

CHEM 362

Basic Goals of Chapters 1 and 2

1. Know the classes of inorganic substances.
2. Know the main differences between organic and inorganic structural complexities.
3. Know the main **classes of inorganic reactions**. You don't need to know them in detail because we haven't had a chance to see many examples yet.
4. Know **Basic Thermodynamics: Enthalpy, Entropy, Gibbs Free Energy** and the relationship among them.
5. Know what is an **Equilibrium Constant**, K , and what affects it.
6. Know how **Gibbs Free Energy** and **Equilibrium** are related ($\Delta G = -RT\ln K$).
7. Know what is meant by the **Kinetics of a reaction** and how to diagram a **Reaction Profile**. Recognize that reactions proceed because of a favorable combination of Thermodynamics and Kinetics.
8. Know how to write a simple **Rate Law** for a **first or second order** reaction. Know what is meant by **Order of the Reaction**. Know what affects Reaction Rates.
9. Know what is a **Catalyst** and how it affects rates of reaction.
10. Know that **Atomic Electronic Structure** is (1) the **number of electrons** and (2) the **distribution of electrons** around an atom and that these two features essentially dictate the properties of an atom.
11. Know that the absorption and emission properties of atoms allowed **Rydberg** and later **Bohr** to develop the theories of the **Periodicity and Quantized Nature** of the electronic structures of atoms. Understand that **Energy is Quantized** – only discrete energies are allowed and discrete frequencies of light are absorbed and emitted by atoms ($\Delta E = h\nu$).
12. Understand that allowed wavefunctions (energies) of the (one-electron) hydrogen atom are called **atomic orbitals** and are defined by three quantum numbers: n , l , m_l . Know allowed values of the three quantum numbers; know that orbitals with $l = 0$ are s orbitals (represented to the left of the periodic table); those with $l = 1$ are p orbitals (represented to the right of the periodic table); those with $l = 2$ are d orbitals (represented in the middle of the periodic table); those with $l = 3$ are f orbitals (represented at the bottom of the periodic table). The values of m_l take on integer values $+l, \dots, 0, \dots, -l$.
13. Know that the **electron** also has a **spin quantum number**, m_s , whose value is $\pm 1/2$.

14. Know that hydrogen-like orbitals can be represented as a product of a **radial wavefunction** $R(r)$ and an **angular wavefunction** $\Theta(\theta)\Phi(\phi)$; be able to sketch both parts for any given s , p , or d orbitals and understand the distribution of electron density in the various types of orbitals.
15. Know that we picture each electron in a **many-electron** atom as occupying “**hydrogen-like**” orbitals, but recognize that **no more than two electrons may occupy a single orbital and, if two do occupy a single orbital, their spins (m_s) must be paired ($\uparrow\downarrow$)**.
16. Know that in a many electron atom, inner electrons **shield outer electrons** from the nuclear charge. However, **ns electrons are less-shielded than np electrons** because **s electrons penetrate the nucleus much better than the p electrons**.
17. Know that an electron configuration results when electrons are added to hydrogen-like orbitals of progressively increasing energy. The Periodic table is a representation of the electronic configurations of the elements and follows the **Aufbau Principle**, **Hund’s Rule of Maximum Multiplicity**, and the **Pauli Exclusion Principle**.
18. Understand that because of s orbital penetration, the ns orbital will fill before the $(n-1)d$ orbital; that half-filled and filled orbitals have unusual stability (*e.g.* $[\text{Ar}]3d^54s^1$; $[\text{Ar}]3d^{10}4s^1$; $[\text{Xe}]4f^76s^2$).
19. Know that atomic radii **decrease** as electrons are added to the same shell with corresponding nuclear charge increase (Z_{eff} increases); that atomic radii **increase** as electrons are added to higher principal quantum numbers.
20. Understand that for comparably charged ions, **ionic radii decrease** as do atomic radii; that **ionic radii decrease with increasing positive charge and increase with increasing negative charge**.
21. Understand the **ionization energy** concept; relative to the periodic chart, I.E. generally **increases** from left to right and from bottom to top. Exceptions: s^2 , s^2p^1 ($d^{10}s^2p^1$) and p^3 , p^4 .
22. Understand the **Electron Affinity** (or **Electron Attachment Enthalpy**) concept; relative to the periodic chart, E.A. generally **increases** from left to right and from bottom to top. To avoid any future confusion, let us all agree **that the meaning of a high E.A. is that the atom has a high tendency to accept an electron**. So a **high E.A. is a large negative number**. The negative sign simply means that the process is spontaneous (in other words, A^- is more stable than A).
23. Understand the electronegativity (χ) or **E.N.** may be viewed as an empirically derived property (Mulliken, Pauling, and others have come up with different ways to describe E.N.). Electronegativity **increases** as we go from left to right and from bottom to top in the periodic table.