Reflect

Imagine a piece of aluminum foil about the size of a sheet of paper. If you cut this piece of foil in half, have you changed the identity of the matter making up the foil? What happens if you cut the foil in half again? And if you cut this sample in half, what do you have? You know that cutting aluminum foil in half does not change its chemical identity. You still have aluminum foil every time you cut a piece in half. But suppose you could keep cutting the foil in half indefinitely. Would you ever reach a point where you could not cut any further without destroying the aluminum's identity?

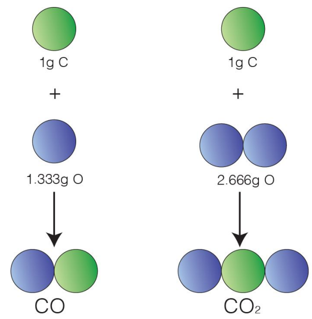
The Greeks hypothesized that matter is not infinitely divisible. Can matter be divided into infinitely smaller portions? Long before the microscope was invented, the ancient Greeks debated this question. In 435 BCE, the Greek philosopher Leucippus reasoned that matter must be composed of finite particles. If something is finite, it has a limit; in other words, it cannot be divided into smaller and smaller pieces forever. His student, Democritus, developed this idea. Democritus used the idea of tiny, indivisible particles called atoms to explain the properties and behavior of matter. (In Greek, *atom* literally means "uncuttable.") Democritus hypothesized that these particles are always moving, and therefore their arrangement in space must be constantly changing. According to Democritus, the movements and changing arrangements of atoms determine all of the phenomena that we observe in the natural world. Neither of these men had any experimental evidence supporting their ideas. They based their ideas on philosophical thought and reasoning.

Democritus hypothesized that all matter and its behavior could be explained by the nature of tiny indivisible particles that comprise it.

John Dalton presented the first experimental evidence to support the concept of atoms. Many centuries passed before anyone was able to find experimental evidence for the existence of atoms. John Dalton, an English schoolteacher, carried out experiments that provided indirect evidence of atoms. He showed that oxygen combined with another gas, nitric oxide, in specific volume and weight ratios. In other words, no matter how much oxygen and nitric oxide he combined, the two always reacted with one another in a constant weight ratio. Dalton concluded that this could occur only if the gases were made up of atoms reacting with one another in whole number combinations. He proposed his atomic theory in 1803 as a series of statements called postulates:

- **All matter is composed of indivisible particles that cannot be created or destroyed.** Like Leucippus and Democritus, Dalton hypothesized that the smallest unit of matter is an indivisible particle. He reasoned that these particles—which he also called atoms—are present in a fixed number in the universe. Atoms cannot be created or destroyed.
- **Atoms of the same element are identical, but atoms of different elements have different properties.** Scientists were well aware of different elements—including iron, mercury, gold, and silver—at the time Dalton developed his atomic theory. Dalton used his concept of atoms to explain the differences between elements. He proposed that each element is composed of the same type of atoms, but that the atoms in one element differ from the atoms in another element. A chief difference is atomic weight. Different types of atoms have different weights.

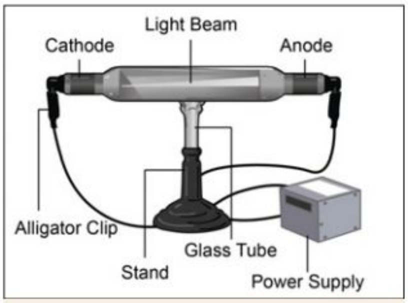
- **Atoms group together in whole number ratios to form compounds.** Dalton's experiments confirmed what other scientists had discovered about chemical compounds—they are made up of two or more different elements. He also confirmed that in any compound, the same weight ratio of the elements is always present. For example, carbon monoxide always has one gram of carbon for every 1.333 g of oxygen. However, carbon dioxide always has one gram of carbon for every 2.666 g of oxygen. Dalton used this information as evidence for his atomic theory. According to Dalton, the fixed weight ratios of elements in a compound result from the fixed ratios of atoms of those elements.
- **Chemical reactions are the result of the rearrangement of atoms**. By the time Dalton began his work, several scientists had already published results of their experiments involving chemical reactions. These results indicated that two substances combined in constant proportions. Dalton used his concept of atoms to explain that the constant proportions indicate a specific whole number of atoms of one reactant rearranging with another whole number of atoms of a second reactant.



Dalton's theory explains the observation that carbon and oxygen can combine in two ways.

Although Dalton's first postulate overlapped with the ideas put forth by the Greeks 2,000 years earlier, his ideas carried more weight because they were backed by experimental evidence. Dalton was cautious in that regard, relying on investigational data rather than philosophical thinking as he developed his atomic theory.

