Unit 1: Atomic Structure and Periodicity

Parent Guide

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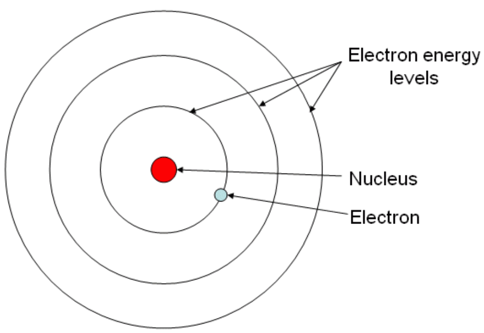
SC1. Obtain, evaluate, and communicate information about the use of the modern atomic theory and periodic law to explain the characteristics of atoms and elements.

SC1. a. Evaluate the merits and limitations of different models of the atom in relation to relative size, charge, and position of protons, neutrons, and electrons in the atom.
From ancient Greek philosophers to 18^{th} century Europeans, humans in history longed to describe and understand the composition of their surroundings. Democritus used the term “atomos” to name the smallest pieces of the matter around him. His contemporary, Aristotle, disagreed. The atomic theory was stagnant for nearly 2,000 years. In 1803, an Englishman, John Dalton rekindled the atomic theory.

An English physicist, J.J. Thomson, discovered the electron in the late 1800s and proposed his “Plum Pudding” model of the atom at this time. Ernest Rutherford began experiments to prove the Plum Pudding Model correct. In 1909, Rutherford’s Gold Foil experiment led to the discovery of the atom’s nucleus.



A Danish physicist, Neils Bohr explained that electrons travel around the nucleus in circular paths at specific, fixed distances from the nucleus. He called these paths energy levels. In 1924, Louis de Broglie proposed that electrons do not behave simply as particles. Three years later, de Broglie proved his theory. Eleven years later, an Austrian physicist, Erwin Schrodinger, described the abstractness of the modern atomic theory with a thought experiment about a cat. He tried to express that the uncertainties and complexities of atomic structure and electron motion cannot be described in the macroscopic world.

SC1. b. Construct an argument to support the claim that the proton (and not the neutron or electron) defines the element’s identity.
The proton, found in the nucleus, has a positive charge. The number of protons is unique to each element and symbolized by the atomic number on the periodic table.

The neutron is also found in the nucleus. It has no charge. The neutron is credited with preventing the like charged protons from repelling one another. For practical purposes, protons and neutrons are considered to have the same mass. Two atoms of the same element can have different numbers of neutrons. Neutron differences cause differences in mass, but they do not affect the identity of the atom.

The electrons compose the electron cloud surrounding the nucleus. The electron has a negative charge, and the mass is so small that it is considered negligible.Two atoms of the same element can have different numbers of electrons. The difference in electrons causes differences in charge, but the difference in electrons does not affect the atom’s identity.

SC1. c. Construct an explanation, based on scientific evidence, of the production of elements heavier than hydrogen by nuclear fusion.
During the formation of the universe, only the lightest elements were formed. Other elements found in nature were created in nuclear reactions in stars and in supernovae. Stars fuse hydrogen into helium in their cores. Two atoms of hydrogen are combined in a series of steps to create helium-4. These

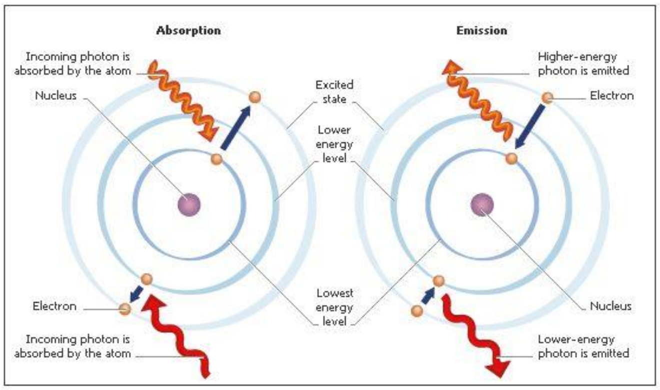
reactions account for 85% of the Sun's energy. The remaining 15% comes from reactions that produce the elements beryllium and lithium.
Examples of element making (nucleogenesis) in helium burning reactions:

- 3 helium atoms fusing to give a carbon atom: 3 @ $^{4}He \rightarrow {}^{12}C$
- carbon atom + helium atom fusing to give an oxygen atom: $^{12}C + {}^{4}He \rightarrow {}^{16}O$
- oxygen atom + helium atom fusing to give a neon atom: $^{16}O + {}^{4}He \rightarrow {}^{20}Ne$
- neon atom + helium atom fusing to give a magnesium atom: $^{20}Ne + {}^{4}He \rightarrow {}^{24}Mg$

