavelength.

The question of whether light is a particle or a wave is really not appropriate. Rather, light somehow has a dual nature, possessing the properties of both particles and waves.

Line Spectra

The fact that atoms can gain or lose energy only in chunks, or quanta, brings us closer to answering the question of how electrons are arranged in atoms. Before we can arrive at that answer, we must first consider another observation that puzzled scientists at the start of the 20^{th} century.

Samples of all elements emit light when they are vaporized in an intense flame or when electricity is passed through their gaseous state. When this light is passed through a prism, a spectrum is formed that contains only a few colors, or wavelengths. Such a spectrum is called a **line spectrum**, or an **atomic emission spectrum**.

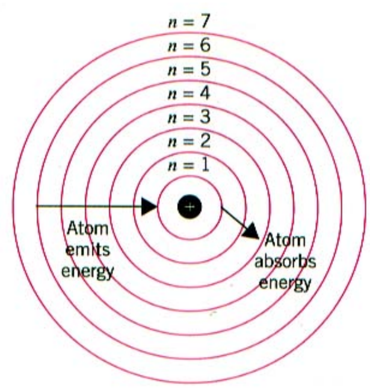
The atomic emission spectrum of each element is unique. Thus, line spectra can act as a fingerprint of sorts that can be used to identify elements.

Bohr Model of Hydrogen

In 1911, Bohr applied Planck's idea of quantization to Rutherford's model of the atom in order to explain the line spectra of elements. He started with the simplest atom, hydrogen, which only has one electron.

Bohr proposed that to get spectral lines, the energy of the electron in a hydrogen atom must be quantized. In terms of Rutherford's model, this means that the electron is only allowed to be in certain orbits corresponding to different amounts of energy.

Bohr labeled each energy level (each orbit) with a **quantum number**, n. For the lowest energy level, or **ground state**, $n=1$. This energy level corresponds to the orbit closest to the nucleus. When the electron absorbs the appropriate amount of energy, it jumps to a higher energy level, called an **excited state**. The excited states have quantum numbers $n=2$, $n=3$, $n=4$, and so forth. The excited states represent orbits that are farther from the nucleus, as shown below.



To explain hydrogen's spectral lines, Bohr proposed that when radiation is absorbed, an electron jumps from the ground state to an excited state. Radiation is emitted when the electron falls back from the higher energy level to a lower one. The color (frequency and energy) of the absorbed or emitted light is determined by the difference between the two energy levels involved.

Bohr's model worked well for hydrogen, but it was not able to explain the spectra of atoms with more than one electron. Despite this fact, the Bohr model represents an important step in the development of the current model of atomic structure.

Matter Waves

Until 1900, scientists believed that there was a clear distinction between matter and energy. As scientists came to understand the dual nature of light, however, they began to question whether or not matter might also have such a dual nature.

In 1924, Louis de Broglie proposed that particles of matter should behave like waves and exhibit a wavelength, just as waves of light behave like particles of matter. He referred to the wavelike behavior of particles as **matter waves**. Three years later, experiments performed by Clinton Davisson and Lester Germer confirmed de Broglie's predictions.

Today, the wave nature of electrons is used to magnify objects using electron microscopes. It is also used to determine the distances between atoms or ions.

Heisenberg's Uncertainty Principle

In 1927, Werner Heisenberg proposed his uncertainty principle, which states that the position and momentum of a moving object cannot simultaneously be measured and known exactly.

This principle is critical to the determination of atomic structure. To locate an electron, you must strike it with a photon or another particle, which then bounces back to some detection device. Because the electron has such a small mass, the collision moves it in some unpredictable way. So in this case, the very act of measurement changes what you're trying to measure.

One of the problems with the Bohr model is that there is no way to observe or to measure the orbit of an electron in an atom. In fact, the uncertainty principle suggests that it is not even appropriate to think of electrons as traveling in well-defined orbits.