J. J. Thomson discovered the first subatomic particle: the electron. Almost 100 years after Dalton proposed his atomic theory, English scientist J. J. Thomson made another discovery about the nature of matter. By this time, scientists had accepted Dalton's postulates about the atom as the fundamental particle. However, nothing was known about the makeup of the atom itself.

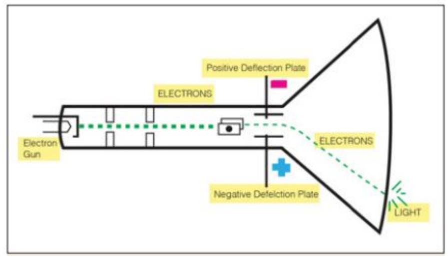


A cathode ray tube uses high voltage to produce a light beam inside a glass tube.

In the late 1800s, J. J. Thomson was using a device called a cathode ray tube to explore some mysterious rays that caused a fluorescent glow. A cathode ray tube is a glass tube with two metal wires at each end. When Thomson used a power supply to apply a high voltage across the two wires, he could observe a greenish glowing light inside the tube. Scientists at that time were not sure about the nature of this light. They didn't know if it was some kind of light wave or a stream of particles. A number of scientists ran experiments to try to uncover the nature of the mysterious rays.

J. J. Thomson succeeded in devising a series of experiments that solved the mystery. Thomson built several new cathode ray tubes, one of which is shown in the diagram below. He designed this tube to test whether the cathode rays could be bent by an external electric field. He added two plates to the interior of the tube: one with a positive charge and the other with a negative charge. These plates were positioned above and below the path of the cathode rays. When he tested his newly built tube, Thomson found that the cathode rays bent away from the negatively charged plate. Thomson concluded the rays were composed of particles that carried a negative charge. More experiments with other tubes enabled Thomson to estimate the size of these particles. His work established that they were much smaller than atoms.

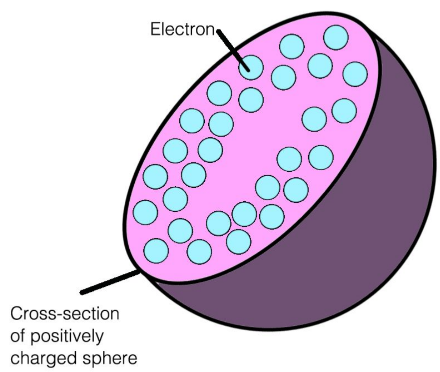
Thompson did not use the term electron to describe the negatively charged particles he discovered. However, Thompson did hypothesize that these particles represented parts of an atom. Thompson proposed that the structure of an atom could be likened to an English dessert known as plum pudding. He envisioned a model in which the negatively charged particles are embedded in a sphere of positive charge much like raisins are embedded in a plum pudding.



J. J. Thomson's cathode ray tube experiments demonstrated that cathode rays are composed of negatively charged particles much smaller than atoms. Later scientists would call these negative particles "electrons."

A hunch led Ernest Rutherford to discover that atoms contain a nucleus. Thomson's plum pudding model of the atom was later discarded in favor of a model in which electrons occupy the space around a positively charged nucleus. One of Thompson's former students, Ernest Rutherford, came up with the nuclear model. Rutherford had worked in Thompson's lab in England in the 1890s before becoming a professor at another English university.

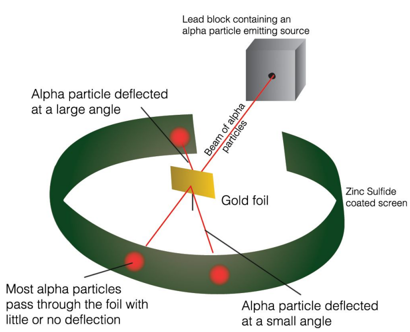
There, he worked with his own research students to study the nature of a type of **radioactivity** known as alpha radiation. Alpha radiation results from the emission of large particles, called alpha particles, from unstable atoms.



J. J. Thomson developed the "plum pudding model," in which negatively charged particles are embedded in a positively charged sphere.

In 1908, Rutherford directed his students to conduct some experiments involving the emission of alpha radiation at squares of thin, gold foil. He wondered what kind of backscatter his students might find when the alpha radiation hit the gold atoms. His students carried out the experiments and reported that most of the alpha particles passed through the foil while a small proportion was deflected backward.

