

B3 Describe ideas used in interpreting the chemical nature of matter, both in the past and present, and identify example evidence that has contributed to the development of these ideas

B3.2 distinguish between observation and theory, and provide examples of how models and theoretical ideas are used in explaining observations (e.g., describe how observations of electrical properties of materials led to ideas about electrons and protons; describe how observed differences in the densities of materials are explained, in part, using ideas about the mass of individual atoms)

Scientific Theory

A scientific theory is a set of ideas that help to explain certain facts. A theory is not a guess; it is an explanation that has been tested and supported by the results of experiments. A theory is accepted to be correct according to the information that is available at the time.

However, theories can be disproved as new information becomes available or as technology improves. As more scientific knowledge is gained, existing theories are challenged or re-evaluated. When an accepted theory is disproved and no longer explains certain facts, scientists develop a new theory to explain the available information.

A good Example of how new discoveries and technology change theories is the development of the atomic theory. The atomic theory originated with the ancient Greeks and continues to develop today.

General Timeline of Atomic Theory

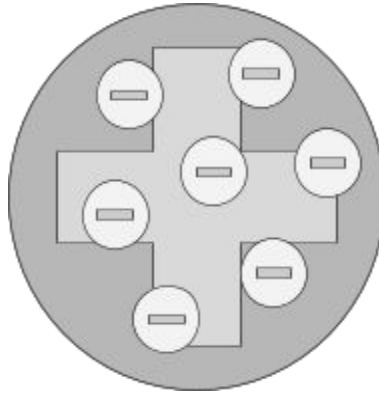
Around 400 BCE, the Greek philosopher Democritus used the word *atomos* to describe the smallest particles of matter that could not be broken down any further. The Greek word *atomos* means indivisible.

In the 1660s, Robert Boyle studied the properties of gases. He believed that all matter was made up of tiny particles and that these particles combined to make different substances.

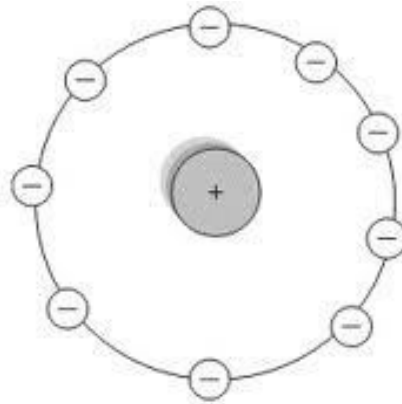
John Dalton suggested that all matter is made up of elements in 1808. He was the first to suggest that these elements are pure substances made of particles called **atoms**. Dalton believed that atoms were tiny, solid spheres of matter.

J. Plücker built a cathode ray tube in 1859. The cathode ray tube would be used by others to study gas molecules in more detail.

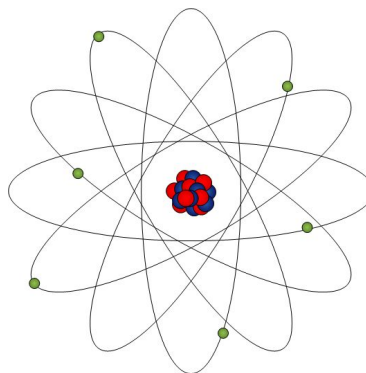
In 1897, J.J. Thompson used the cathode ray tube to show that the cathode rays were made of negatively charged particles. He called these particles electrons, and he showed that these particles were much smaller than atoms. He describe the Raisin-bun model of the atom. This model had a large positive particle in the centre, with small, negative electrons stuck around the outside of it.



Hantaro Hugaoka introduced the Saturn model of the atom in 1904. This model had a central positive charge and the electrons orbiting around the centre like planets around the sun.

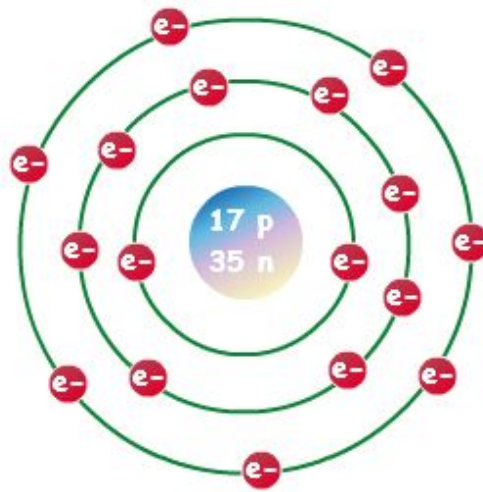


In 1911, Ernest Rutherford projected positively charged particles through thin gold foil. He suggested that the nucleus was both very small and positively charged. Rutherford also believed that the electrons were located outside the nucleus.

