1.2: Atomic Theories

Dalton’s Atomic Theory (1803)
By the late 1700’s, people experimenting in chemistry had discovered a great deal of empirical knowledge, which they called laws. Examples of these laws included the law of conservation of mass and the law of definite proportions. In 1803 an English schoolteacher named John Dalton proposed a theory that explained the empirical evidence. His theory included the following ideas:

- All matter consists of very tiny particles called atoms.
- Each element is made up of its own type of atom.
- Atoms of different elements have different physical and chemical properties.
- Atoms of two or more elements can combine in constant ratios to form new substances.
- Atoms cannot be created, destroyed, or altered during a chemical reaction.

To help picture the Dalton theory, think of a smooth hard sphere, like a marble or a billiard ball.

Dalton’s model was a good explanation of the facts of the time, but in time things changed.

Thomson’s Atomic Theory (1897)
During the Industrial Revolution of the 1700’s and 1800’s technology improved and scientists had access to increasingly sophisticated technology which they used to probe matter. In 1897 J. J. Thomson conducted experiments with his cathode ray tubes. He discovered that the atom wasn’t the smallest particle of matter: there were smaller charged particles inside the atom!

Thomson revised Dalton’s model of the atom. In Thomson’s model the atom is like a chocolate chip muffin. Inside the muffin are tiny negatively charged particles, called electrons. Since the atom is neutral overall, Thomson decided that the dough represents a positively charged material.

Rutherford’s Atomic Model (1911)
In 1911 Ernest Rutherford tested Thomson’s atomic model with his famous gold foil experiment. Rutherford discovered that an atom actually has a small positively charged nucleus, surrounded mostly by empty space. Some of the space is filled by the electrons. Rutherford later discovered that the positive charge in the nucleus is caused by small subatomic particles called protons. Later, in 1932 James Chadwick modified Rutherford’s theory after discovering another particle in the nucleus called the neutron. Neutrons have the same mass as protons but no charge.
Rutherford’s model is like a beehive. The nucleus is made of a dense positive core of protons and neutrons, while the electrons fly around it like bees buzzing around a hive.

**The Bohr Model of the Hydrogen Atom**

Scientists had observed that different gases emit different patterns of light when a large voltage is applied to them.

A Danish physicist named Neils Bohr developed a new atomic theory to explain the patterns of lines observed in the emission spectrum of gas discharge tubes. Bohr suggested that the electrons were going around the nucleus in orbits of fixed energy. An electron in one of these orbits had a specific amount of energy. The further the orbit was from the nucleus the more energy the electron had. A good analogy to this is a flight of stairs. The higher you climb up the stairs the more energy it takes to do this.

When electrons absorb just enough energy, such as from heat or electricity, they can jump up from one level to a higher energy level. The more energy they absorb the higher they can jump. When this happens the electron is in an **excited state**. To jump to a higher energy level the electron must gain just the right amount of energy to make the jump.

Atoms do not stay excited for very long. Eventually the electron will fall back down to its original energy level. When it does this it will lose energy. This energy is released as particles of light energy called **photons**. The amount of energy released determines the colour of the light. If it is in the visible part of the spectrum we can detect it with our eyes. Since the energy released is in very specific amounts, we will only observe very specific colours in the spectrum of light. When an electron falls back to its lowest energy level it is said to be in the **ground state**.
According to this model each energy level could hold only so many electrons. The first orbit holds only two electrons, the second orbit holds eight electrons, and the third energy level can hold 18, but is stable with only eight.
Worksheet 1.2: Atomic Theory

1. What part of Dalton’s atomic model explains the empirical rule known as the Law of Conservation of Mass?

2. Atomic theory has been changed many times over the past 200 years. What was the empirical evidence that suggested Dalton’s theory was not entirely correct?

3. In Thomson’s model of the atom, what electrical charge is represented by the chocolate chips and what electrical charge is represented by the bun?

4. Why did it take so long to come up with the Rutherford model of the atom?

5. Which atomic theory first suggested each of the following ideas?
   a) Atoms contain a dense positively charged nucleus.
   b) Atoms contain electrons.
   c) Electrons surround a central positive core.
   d) Atoms cannot be divided.

6. Using the development of atomic theory as an example, explain the statement “Scientific knowledge is subject to change.”

7. Bohr’s atomic theory is often described using the solar system model. What part of the atom does the sun represent in this model? What part of the atom do the planets represent?

8. How is Bohr’s theory of the atom the same as Rutherford’s theory? In what way is it different?

9. Electrons can be found in the ground state or the excited state. What is the difference in these two states?

10. Why do electrons emit light energy when they drop from a higher energy level to a lower energy level?

11. How could we use our knowledge of Bohr’s atomic theory to identify a sample of an unknown gas?