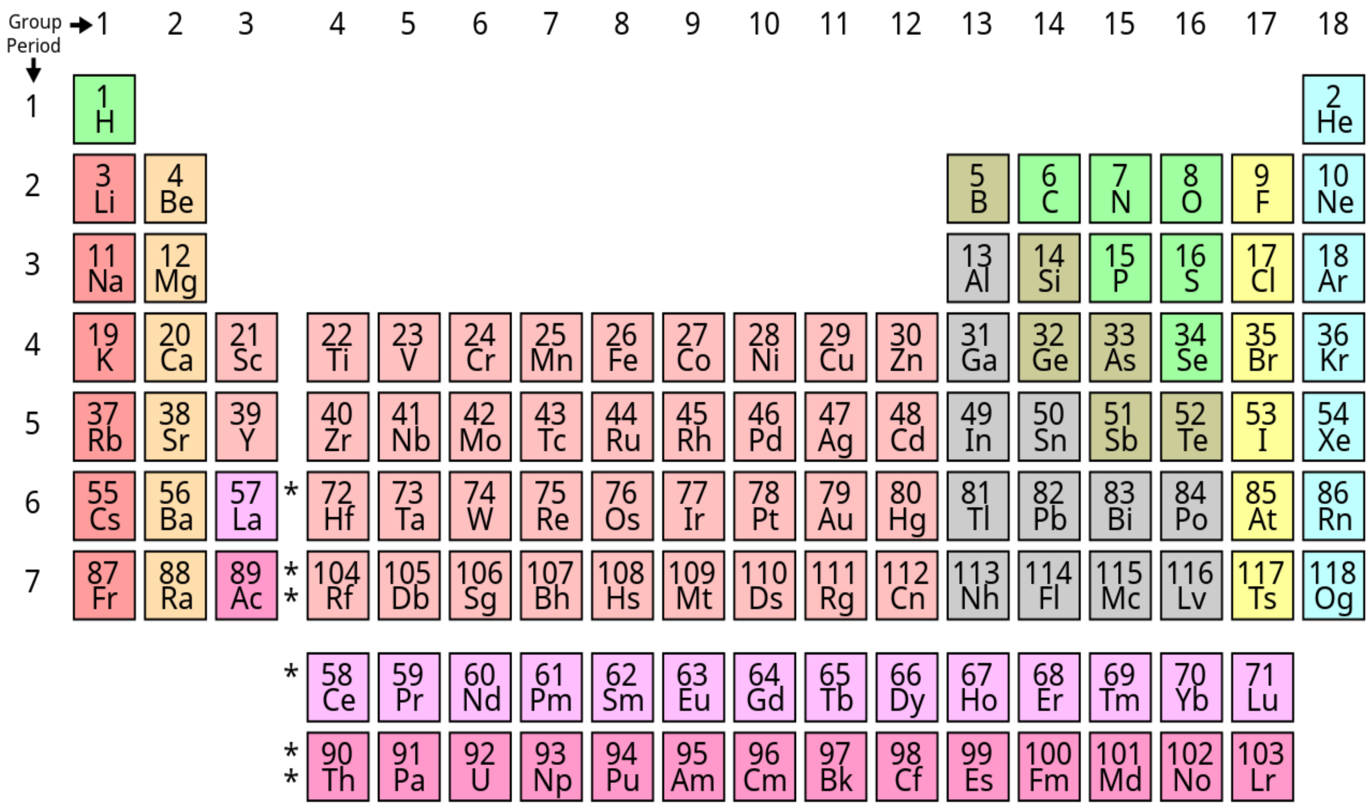
Periodic table

The **periodic table**, also known as the **periodic table of elements**, is a tabular display of the chemical elements, which are arranged by atomic number, electron configuration, and recurring chemical properties. The structure of the table shows periodic trends. The seven rows of the table, called periods, generally have metals on the left and nonmetals on the right. The columns, called groups, contain elements with similar chemical behaviors. Six groups have accepted names as well as assigned numbers: for example, group 17 elements are the halogens; and group 18 are the noble gases.

The elements from atomic numbers 1 (hydrogen) through 118 (oganesson) have all been discovered or synthesized, completing seven full rows of the periodic table. The first 94 elements, hydrogen through plutonium, all occur naturally, though some are found only in trace amounts and a few were discovered in nature only after having first been synthesized. Elements 95 to 118 have only been synthesized in laboratories, nuclear reactors, or nuclear explosions. Numerous synthetic radioisotopes of naturally occurring elements have also been produced in laboratories.

The organization of the periodic table can be used to derive relationships between the various element properties, and also to predict chemical properties and behaviors of undiscovered or newly synthesized elements. Russian chemist Dmitri Mendeleev published the first recognizable periodic table in 1869, developed mainly to illustrate periodic trends of the then-known elements. He also predicted some properties of unidentified elements that were expected to fill gaps within the table. Mendeleev's idea has been slowly expanded and refined with the discovery or synthesis of further new elements and the development of new theoretical models to explain chemical behaviour.



