

CHE 2060: What you need to recall from inorganic chemistry, CHE1031

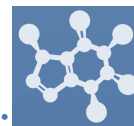
Organic chemistry is all about chemical bonds and the behavior of electrons. Some of the material covered in inorganic chemistry serves as essential background information for organic chemistry. That essential material is summarized here. We will review some of it briefly, but if you feel that you need more of a refresher I strongly advise you to go back to my website for CHE1031 and look through:

- Lecture 6: Atoms, subatomic & quantum structure
- Lecture 7: Chemical bonding

You may find it particularly useful to revisit the electron configuration and Lewis structure worksheets.

Here's what you need to remember or refresh:

- **Science** uses **logic and experimentation** to explain the world around us through **systematic study**.
- **Chemistry** is the science of the **changes** that matter undergoes.
- **Atoms** are the smallest unit of matter that has the chemical and physical properties of a specific type of matter (an element) (Dalton).
- Atoms have tiny, dense, central **nuclei** of protons & neutrons (Rutherford). The nucleus is so tiny that it is like a gnat flying in the center of a huge cathedral.
- **Electrons** occupy most of the **volume** of the atom.
- Electrons are either core or **valence electrons**. Valence electrons are the outermost electrons while core electrons are all other electrons and sit below the valence shell.
- The number of valence electrons in an atom is equal to the atom's **column number**.
- Only valence electrons form **chemical bonds** between atoms.
- Classical Newtonian physics does not explain (or govern) the atoms and subatomic particles. **Quantum mechanics** describes the physical behavior of electrons.
 - Electrons behave as '**wavicles**', neither waves nor particles, but both.
 - The behavior of electrons is dualistic and **uncertain** (Heisenberg).
 - Electrons exist at discrete, **quantum** levels of energy and not at energy between these quantum levels.
- Electrons occupy **orbitals**: three-dimensional shapes that are the electron's most probable location in the space surrounding the nucleus. The shape and energy level of orbitals are defined mathematically by wave functions (Schrodinger).
 - There are **s, p, d and f orbitals** that hold a total of 2, 6, 10 or 14 electrons.
 - All orbitals have **nodes** where electrons do not exist. Nodes are a consequence of wave functions.
 - P, d and f orbitals have **positive and negative lobes** separated by nodes.



- Orbitals **fill** in order of increasing energy levels and the order of filling is not as expected:
 - Each orbital can be occupied by a maximum of two electrons with **opposite spin** (Pauli). And orbitals of equal energy must each be filled one electron before any have two (Hund).
- ~~1s~~
~~2s 2p~~
~~3s 3p 3d~~
~~4s 4p 4d 4f~~
~~5s 5p 5d 5f~~
~~6s 6p 6d 6f~~
~~7s 7p 7d 7f~~
- All atoms would like to have a **full valence shell** (the octet rule).
 - **Ions** form so that atoms can fill their valence shells, sometimes by emptying their valence shell. Ionic bonds are just the electrostatic attraction of oppositely charged ions.
 - **Covalent bonds** form to allow atoms to allow atoms to fill each other's valence shells.
 - **Covalent bonds** are formed by a pair of electrons shared between two atoms with overlapping valence shells. Generally, each atom donates one valence electron to the covalent bond.
 - Covalent bonds can be **non-polar** (the electron pair is equally shared) or **polar** (the electron pair is closer to one atom than the other. Polar bonds are more reactive.
 - Bond polarity is determined by the **electronegativity** values of the bonded atoms. The greater the difference in electronegativity values, the more polar the bond. Electronegativity is a measure of an atom's ability to hold on to its own valence electrons and to grab valence electrons from other atoms.
 - Lewis dots and lines are used to show the bonding structure of molecules as Lewis dot structures. **Lewis dot structures** are drawn using this process:
 - Sum a molecule's valence electrons.
 - Draw the symbols of atoms with by surrounding single atoms with atoms present in multiple copies.
 - Connect all atoms with a line. One line is a bond created by two, shared electron.
 - Use the remaining electrons to complete (fill) the valence shells of outer (peripheral) atoms. Then fill the valence shell of the central atom(s).
 - If there are not enough electrons to fill the valence shell of central atom(s), then convert free electron pairs on the peripheral atoms to pi (multiple) bonds to the central atom(s).
 - Calculate formal charge of atoms as: formal charge = #ve - (dots+sticks)
 - **Resonance structures** are identical except for their arrangement (or placement) of electrons as free pairs or pi (double) bonds.
 - Resonance structures are easy to draw, but **resonance hybrids** are a better representation of reality: pi bonds are shared between equivalent atoms.
 - **Resonance arrows** show the direction in which electrons are pulled.