**NTI Day 4 Lesson**
**Content Standard: Atomic Structure and Theory**
**Class: Chemistry Teacher: K. Kelly**

# History of Atomic Theory

Modern atomic theory is rooted in the study of the structure of matter. Scientists for thousands of years have hypothesized about the makeup of matter. The time line below shows some important developments in scientific knowledge about matter and its building block, the atom.

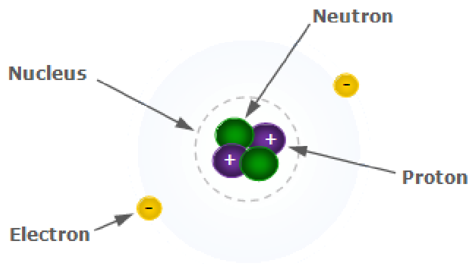
| | |
|---|---|
| c. 400 B.C. | The Greek philosopher Democritus proposed that matter is composed of relatively simple particles that are too small to be visible to the naked eye. These particles, which he called *atoma* (a Greek word meaning, "things that cannot be cut or divided"), could not be further broken down into smaller particles. |
| early 1800s | British scientist John Dalton proposes that atoms of different substances can be distinguished by their mass. |
| 1897 | British physicist Joseph Thomson performs an experiment that proves that atoms contain negatively charged particles-electrons. |
| 1901 | American researcher Robert Millikan experimented with drops of oil in order to calculate the mass of electrons. |
| 1911 | British chemist Ernest Rutherford proposed that an atom contained a nucleus, which was positively charged, dense, and very small. Rutherford is also credited with having discovered the proton in 1919. |
| 1932 | British physicist James Chadwick, building upon Rutherford's experiments, discovered that atomic nuclei contain neutrally charged particles-neutrons-in addition to protons. |

Today, the discoveries of Dalton, Thomson, Millikan, Rutherford, Chadwick, and many other scientists are combined to form the **quantum mechanics** atomic theory.

# Atomic Structure

Matter is made of minute particles called atoms.

An **atom** is the smallest part of an element that still retains properties of that element.




Atoms contain three main subatomic particles:

- protons
- neutrons
- electrons

Some general properties of these subatomic particles are described in the table below:

| subatomic particle | charge | location | actual mass (in grams) | relative mass (in atomic mass units) |
|---|---|---|---|---|
| *protons* | positive | in the nucleus | $1.673 \times 10^{-24}$ | 1 |
| *neutrons* | no charge (neutral) | in the nucleus | $1.675 \times 10^{-24}$ | 1 |
| *electrons* | negative | outside of the nucleus (in electron clouds) | $9.11 \times 10^{-28}$ | 1/1840 (0.0005) |

The number of protons in an atom's nucleus is equal to that atom's **atomic number**. This number, and hence the number of protons in an atom, never changes for a particular atom. If the number of protons does change (as it does during radioactive decay, for example), a totally new element results.

The **periodic table** arranges elements from left to right, and top to bottom, in order of their atomic number. Each element in the periodic table has one more proton in its nucleus than the element before it.

The **mass number** of an atom is the total number of protons and neutrons in that atom. The electrons are not included in the mass number, because their masses are so small compared to the masses of protons and neutrons. It would be like adding a penny to $1,840.

## *General Properties of Atoms*

Although scientists have been unable to get clear pictures of atoms, they have been able to get fuzzy pictures of them using electron microscopes. From these images and other resources, scientists have learned that atoms are composed of small, dense nuclei surrounded by large electron clouds.

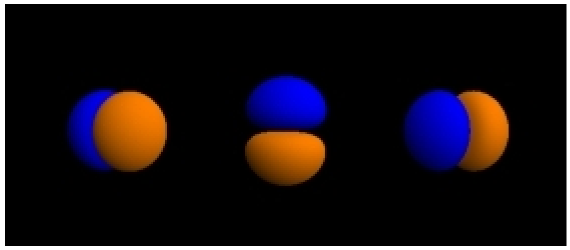
To get an idea of these relative sizes, imagine this model: if the nucleus of an atom were the size of a marble, then the whole atom including the electron clouds would be the size of a professional football stadium!