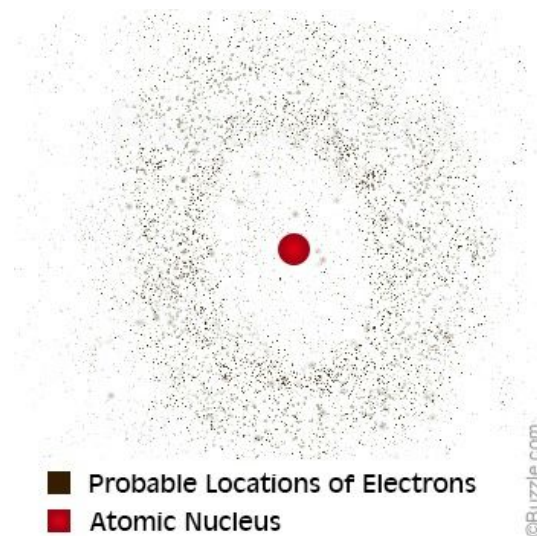


In 1914, Henry Gwyn Jeffreys Moseley used X-ray tubes to determine that the atomic number of an element was the same as the number of protons in the nucleus. As a result, the periodic table was reorganized according to the atomic numbers of elements instead of their atomic masses.

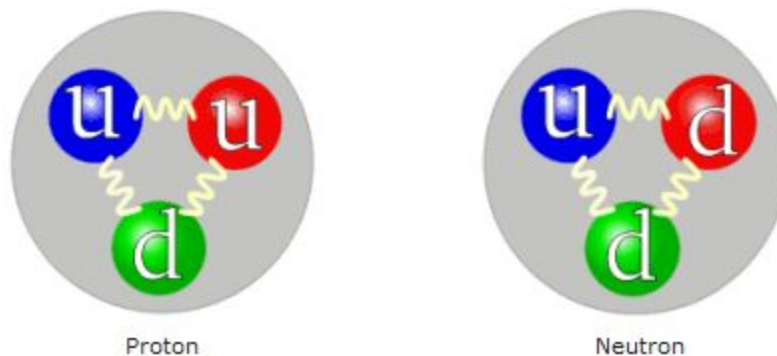
In 1922, Niels Bohr suggested that electrons do not orbit the nucleus randomly. He showed that they orbit the nucleus in specific shells around it.



In 1930, Erwin Schrödinger described the movement of electrons as forming continuous clouds around the nucleus.



In 1968, it was shown that the nucleus contains particles smaller than either protons or neutrons. These particles were called quarks.



From 1978 to 1995, several experiments confirmed the existence of quarks. Researchers also tried to measure some of the properties of quarks.

From 1995 to today, nuclear physicists have continued to explore the properties of the nuclei of different elements.

As you can see from this timeline of the development of atomic theory, scientific knowledge is always changing. As more information is found and technology continues to develop, scientists continue to re-evaluate accepted theories in order to explain the new information.

B3.1 demonstrate understanding of the origins of the periodic table, and relate patterns in the physical and chemical properties of elements to their positions in the periodic table—focusing on the first 18 elements

The Periodic Table

History and Trends

Matter has been studied for many centuries. In the late 1800s, Dmitri Mendeleev organized the elements into the periodic table. The periodic table is separated by a staircase, with metals on the left of the staircase and non-metals on the right. Metals have some similarities in their chemical properties because they all lose electrons to form positive ions. Typically, metals are shiny, flexible, and good conductors of heat and electricity. Almost all metals are solids at room temperature and normal pressure. Non-metals are to the right of the staircase. The chemical properties of non-metals are related to the fact that they accept electrons to form negative ions. Physically, non-metals are typically poor conductors of heat and electricity. They can be solids, liquids, or gases at normal temperature and pressure.

Each horizontal row of the periodic table is called a **period**. The period of an element is determined by the number of energy levels that are occupied by electrons. Each vertical column of the periodic table is a **group**, or family. All the members of a chemical properties. That is, they react with the same elements in a similar way. Group 1, the alkali metals, all have one valence electron and are all highly reactive metals. Group 2, the alkaline-earth metals, can all combine by ionic bonding with two atoms of members of Group 17 (BeCl_2 , MgCl_2 , CaCl_2 , etc.). Group 17, the most reactive non-metals, are called halogens. Group 18, the noble gases, do not have any unpaired electrons in their valence energy levels. They are inert and do not react to form compounds under normal circumstances.

B3.3 use the periodic table to identify the number of protons, electrons and other information about each atom; and describe, in general terms, the relationship between the structure of atoms in each group and the properties of elements in that group (e.g., use the periodic table to determine that sodium has 11 electrons and protons and, on average, about 12 neutrons; infer that different rows (periods) on the table reflect differences in atomic structure; interpret information on ion charges provided in some periodic tables) [Note: Knowledge of specific orbital structures for elements and groups of elements is not required at this grade level.]