Periodic properties of elements

Atomic radius

The atomic radius of an element is half of the distance between the centers of two atoms of that element that are just touching each other. Generally, the atomic radius decreases across a period from left to right and increases down a given group. The atoms with the largest atomic radii are located in Group I and at the bottom of groups.

Moving from left to right across a period, electrons are added one at a time to the outer energy shell. Electrons within a shell cannot shield each other from the attraction to protons. Since the number of protons is also increasing, the effective nuclear charge increases across a period. This causes the atomic radius to decrease.

Moving down a group in the periodic table, the number of electrons and filled electron shells increases, but the number of valence electrons remains the same. The outermost electrons in a group are exposed to the same effective nuclear charge, but electrons are found farther from the nucleus as the number of filled energy shells increases. Therefore, the atomic radii increase.

Ionization Energy

The ionization energy, or ionization potential, is the energy required to remove an electron from a gaseous atom or ion completely. The closer and more tightly bound an electron is to the nucleus, the more difficult it will be to remove, and the higher its ionization energy will be. The first ionization energy is the energy required to remove one electron from the parent atom. The second ionization energy is the energy required to remove a second valence electron from the univalent ion to form the divalent ion, and so on. Successive ionization energies increase. The second ionization energy is always greater than the first ionization energy. Ionization energies increase moving from left to right across a period (decreasing atomic radius). Ionization energy decreases moving down a group (increasing atomic radius). Group I elements have low ionization energies because the loss of an electron forms a stable octet.

Electron Affinity

Electron affinity reflects the ability of an atom to accept an electron. It is the energy change that occurs when an electron is added to a gaseous atom. Atoms with stronger effective nuclear charge have greater electron affinity. Some generalizations can be made about the electron affinities of certain groups in the periodic table. The Group IIA elements, the alkaline earths, have low electron affinity values. These elements are relatively stable because they have filled *s* subshells. Group VIIA elements, the halogens, have high electron affinities because the addition of an electron to an atom results in a completely filled shell. Group VIII elements, noble gases, have electron affinities near zero since each atom possesses a stable octet and will not accept an electron readily. Elements of other groups have low electron affinities.

In a period, the halogen will have the highest electron affinity, while the noble gas will have the lowest electron affinity. Electron affinity decreases moving down a group because a new electron would be further from the nucleus of a large atom.

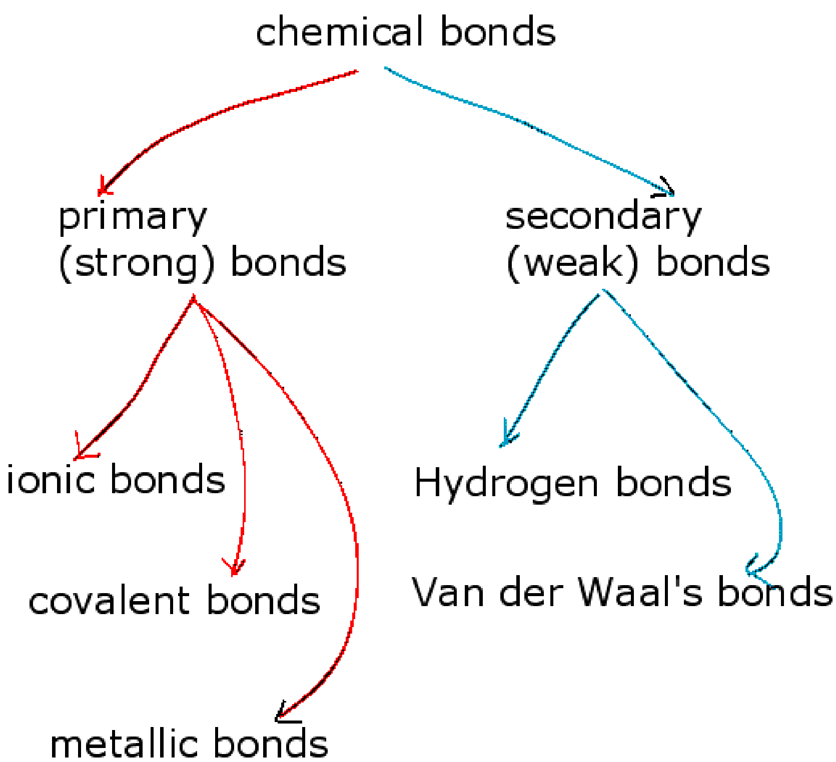
Electronegativity

Electronegativity is a measure of the attraction of an atom for the electrons in a chemical bond. The higher the electronegativity of an atom, the greater its attraction for bonding electrons. Electronegativity is related to ionization energy. Electrons with low ionization energies have low electronegativities because their nuclei do not exert a strong attractive force on electrons. Elements with high ionization energies have high electronegativities due to the strong pull exerted on electrons by the nucleus. In a group, the electronegativity decreases as the atomic number increases, as a result of the increased distance between the valence electron and nucleus (greater atomic radius). An example of an electropositive (i.e., low electronegativity) element is cesium; an example of a highly electronegative element is fluorine.