Resource: https://www.sciencelearn.org.nz/resources/1727-how-elements-are-formed

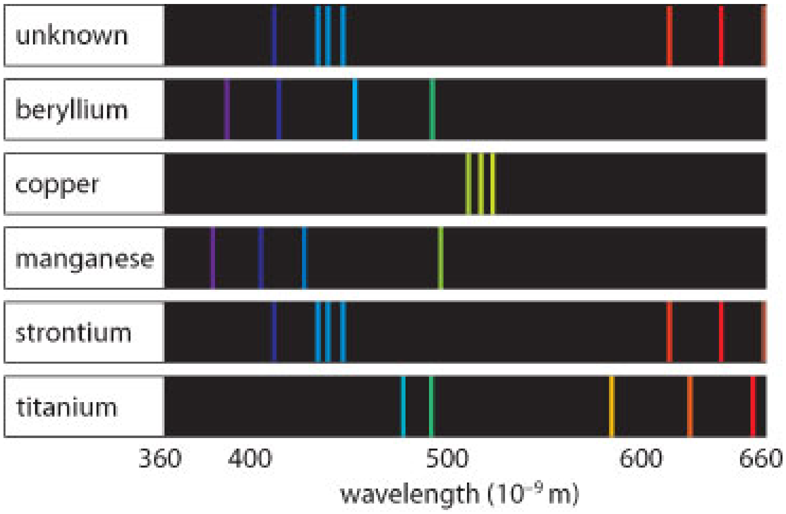
SC1. d. Construct an explanation that relates to the relative abundance of isotopes of a particular element to the atomic mass of the element.
Isotopes are atoms of the same element, therefore containing the same number of protons, with different numbers of neutrons. A difference in neutrons causes a variation in atomic weight, not atomic identity. When scientists catalog properties of the elements on the blocks of the periodic table, they display an atomic mass. Atomic mass is not the actual mass of any particular atom of the element. Instead, atomic mass is the weighted average of all of the isotopes of the element.

SC1. e. Construct an explanation of light emission and the movement of electrons to identify elements.



Each electron exists at a particular energy level called ground state. The electron's ground state is the lowest allowable energy for that electron. When an atom is exposed to an energy source, the electrons absorb the incoming energy. The left side of Figure 5 illustrates the absorption of a photon, a particle of light energy. The photon causes the electron to jump from its ground state to a temporary, excited state at a higher energy level.

On the right side of Figure 5, electrons cannot exist permanently at the higher energy state. The recently absorbed energy will be released by the electron as it returns to its ground state. The released photon of energy will leave the atom. Photons, particles of light energy, travel in energetic waves. If the photon's wave has a wavelength between 400 nm and 750 nm, our eyes will detect visible light.



When energy is absorbed by a particular element's atoms, the electrons eventually release many waves of light energy. The particular set of wavelengths emitted by an element is called the atomic emission spectrum. Much like the fingerprint of a human, the atomic emission spectrum is unique to each element. Therefore, an element can be identified by the light emitted from its electrons.

SC1. f. Use the periodic table as a model to predict the relative properties of elements based on the patterns of electrons in the outermost energy level of atoms (including atomic radii, ionization energy, and electronegativity of various elements.)

The periodic table is arranged in order of increasing atomic number. The periodic law states: "when elements are arranged in order of increasing atomic number, many properties will repeat every eight elements."

