Properties of Atoms

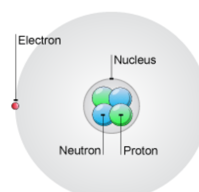
Units and measures:

Time – seconds (s)
Length – meters (m) & angstroms (Å)
(1 Å = 10^{-8} cm = 10^{-10} m, about the length of a covalent bond)
Area – square angstroms
Volume – cubic angstroms
Mass – grams (g), one mole of hydrogen has a mass of 1 gm
Avogadro's number – 6.022 x 10^{23} mol^{-1} – number of atoms/molecules in one mole
Mole – Avogradro's number of atoms/molecules
Temperature – Celsius (C), Kelvin (K)
T(K)=T(C)+273.15

Metric units

Prefix	symbol	factor	Typical use
kilo	k	10^{3}	Kg (kilogram)
centi	c	10^{-2}	cm (centimeter)
milli	m	10^{-3}	ml (milliliter)
micro	μ	10^{-6}	μmole (micromole)
nano	n	10^{-9}	nmole (nanomole)
pico	p	10^{-12}	ps (picoseconds)

Elements and the Periodic Table.



Atomic stucture: Atoms consist of a nucleus that contains positively charged protons and neutral (uncharged) neutrons. The number of neutrons usually equals the number of protons (e.g. Helium), but there are exceptions, such as hydrogen. Electrons are found outside the nucleus. The nucleus is small and the volume of atoms is defined by the extent of the electrons.

Protons = elementary particle with charge of +1, found in nucleus of atom.

Neutrons = elementary particle with charge of 0, found in nucleus of atom.

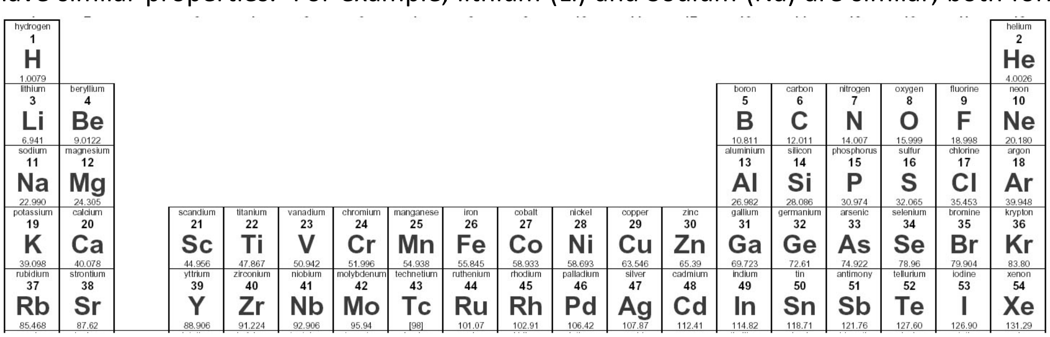
Atomic number = number of protons in the nucleus of the atom, defines the chemical properties of the atom. Hydrogen's atomic number is 1, its nucleus contains one proton.

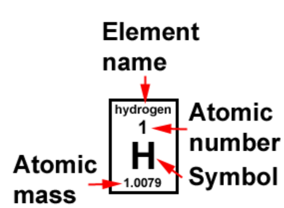
Atomic mass = mass of one mole of atoms, in grams, averaged over all stable isotopes. The atom mass of a pure isotope is equal to the number of protons and neutrons, the most common form of carbon contains 6 protons and 6 neutrons and has an atomic mass of 12. The unit of mass is also called a Dalton (Da).

Isotope = variation of an element that contains a different number of neutrons, e.g. a common isotope of hydrogen is deuterium, its nucleus contains one proton and one neutron and its symbol is ^{2}H. Some isotopes are stable, e.g. ^{2}H, ^{15}N, ^{13}C. Some are unstable (radioactive) and decay to more stable isotopes, e.g. ^{14}C decays to nitrogen 14 (^{14}N) – the most common isotope of nitrogen.

Electrons = elementary particle with charge of -1, found outside of the nucleus, in orbitals (often simplified as shells). Atoms have equal numbers of electrons and protons, i.e. they are electrically neutral. Atomic ions have lost or gained an electron, so they are positively charged, or negatively charged, respectively.

