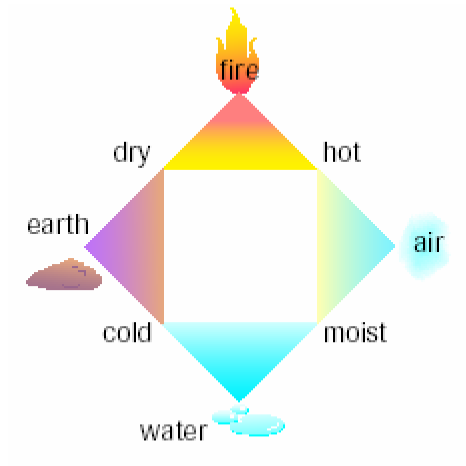
Guided Reading Activity - 1.3 Early Models of the Atoms - Page 13 through 15

Ancient Models of Matter



Empedocles in the fifth century BC, proposed that all matter was made up of the four basic particles of matter, **earth, fire, water** and **air**. Later on **Democritus**, doing a thought experiment came up with the idea that all matter was made up of tiny **indivisible particles**. These particles were so small that they could not be seen with the naked eye and they could not be cut or **divided** into smaller bits. He called these particles "**atoms**". Both Empedocles and Democritus came up with their models of matter without any **experimentation**. Experimentation gained importance during the Middle Ages, when **alchemists** were trying to turn base metals, such as copper and nickel, into precious metals, such as silver and gold. Methodical **observation** and **experimentation** were crucial parts of their quest. Through experimentation, they accumulated invaluable **empirical knowledge**, which led to more **theoretical knowledge** and more sophisticated equipment for conducting experiments.

Dalton's Atomic Theory
150 years ago John Dalton formulated the idea that the world around us is made up of large numbers of identical very small particles called **atoms**, and that the many types of different kinds of molecules are simply differently arranged groups of **atoms.** It is this idea of the molecule that lies at the heart of chemistry.

The understanding of the nature of matter which is called the atomic theory of matter, first postulated by John Dalton, is the basis of all modern chemistry. As stated by Dalton, the atomic theory of matter consists of five postulates: (fill in the blanks below)

1. **Matter consists of definite particles called atoms.**
2. **Each element is made up of its own type of atom.**
3. **Atoms of different elements have different properties.**
4. **Atoms of two or more elements can combine in constant ratios to form new substances.**
5. **Atoms cannot be created, destroyed, or subdivided in a chemical change.**

Sub Atomic Particles
In the 1880's advancements in technology like glass blowing, electricity and vacuums allowed for the creation of **cathod ray tube** J.J. Thomson, these tubes theorized that there was something smaller than the atom and that there was a small, **negatively**

charged particle inside the atom which were later called **electrons.** Thomson came up with a model which he called his **raisin bun** Model.
The **raisins** were the negatively charged electrons and the **bun** was the positive charged remainder of the atom.

Thomson's model of the atom was tested in 1911, when Ernest Rutherford conducted his famous **gold foil** experiment. From the results of his experiment, he concluded that an atom contains a **positively** charged **nucleus** surrounded by mostly **empty** space. Some of this space is occupied by negatively charged electrons . A few years later, Rutherford further concluded that the atom's nucleus consists of positively charged subatomic particles, which he named **protons.**

In 1932, James Chadwick discovered a third particle that had no charge, which he called a **neutrons.**
Neutrons are the same as a proton except that they have no **charge.**

