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.


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 if 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 hat 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 $+l,...,0,...,-l$.

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 an single orbital and, if two do occupy a single orbital, their spins \( (m_s) \) must be paired (↑↓).

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 \((Z_{\text{eff}} \text{ increases})\); 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: \(s^2\), \(s^2p^1(d^{10}s^1p^1)\) 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 wards, \(A^-\) is more stable than \(A\)).

23. Understand the electronegativity \((\chi)\) 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.