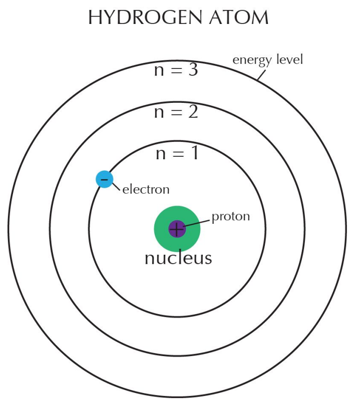
From these results, Rutherford concluded that atoms contain extremely small, dense, and positively charged nuclei. If an alpha particle collided with an atom's nucleus, the particle bounced off the gold foil. Because most particles passed through the foil, however, Rutherford concluded that the area around the nucleus is mostly empty space with a few negative electrons. He published these conclusions in 1911, refuting the plum pudding model.



Ernest Rutherford's gold foil experiment demonstrated that the atom could not be described by the plum pudding model.

What Do You Think?

What results would Rutherford and his students have observed if the plum pudding model had been accurate? Note that in this model, the matter making up the atom was thought to be spread more evenly within the atom and not concentrated into a small space at the center.



Bohr's model of the hydrogen atom placed electrons in specific orbits around the nucleus.

Niels Bohr refined the atomic model to explain how electrons were positioned around the nucleus. Rutherford's work shed light on the presence of an atomic nucleus, but it didn't reveal any new information about the positions of electrons in an atom. Niels Bohr, a student in Rutherford's lab, worked on extending Rutherford's model to explain how electrons were arranged in the atom.

At that time, scientists knew that different elements emitted different patterns of light when they were heated to high temperatures. They also knew that the pattern of light emitted by an element could be described by a mathematical relationship. Bohr took this idea one step further and hypothesized that the mathematical relationship indicated the presence of different energy levels for electrons within the atom. From this, Bohr developed a model of an atom with a central nucleus surrounded by electrons moving in orbits. Moving from the nucleus outward, each orbit represented higher energy than the previous one.

In Bohr's model, electrons could move from one orbit to another. If an electron moved from a lower energy level to a higher one, it had to absorb energy to make the move. If an electron moved from a higher energy level to a lower one, it lost energy in the form of light. This explained the light emission spectrum of a hydrogen atom.

However, scientists gradually realized that hydrogen is the only element whose light emission spectrum fits Bohr's mathematical calculations. Despite this serious flaw, the Bohr model of the atom continued to be important to scientists. The idea of specific energy levels for electrons became an important one that endured even after the model itself was found to be inadequate.

All of these scientists were important in the development of the atomic theory, the design of our current atomic model, along with the role of the different subatomic particles discovered through their many experiments.

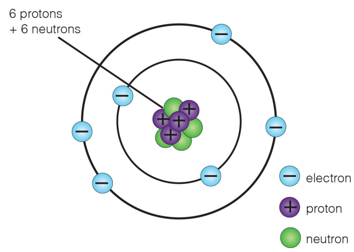
Atomic Structure

All atoms have the same general arrangement. An atom contains a *nucleus*, which is in the center of the atom. An atom also has an area of space surrounding the nucleus called an *electron cloud*. Subatomic particles, the small particles within the atom, are known as protons, neutrons, and electrons. They differ in mass, charge, and location in the atom.

- **Neutron (*n*^{o})**: Does not have an electrical charge. They are neutral. Neutrons are found in the nucleus of the atom and are represented by the red spheres in the atomic diagram..

Models and Properties of Atoms

- **Proton (p^+)**: Protons are positively charged particles ($^+1$). They are found in the nucleus of an atom and are represented by the blue spheres in the atomic diagram to the right. Since protons are the only charged particle in the nucleus, an atom's nucleus is always positively charged. The number of protons in the nucleus determines the identity of the atom. This means that each element has a unique number of protons (atomic number *Z)* in the nucleus of each atom. For example, all carbon atoms contain six protons, all gold atoms contain 79 protons, and all lead atoms contain 82 protons.
- **Electron (e^-)**: Electrons are negatively charged particles (-1). They travel around the nucleus in orbitals, within the electron cloud surrounding the nucleus. Electrons are constantly moving. The green spheres in the atomic diagram represent electrons.



Carbon Atom

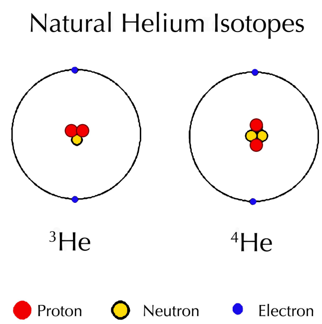
Classic model of a carbon atom with atomic number of 6 (six protons). In the nucleus are six protons and six neutrons.