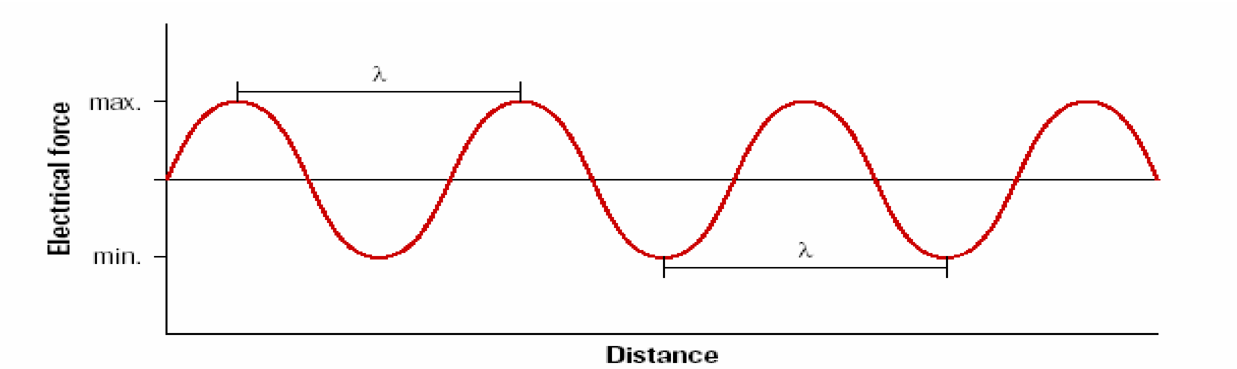
Problems with the Rutherford model
If the atom's nucleus is positively charged and the electrons surrounding it are negatively charged, and opposite charges attract, why do electrons not **collapse** into the nucleus? Even though Rutherford's atomic model was an improvement over previous atomic models, it could not explain this phenomenon. A new model was needed.

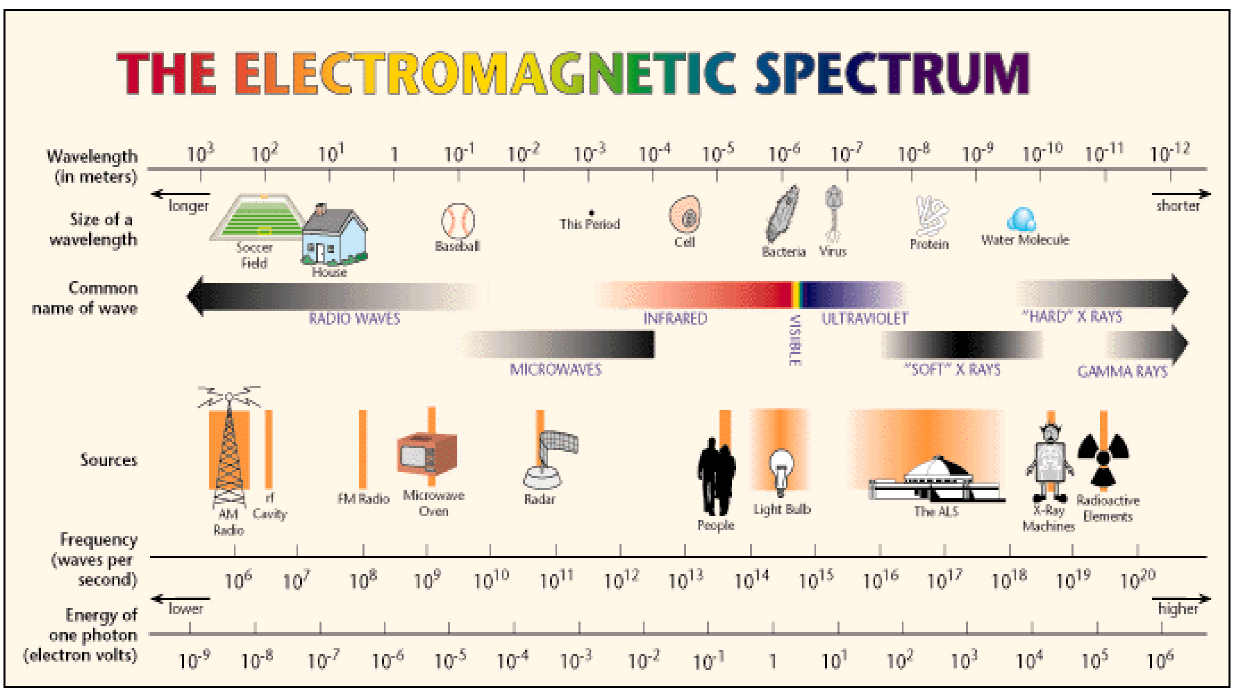
Fundamental Sub Atomic Particles of the Atom Summary

Particle Name	Location in Atom	Particle Symbol	Mass	Charge
Proton	**nucleus**	**p+**	**1**	**+ve 1**
Neutron	**nucleus**	**n**	**1**	**none**
Electron	**Around the nucleus**	**e-**	**1/2000**	**-ve1**