Periodic Table: The elements are arranged in the period table according to common atomic properties that arise from a similar configuration of the electrons around the atom. Elements in each column of the table have similar properties. For example, lithium (Li) and Sodium (Na) are similar; both form +1 ions in water.

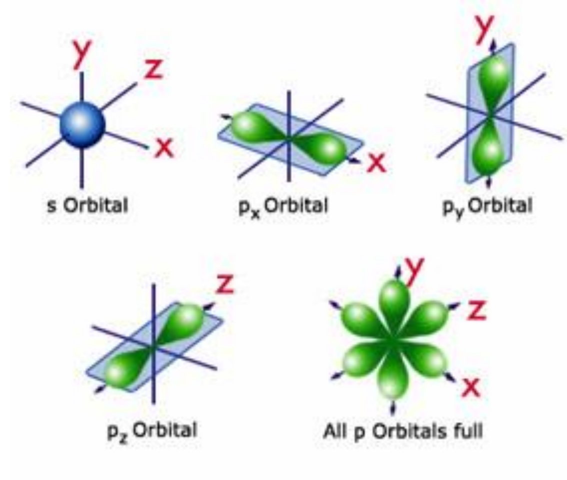




Oxygen and sulfur are similar, both can form two bonds. Each row of the table represents electrons filling additional shells. The first row represents the first shell, the second the 2nd shell, etc.

Electron Orbitals & Shells

Electrons exist in regions of space surrounding the atom called orbitals. Each orbital can hold at most two electrons. Electrons have a property called spin, and can have one of two values of spin, “spin-up” and “spin-down”, often represented by arrows: ↑, or ↓. A single orbital can only contain electrons of opposite spin.



Orbitals of importance in biology are s, p, and d orbitals.

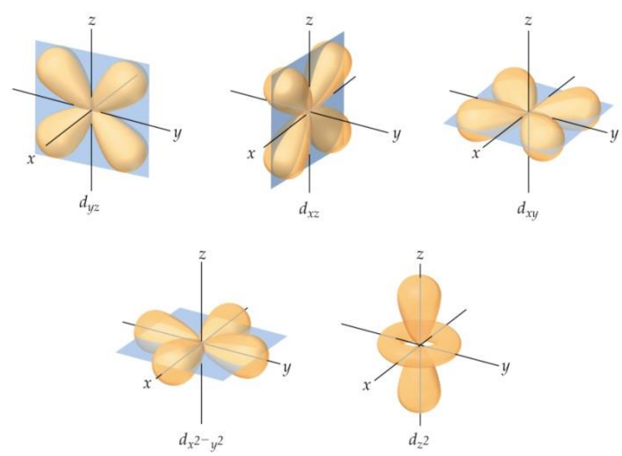
s orbital – spherically symmetric, holding 2 electrons, closest to the nucleus.

p orbital – There are three possible p orbitals, all have the same bi-lobed shape, oriented along the x, y, or z axis. A total of 6 electrons can occupy the 3 different p orbitals, again two electrons/orbital. These are further away from the nucleus.

d-orbitals – There are five possible d orbitals, holding a total of 10 electrons.

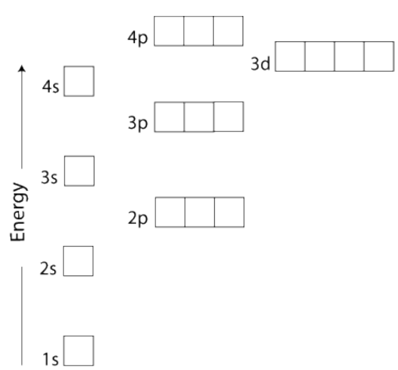
Each row in the periodic table corresponds to a different level. The orbitals found within each level are:

Level	**Orbitals**
1	1s
2	2s and $2p_x$, $2p_y$, $2p_x$
3	3s and $3p_x$, $3p_y$, $3p_x$, and five (5) 3d orbitals
4	4s, 4p, 4d, and 4f orbitals



The levels are often simplified and referred to as shells, e.g. the 1st shell can hold 2 electrons, the 2nd shell can hold 8 electrons. A simpler model of the 3rd shell, ignores the 3d orbitals because they are filled after the 4s orbitals, so it hold 8 electrons as well.