Different atoms have different combinations of subatomic particles. However, there are some general rules regarding the electrical charge of an atom. When the number of protons equals the number of electrons, the atom is neutral because the positive nucleus balances with the negatively charged electrons. If there are more protons than electrons, the atom becomes a positively charged ion. If there are more electrons than protons, the atom becomes a negatively charged ion.

Why do you think the number of neutrons in an atom does not affect the overall charge of the atom? Hint: Think about the characteristics of protons and electrons.

Atomic structure affects the properties of elements.
Chemical reactions involve either the transfer or the sharing of electrons between atoms. Therefore, the chemical reactivity properties of an element primarily depend on the number of electrons in an atom of that element. Protons are significant because the tendency for an atom to either lose, gain, or share electrons depends on the total charge of the protons in the nucleus. The chemical reactivity of an atom depends on the number of electrons and protons, and is independent of the number of neutrons. The mass and radioactive properties of an atom depend on the number of protons and neutrons in the nucleus.

The Masses of Subatomic Particles
Protons and neutrons have similar masses. Each proton and neutron has a mass of approximately 1.67×10^{-27} kilograms. The mass of one electron is even smaller: 9.11×10^{-31} kilograms. Protons and neutrons located in the atom's nucleus represent where most of the atom's mass is located. *Atomic mass* is the sum of the mass of all of the protons and neutrons in the atom. Electrons are so small that they not do change the mass of the atom significantly. They do, however, take up a large amount of the volume of the atom. The following formula can be used to find the atomic mass where P = number of protons and N = number of neutrons: **P + N = atomic mass.**



Atoms of the same element can have different numbers of neutrons. Elements in the periodic table are ordered by their atomic number: the number of protons in the atom's nucleus. The number of protons defines the type of element, yet atoms of the same element may have different numbers of electrons and neutrons. Because both atoms in the diagrams above have seven protons, they are atoms of the same element: nitrogen (N). They are also isotopes. What are isotopes? Isotopes are atoms of the same element that have different numbers of neutrons. Let's consider another example. There are two natural isotopes of helium (He). One isotope has one neutron, and the other isotope has two neutrons. Both isotopes have two protons. Therefore, the mass number—the total number of an atom's protons and neutrons—of the first helium isotope is three (two protons plus one neutron). The mass number of the second helium isotope is four (two protons plus two neutrons).

Helium has only two isotopes, but most elements have more. Iron (Fe) has more than 25 isotopes, but many of these isotopes are not very stable. The most abundant iron isotope has 26 protons and 30 neutrons. Therefore, this isotope has a mass number of 56.

Look Out!

There are several ways to write an isotope of a particular element. One way is to write the mass number as a superscript before the elemental symbol. Following this method, the two isotopes of helium are written ^{3}He and ^{4}He. Another way is to use a hyphen to separate the name of the element and its mass number: for example, helium-3 and helium-4.

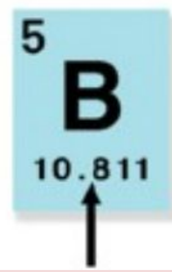
The average atomic mass of an element depends on its isotopes.
On the periodic table, each element is listed with its atomic number above the chemical symbol. Below the atomic symbol is a decimal number that is the element's average atomic mass. An element's average atomic mass is the weighted average of the mass numbers of all the naturally occurring atoms and isotopes of that element. Most elements have more than one naturally occurring isotope. Do not confuse an isotope's mass number with the average atomic mass of the element.

For example, the element boron (B) has two naturally occurring isotopes: boron-10 and boron-11. To calculate a weighted average, you need to know the abundance of each isotope: How frequently does it occur in nature? The abundance of boron-10 is 19.9 percent, and the abundance of boron-11 is 80.1percent. (In other words, approximately four out of every five boron isotopes in nature have five protons and six neutrons. The remaining boron isotopes in nature have only five neutrons.) Average atomic mass is calculated by multiplying the mass number of each isotope by the abundance of that isotope, and then taking the sum.

For example, here is how to calculate the average atomic mass of boron:

$$\begin{aligned}\text{Average atomic mass}_B &= (10)(19.9/100) + (11)(80.1/100)\\ &= 1.99 + 8.811\\ &= 10.801\end{aligned}$$

We just calculated boron's average atomic mass to be 10.801. However, the periodic table lists boron's average atomic mass as 10.811. What happened? In fact, an isotope's mass number does not exactly equal the sum of its protons and neutrons. The mass number of boron-10 is actually closer to 10.013 atomic mass unit, and the mass number of boron-11 is actually closer to 11.009 atomic mass unit. (Atomic mass is measured in atomic mass units: amu.) If you use these numbers in your calculation, you will get an average atomic mass closer to 10.811 atomic mass units.