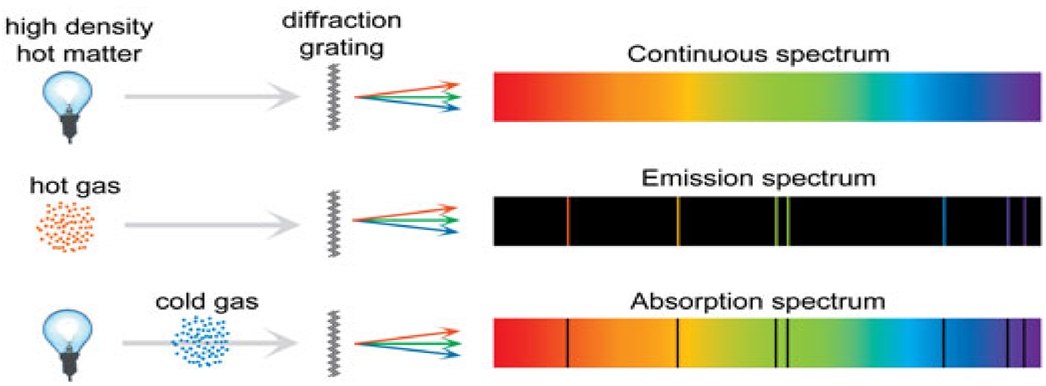
Guided Reading Activity - 1.4 The Electromagnetic Spectrum Page through

Electromagnetic energy is commonly known as **light energy**. Visible light, infrared light, ultraviolet light, and X rays are examples of light energy. Light energy is thought to move in the form of **waves**. One of the ways in which light waves differ from each other is **frequency**. Frequency is defined as **cycles** per unit **time**. The frequency of a light wave is the number of cycles that pass a point in one second. A wave has maximum and minimum values called **crests** and **troughs**, respectively. The distance between successive crests or successive troughs is known as the wavelength. The wavelength of visible light is usually measured in nanometres (nm). The symbol for wavelength is the Greek letter lamda (λ).





Another characteristic of light is speed. All forms of light travel at the same speed: 3.0x 10^8 m/s, or the speed of light. When the distance between the successive peaks of a light wave is short (the wave has a short wavelength), the time that the wave takes to pass a point is short, which means that the wave has a **high** frequency. Alternatively, when a light wave has a long wavelength, fewer cycles pass a point in a given time and the wave has a **low** frequency. A higher-frequency wave has more **energy** than a lower-frequency wave. The electromagnetic spectrum consists of light waves of different **frequencies** For example, radio waves have **low** frequencies. Therefore, they have **long** wavelengths and **low** energy. Radio waves do not pose a risk to humans. X rays are **high -** frequency, **high** -energy waves. Exposure to high-energy waves over a prolonged period of time can be harmful to human health. For this reason, an X-ray technician leaves the room when X-ray images are taken. Humans can detect certain types of light waves but not other types. For example, you can detect **infrared light** waves as heat. You cannot detect radio waves without a radio receiver, even though they pass above, below, around, and through you. The special band of light waves that the human eye candetect is known as the **visible** spectrum. The visible spectrum is composed of light waves with wavelengths in the range of **400 nm** to **700 nm.** Each wavelength of visible light is seen as a different **colour**. A rainbow contains all the colours in the visible spectrum and is an example of a continuous spectrum: a spectrum