The energy level associated with each orbital is

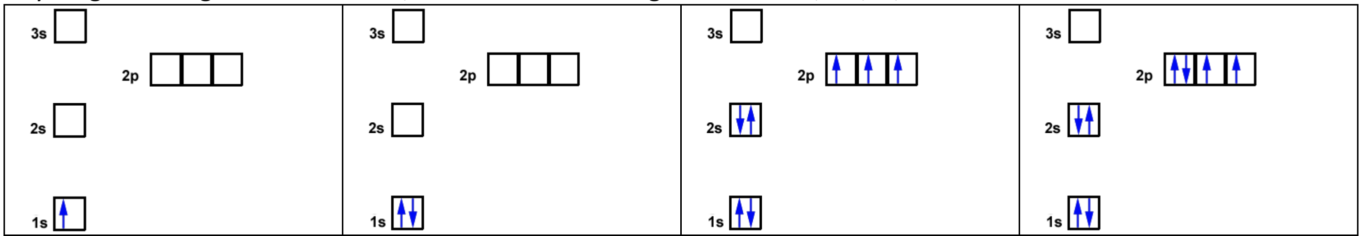


When electrons are added to the oribitals, the following rules are applied:

i) the lowest orbitals are filled first.

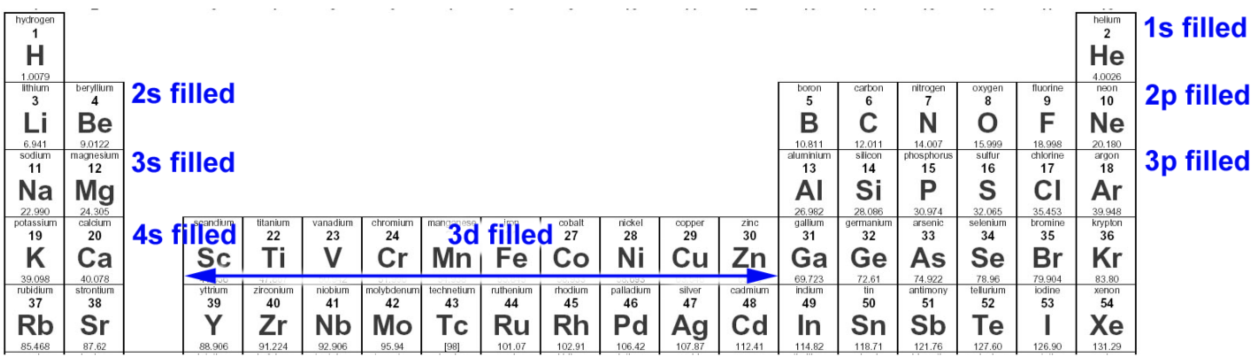
ii) ½ full orbitals are more stable than completely full orbitals, because it decreases repulsion between the two electrons in the same orbital. For example, it is lower in energy to put a single electron in each 2p orbital, than to have one with two electrons, one orbital with one electron, and one orbital with no electrons. Note that the 4s orbitals are filled before the 3d since the 4s orbitals are lower in energy.

Electronic configuration: The number of electrons in each orbital is indicated by a superscript, e.g. $1s^1$ for hydrogen. Using the above rules, the electronic configuration for H, He, N, and O are shown:



Hydrogen (H): $1s^1$	Helium(He): $1s^2$	Nitrogen(N): $1s^2\,2s^2\,2p^3$	Oxygen(O): $1s^2\,2s^2\,2p^4$

The period table reflects the order of filling of the electronic orbitals:



Elements in the second column on the left all have their highest energy s-orbital full, while elements in the right-most column have both their s and p orbitals filled (as well as their 3d orbitals for row 4). The 3d orbital are filled after the 4s because they are higher in energy.

Given the organization of the periodic table, it is quite easy to write out the electronic configuration of any element. For example, Cobalt (Co) is in the fourth row, so its 1s, 2s, 2p, 3s, 3p, and 4s orbitals must be full and the 3d orbital is partially filled, giving an electron configuration of $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^7$. The number of electrons add to 27, which is equal to the atomic number of Co. There are some exceptions to this simple rule, but for the lighter elements, the electronic configuration can be accurately predicted using these rules.