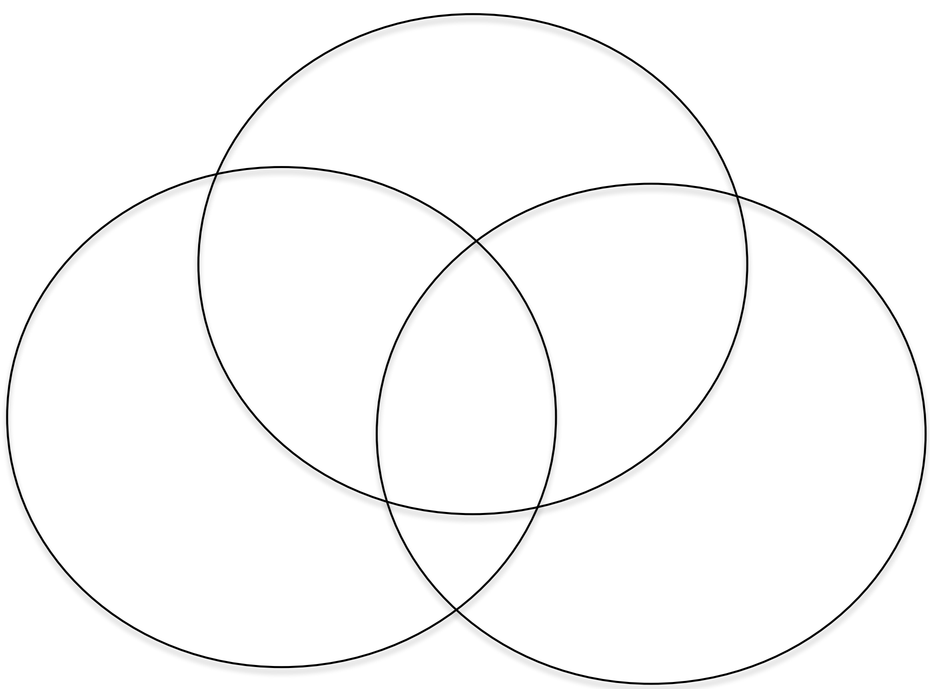
The atomic number of boron (B) is 5. The average atomic mass of boron is 10.811.

Try Now

Atoms are made of protons, neutrons, and electrons. Read the characteristics of these subatomic particles in the box below. Decide whether each characteristic describes protons, neutrons, or electrons. Then write each characteristic in the correct section of the Venn diagram.

Characteristics of Subatomic Particles

- Negatively charged
- Positively charged
- Have no charge
- Smallest mass
- In the nucleus
- In orbitals around the nucleus
- Subatomic particles
- Discovered by J. J. Thomson
- Make up the mass of the atom
- When equal, atom is neutral



Connecting With Your Child

Determining Averages at Home

To help your child practice the calculations necessary to determine average atomic mass, calculate the average masses of everyday objects such as eggs in an egg carton. You may use a small kitchen scale to make the measurements; all eggs should be roughly the same size (small, medium, large, or extra large). Use two cartons, each with different-sized eggs (small, medium, large, or extra large). One carton should contain a dozen eggs, all the same size; the other carton should contain a half dozen eggs, all the same size.

Have your child create a two-column table with these headings: “Carton 1” and “Carton 2.” Include 12 rows, one for each egg in carton 1; you will use only the first six rows for carton 2. Students should record their measurements in their tables as they measure each egg. After your child measures all the eggs in the first carton, add their masses and divide by 12 to determine the average mass of these eggs. Have your child repeat the experiment for the second carton, dividing by six rather than 12 to determine the average mass of those eggs.

Finally, determine the weighted average mass of all the eggs, using what you know of the abundance of each size of egg. For example, if the average mass of the large eggs in carton 1 is 55 grams, then 67 percent of your eggs (12 of 18) have an average mass of 55 grams. If the average mass of the small eggs in carton 2 is 45 grams, then 33 percent of your eggs (6 of 18) have an average mass of 45 grams. You can calculate weighted average mass as follows:

$$\begin{aligned}\text{Average mass}_{\text{eggs}} &= (55\text{ g})(67/100) + (45\text{ g})(33/100)\\ &= 36.85\text{ g} + 14.85\text{ g}\\ &= 51.7\text{ g}\end{aligned}$$

The weighted average mass of the eggs is 51.7 grams.

Here are some questions to discuss with your child:

1. Why is average atomic mass an important value for each isotope? What can average atomic mass tell you about an element?
2. How does the average egg mass relate to the average atomic mass? How does the individual mass of each egg relate to an atom’s mass number?
3. How do the different cartons relate to isotopes?