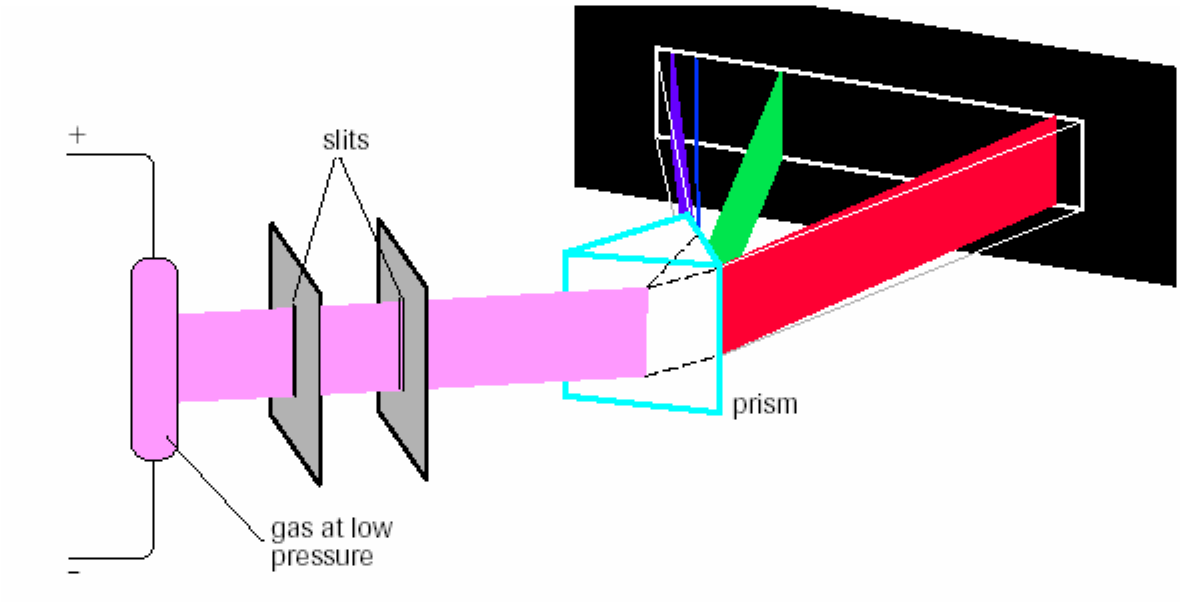


in which all the wavelengths of light are represented as an uninterrupted sequence. Different types of matter emit different wavelengths of light. Different types of matter also emit characteristic **line** spectra. Unlike a continuous spectrum, a line spectrum consists of **distinct coloured** lines rather than a rainbow. Analyzing the line spectrum of a sample is one way to distinguish between different types of matter. An element's line spectrum is analogous to a human's **finger print**. Line spectra can be observed using a **spectroscope** an instrument that separates light into its component colours using a prism or a diffraction grating. A line spectrum can be compared to playing a scale on a **piano**. A pianist can play tones and semitones, but no notes in between. A continuous spectrum is like playing a scale on a **violin**. A violinist can play all the intermediate notes between tones. A continuous spectrum is also analogous to a ramp, which provides a smooth transition from one level to the next. A line spectrum can be compared to a staircase, where the transition between levels is broken into steps.

REAL WORLD - Infrared Energy and Fever Doctors and nurses can measure a patient's temperature without touching the patient. They insert a hand-held device into the patient's ear and measure infrared (heat) energy. A person with a fever emits more infrared energy than a healthy person.

Guided Reading Activity - 1.5 Identifying Gases Using Line Spectra
pages ____ through ____

The light emitted by **__elements__** consists of many different **__wavelengths__** If this light is directed through a **__spectroscope__** a bright **__line__** spectrum is seen Since each element's line spectrum is **__unique__** line spectra can be used to **__identify__** known elements and to **__predict__** the existence of new elements. **__Spectroscopy__** is a branch of science in which light is used to identify and/or quantify substances.



When **__electricity__** is passed through a **__gaseous__** element, the atoms emit certain wavelengths of light. A **__spectroscope__** separates the mixture of wavelengths to produce a line spectrum. **The spectroscope may not show all the lines near the edges of the visible spectrum.**

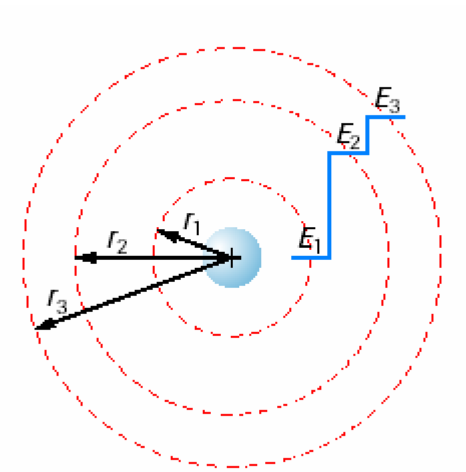
Guided Reading Activity - 1.6 The Bohr Model of the Hydrogen Atom pages ____ through ____