Although most of the volume or space of an atom is taken up by electron clouds, electron clouds have a very low density. This is because electron clouds have relatively little mass per unit volume since electrons only possess about 1/1840th the mass of protons or neutrons. The nucleus, however, is very dense because both the protons and neutrons are held together in a small area by strong forces.

## *Electron Clouds*

Previously, scientists believed that electrons moved around the nucleus in specific orbitals or circular paths. Now, however, scientists realize that it is impossible to know the exact location of an electron at any given time because electrons move very, very fast, and their motion is random. So, the locations of electrons are described by regions of probability known as electron clouds.

Three possible electron cloud shapes for carbon's two outermost electrons



*This image is courtesy of Wikimedia*

In the electron cloud model of an atom, electrons travel around the nucleus in unpredictable patterns, but it is possible to predict an area in which the electrons will be found. Think of it this way: imagine a balloon that contains a single bead. The bead represents an electron and the balloon represents an electron cloud. If you shake the balloon, you can definitely say that the bead is in the balloon somewhere, but you cannot pinpoint its exact location because by the time you point at a location, it will have moved somewhere else.

Electrons travel around the nucleus in specific energy levels. Each energy level contains a certain number of electron clouds, and electrons have a tendency to fill the clouds at the lowest energy level first. The first energy level can hold a maximum of 2 electrons, the second can hold a maximum of 8 electrons, the third can hold a maximum of 18 electrons, and the fourth can hold a maximum of 32 electrons. The number of energy levels that an element possesses is equal to the period (row) number of the element. The electrons located farthest from the nucleus (in the outermost energy level) are called valence electrons. Valence electrons determine an element's chemical properties, including reactivity.

## *Quantum Particles*

Scientists have discovered that neutrons and protons are made up of even smaller constituents called quarks. Murray Gell-Man discovered the first of these quantum particles, which he named quarks, in 1961. These minute particles can be studied further as scientific technology advances.

## *Isotopes*

Isotopes are atoms of the same element that have the same atomic number but different atomic masses. Since isotopes of an element have the same atomic number, they must have the same number of protons. Thus, the difference in atomic masses must be due to a change in the number of neutrons.

Carbon-14 is an example of a common isotope of carbon. Since the atomic number of carbon is 6, carbon-14 must have 6 protons. The number of neutrons can be calculated by subtracting the number of protons (6) from the mass number (14). Thus, carbon-14 has 8 neutrons.

Carbon-14 can be also be depicted as $^{14}C$ or C-14.

## *Ions*

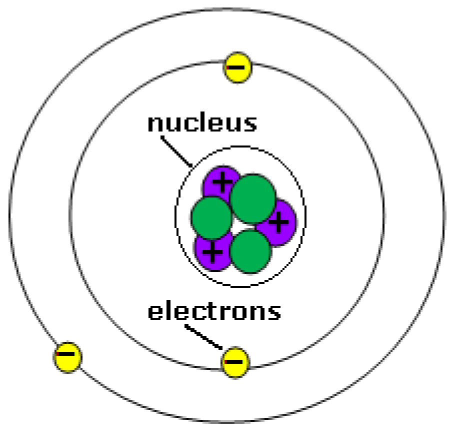
Neutral atoms have no charge. So, in order to be neutral, an atom must possess the same number of positively-charged protons and negatively-charged electrons. Since neutrons have no charge, the number of neutrons has no effect on the overall charge of the atom.

Ions are charged particles. Ions, therefore, must always have an unequal number of protons and electrons. If there are more electrons than protons, the ion has an overall negative charge. If there are more protons than electrons, the ion has an overall positive charge.

For example, iodine forms a negative ion with an overall charge of -1 ($I^{-1}$). This means that iodine has 1 more electron than it does protons.

# Forces in Atoms

The forces in an atom help to hold the atom and its nucleus together.




The nucleus of an atom contains positively charged protons and neutrons that do not have a charge. Outside the nucleus, there are negatively charged electrons. The protons in the nucleus attract the electrons outside of it, causing the electrons to orbit around the nucleus. **Electric forces between the nucleus and electrons** hold the atom together.

There is **a strong force between the protons and neutrons** in an atom's nucleus that is usually stronger than the electric force that would pull the particles apart. This strong force is only evident at nuclear distances and holds the nucleus together against the electrical repulsion that exists between the protons.