Chemical bonding

The Classification of Chemical Bonds

There are two major bond classifications, each with identifiable sub-groups:

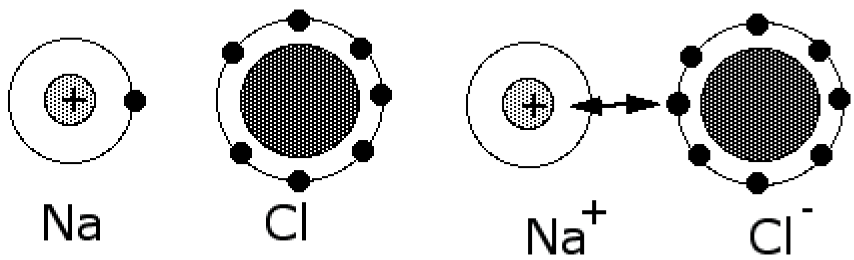


Three Primary Bonds

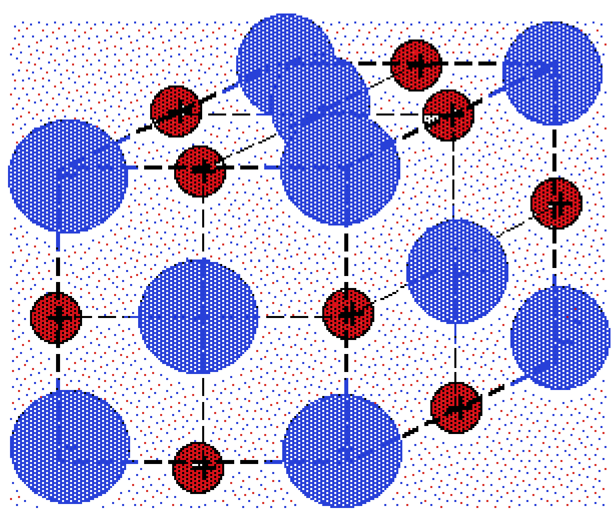
The three types of primary bonding reflect these ways in which atoms can group together by gaining or losing or sharing electrons, so they can get inert gas electron configurations.

Ionic Bonds

Atoms near the left or right sides of the periodic table can lose or gain 1 (or 2) electrons to form charged "ions". For example, a Sodium atom (row 3, column IA) can lose one electron to have 8 valence electrons and become a positively charged "cation". A Chlorine atom (row 3, column VIIA) can gain one electron to have 8 valence electrons and become a negatively charged "anion". These two ions then will be attracted to each other by non-directional electrostatic force and form an *ionic (or electrovalent) bond*.



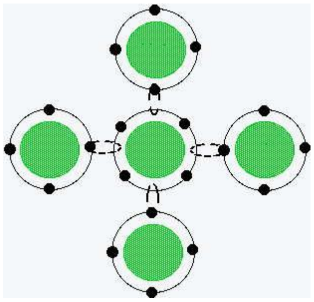
When large numbers of such ion pairs come together an *ionic solid* is formed. Common salt (NaCl) is an ionic solid which has the cubic structure shown on the right.



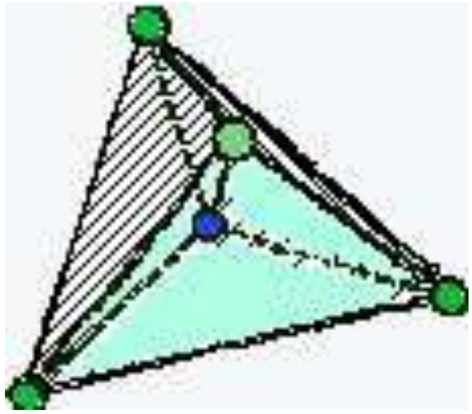
Covalent bonds

Atoms at the center of the periodic table (group **IV**A) have 4 valence electrons. It is difficult to completely loss or gain this many electrons so by compromise they end up sharing electrons.

In the diagram, the central atom is missing the 4 electrons to form an inert gas configuration, but can borrow them for a while from like-minded neighbours to form an electron "cloud" between the two atoms. The atomic cores are then attracted to the negative electron cloud between them, forming a *covalent* bond. The repulsion between electron "clouds" maximises the angles between the bonds.



In 3-dimensions, four equally spaced bonds form a tetrahedral structure. In the diagram, the blue atom is at the center of the tetrahedron forming four bonds with green atoms at the tetrahedral vertices. A sloping light green triangular face is highlighted between two base atoms and the top atom.



Metallic Bonds

A lot of metals fall in the yellow area of the periodic table shown. They share electrons in a different way to covalent bonding.

Metallic bonding occurs between the positive atom cores and the "nearly free" electrons.

In metallic bonding:

• there are no charge requirements,

• there are no directional requirements, and

• there are long range effects.