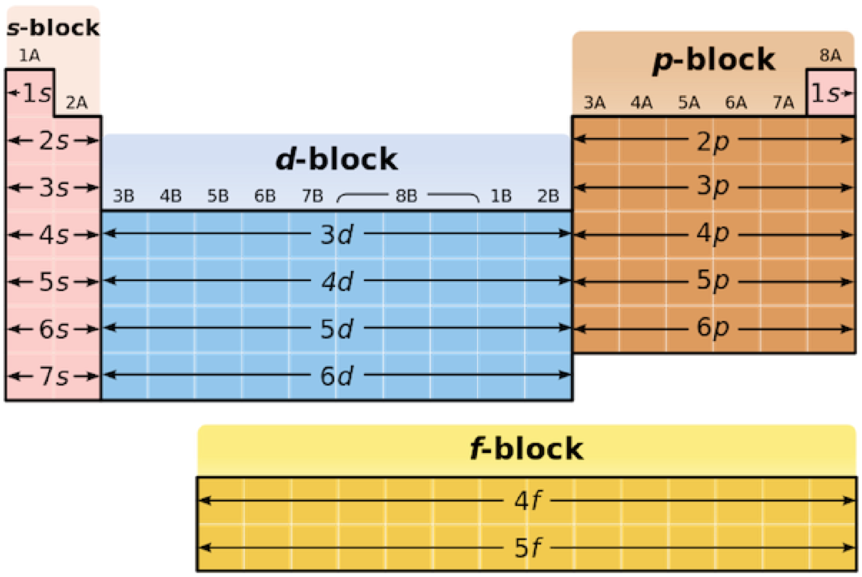
Patterns of Properties

The periodic table is organized into horizontal rows, called periods, and vertical columns, called groups.

- Number of Energy Levels: The elements in each horizontal row have the same number of electron energy levels in their atoms.
- Number of Valence Electrons: The elements in the same vertical column have the same number of valence electrons. Therefore, elements in the same vertical column tend to form similar ions and chemical bonds.
- Strength of Nuclear Charge: The strength of the nucleus' positive charge increases from left to right within a row.
- Atomic Radius: Atomic radius is defined as the one-half the distance between the nuclei of two identical bonded atoms. The relative size of an element's atoms can be predicted by finding the element on the periodic table and applying a few simple ideas about electrostatic attraction, nuclear charge and electron energy levels. Atomic radius decreases from left to right across a period, and the radius decreases from the bottom to top of a group.
- Ionization Energy: Ionization energy is defined as the energy required to remove the most loosely held electron from an atom. The most loosely held electron is typically on the valence energy level. The energy required to remove an electron from an atom increase from left to right across a period and from bottom to top of a group.
- Electronegativity: Electronegativity is defined as the ability of an atom in a chemical bond to draw the bonding electrons closer to itself. Electronegativity values increase from left to right across a period. The values also increase from bottom to top of the group.

SC1. g. Develop and use models including electron configuration of atoms and ions to predict an element's chemical properties.

Energy levels contain sublevels, and sublevels contain orbitals, and orbitals contain electrons spinning on their own axes. Each type of sublevel corresponds to a particular area of the periodic table.



Electron configuration: Electrons are arranged in sublevels in energy levels in the electron cloud. The arrangement follows a predictable pattern.

Examples:

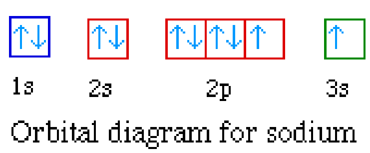
C	$1s^22s^22p^2$
Ne	$1s^22s^22p^6$
Al	$1s^22s^22p^63s^23p^1$
Fe	$1s^22s^22p^63s^23p^64s^23d^6$
Se	$1s^22s^22p^63s^23p^64s^23d^{10}4p^4$

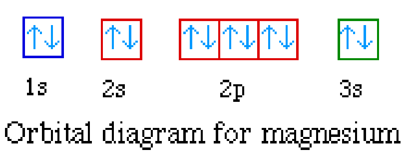
Deciphering electron configurations is simple. The coefficients represent energy levels. The letters represent sublevels, and the superscripts represent electrons in those sublevels and energy levels.

Example: Se $1s^22s^22p^63s^23p^64s^23d^{10}4p^4$

- Selenium have 4 energy levels because the highest coefficient is 4.
- The first energy level contains 2 electrons in an s sublevel.
- The second energy level contains 2 electrons in an s sublevel and 6 electrons in a p sublevel.
- The third energy level contains 2 electrons in an s sublevel, 6 electrons in a p sublevel, and 10 electrons in a d sublevel.
- The fourth energy level contains 2 electrons in an s sublevel and 4 electrons in a p sublevel.
- Selenium has 6 total valence electrons on the fourth energy level.

After mastering neutral atom electron configurations, students can apply their prior knowledge of ions to electron configurations.





Orbital Diagrams: Electron configurations describe the electron cloud, but orbital diagrams give a more visual representation of the electrons' arrangement.

Three significant rules govern electron configuration, and they are most easily understood by looking at an orbital diagram.

1) aufbau principle: In German, "aufbau" means building up. Electrons fill the lowest energy levels of an atom first. No electron will fill a higher level or sublevel until the previous one is full.
2) Pauli exclusion principle: A maximum of two electrons may occupy an orbital. The paired electrons must spin in opposite directions.
3) Hund's rule: In p, d, and f sublevels, multiple orbitals exist. When filling these orbitals, put one electron in each orbital before adding a second electron to any orbital. Notice this illustration uses lines rather than full boxes to illustrate orbitals. Either is acceptable.