As you learned at the end of section 1.3, Rutherford's model of the atom **did not** explain why negatively charged electrons orbiting a positively charged nucleus do not **__collapse__** into the nucleus. The law of moving **__charges__** states that as an electron orbits the nucleus, it should emit energy in the form of electromagnetic radiation.(*and you should see a continuous spectrum….which you don't*) As the electron runs out of **__energy__** it should collapse into the atom's nucleus. Matter is very stable, however. Therefore, scientists have inferred that atoms are stable and that electrons do not collapse into the nucleus. As you saw in section 1.5, a spectroscope separates light into its component wavelengths, revealing a line spectrum that is unique to each element. Line spectra had been known and observed for many years, but no one could explain this phenomenon.

In 1913, **__Neils Bohr__** explained (**1) how line spectra were produced, but also (2) why electrons do not collapse into the atom's nucleus.**

Bohr's theory was based on the line spectrum of **<u>hydrogen</u>** because of the simplicity of hydrogen: the hydrogen atom possesses only one electron. Bohr suggested that electrons **<u>revolve</u>** around the atom's nucleus in **<u>orbits</u>** of **<u>fixed</u>** energy. These orbits are similar to the fixed orbits of the **<u>planets</u>** as they revolve around the Sun. The electrons are **<u>restricted</u>** to certain **<u>energy levels,</u>** and the energy of electrons is **quantized.** In other words, electrons must possess specific amounts of energy at each energy level.

Figure 3 page 21 - copy into space



Quantization is analogous to a **<u>ball</u>** sitting on a **<u>flight</u>** of steps. Since the ball cannot sit between steps, it is **<u>restricted</u>** to specific levels. As well, the ball can possess only a specific amount of energy, depending on which step it is sitting on. Since the ball cannot exist between steps, it cannot possess **<u>energies</u>** that are in **<u>between</u>** specific levels. When the hydrogen atom's **<u>electron</u>** absorbs light that has a specific energy, the electron jumps to a higher energy level. An electron that occupies a higher energy level is said to be in an **<u>excited</u>** state. Using the staircase analogy, if you want to kick a ball from the second step to the fourth step, you must kick it with precisely the right amount of energy. If you put too much energy into the kick, the ball may reach a height between the fourth and fifth steps and fall back down to the landing. If you put too little energy into the kick, the ball may go up only half a step and fall back down to its original position. Similarly, the orbiting electron must absorb electromagnetic radiation with just the right amount of energy in order to jump to a higher energy level.

Figure 5 page 22 - copy into space

(a) An electron gains a quantum of energy and jumps to a higher energy level.

(b) An electron loses a quantum of energy and drops to a lower energy level.

Figure 5

In order for the ball to fall back down to its original position on the staircase, it must **release** the same amount of energy that was required to raise it to the higher step. Similarly, when the hydrogen atom's electron falls back to its original position, it releases the same amount of energy it absorbed in order to reach the higher energy level.

According to Bohr, the lines in the **line** spectrum for hydrogen are a result of the energy released when the hydrogen atom's electron falls from a higher energy level to a lower energy level. If the electron is found in the lowest possible energy level, it is said to be in its **ground** state. Electrons in their ground state do not emit any **light** energy. Bohr's model of the atom was both a **success** and a **failure** It worked very well for explaining the line spectrum of hydrogen, but it did not work as well for the line spectra of other elements. Nevertheless, it was a breakthrough model because it introduced the idea that electrons orbit the nucleus in fixed energy levels. Furthermore, using evidence from the periodic table, Bohr predicted that each energy level could hold a **maximum** number of electrons. He predicted that the first energy level could hold **two** electrons, and the second energy level could hold **eight** electrons. The third energy level can hold **eighteen** electrons, but it is stable with **eight** electrons. You have used these numbers in previous courses to draw Bohr–Rutherford diagrams of atoms.

Guided Reading Activity - 1.7 Flame Tests - pages ____ through ____

When **electrons** in an atom are **excited** by heat or electricity, they **absorb** energy and jump from lower energy levels to higher energy levels. When an electron drops from an **excited** state to its **ground** state, the atom releases energy in the form of light. This light has a characteristic **colour**. Different types of matter produce different colours when subjected to high temperatures. Fireworks contain different compounds, which produce the different colours at a fireworks display.