Atomic Structure and the Properties of Elements

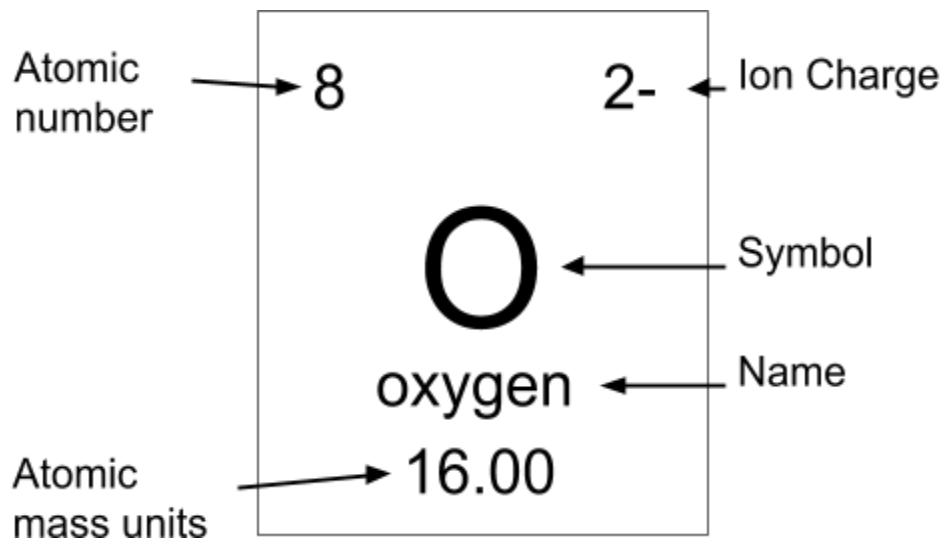
Atoms are composed of three basic subatomic particles: protons, electrons, and neutrons. Each of these particles is identified by specific characteristics.

Particle	Proton	Electron	Neutron
Charge	+ Positive	- Negative	No Charge
Size	Large - same as a neutron	Small - about $\frac{1}{2000}$ the size of the proton	Large - same as a proton
Location	Nucleus	Cloud around nucleus	Nucleus

The **atomic number** of an atom is the number of protons found in the nucleus. This also represents the number of electrons found in the orbits. The elements in the periodic table are arranged in order of increasing atomic number. For example, the first element is hydrogen, which has one proton and an atomic number of 1. The second element is helium, which has two protons and an atomic number of 2, and so on.

The **atomic mass unit** (amu) of an atom is the total mass of the protons and neutrons in the nucleus.

This information can be found for each element in the periodic table.



Oxygen has an atomic number of 8. This means that there are eight protons and eight electrons in an oxygen atom. The mass of oxygen is 16.00 amu. You can calculate the number of neutrons by subtracting the atomic number from the mass number.

$$\begin{array}{ccccccc} \text{Mass} & & & & \text{Atomic} & & & & \text{Number} \\ \text{number} & & + & & \text{number} & & = & & \text{of neutrons} \end{array}$$

For oxygen, the number of neutrons can be calculated as follows:

$$16.00 - 8 = 8 \text{ neutrons}$$

B3.4 distinguish between ionic and molecular compounds, and describe the properties of some common examples of each

Ionic and Molecular Compounds

Atoms of different elements can join together to form compounds. Recall that the periodic table can be divided into metals on the left side of the staircase and non-metals on the right side of the staircase. When a metal and a non-metal are joined together, they are joined by an ionic bond, and the resulting compound is called an **ionic compound**. An example of an ionic compound is sodium chloride, $\text{NaCl}_{(s)}$, also known as table salt. Sodium atoms (metal) are joined by ionic bonds to chlorine atoms (non-metal). All ionic compounds are solids at room temperature and have relatively high melting points. They have distinct crystal structures and can conduct electricity if they have been melted or dissolved in water.

Bonds can also form between two different non-metals. This is called covalent bonding, and the compounds that are formed are called **molecular compounds**.

One of the most common examples of covalently bonded molecular compound is water, $\text{H}_2\text{O}_{(l)}$. Table sugar, $\text{C}_{12}\text{H}_{22}\text{O}_{11(s)}$, is another example of a molecular compound. Several elements are able to form covalent molecules using only their own atoms. Hydrogen, fluorine, and oxygen atoms are found in nature as pairs of atoms. For this reason they are called diatomic elements.

Molecular compounds can be solids, liquids, or gases at room temperature. They generally do not conduct electricity, and they have relatively low melting and boiling points because the bonds holding them together are much weaker than those in ionic bonds.