Organic Chemistry Notes: Chapter 1

Structure and Bonding; Acids and Bases

Brief History of Organic Chemistry's Origin

1770 – Swedish chemist Torbern Bergman was the first person to express the difference between "organic" and "inorganic." Organic chemistry became the study of compounds in living things.

1828 – Friedrich Wohler discovered that the *inorganic salt*, ammonium cyanate, could be converted to urea, an *organic compound* isolated from urine.

Mid-1800's – Evidence suggested that *organic compounds* all contained the element carbon.

-From section 1.3-

1858 – August Kekule proposed that in all *organic compounds*, carbon always has four (4) "affinity units." Carbon is tetravalent and therefore always forms four (4) bonds when it joins with other elements. Archibald Couper also proposed the above in 1858.

1874 – Jacobus van't Hoff and Joseph Le Bel proposed that the four (4) bonds of carbon are not randomly oriented around carbon but have a specific spatial orientation. (2D → 3D).

~1874 – Jacobus van't Hoff proposed that the four atoms to which a carbon atom is bonded sit at the corners of a regular tetrahedron, with carbon in the center.

Organic Chemistry – the study of carbon compounds. (i.e. methane → DNA, medicines, dyes, polymers, plastics, pesticides, etc.)

1.1 Atomic Structure

Atoms made of:

1. nucleus: protons (+) and neutrons (0)
2. electron cloud: electrons (-)

Atom width: 2 angstroms (10^{-10} m)

The Nucleus:

Atomic number (Z): number of protons in the nucleus
  i.e. carbon: 6 protons
Mass number (A): total number of protons + neutrons in the nucleus
  i.e. carbon-12: 6 protons + 6 neutrons; (6 + 6 = 12)
  i.e. carbon-13: 6 protons + 7 neutrons; (6 + 7 = 13)
Isotopes: element with a different number of neutrons
  i.e. carbon-12 and carbon-13
Atomic weight: the *average* mass number of a lot of atomic isotopes. It is determined by each isotopes natural abundance.

The Electron Cloud:

Shells: the different layers around the nucleus where the electrons move.
Subshells: the different paths that the electrons follow in the shells
  s (sharp), p (principal), d (diffuse), f (fundamental)
Orbitals: location of (paired or unpaired) electrons in subshells
Atoms: the number of protons equal the number of electron (no overall charge)
Ions: the number of protons do not equal the number of electron (positive or negative overall charge)
Subshells and total number of electrons (using g-sec representation):

4p ↑↓ ↑↓ ↑↓
3d ↑↓ ↑↓ ↑↓ ↑↓ ↑↓
4s ↑↓
3p ↑↓ ↑↓ ↑↓
3s ↑↓
2p ↑↓ ↑↓ ↑↓
2s ↑↓
1s ↑↓

1.2 Electron Configuration of Atoms

Ground-state electron configuration (g-sec): the lowest energy arrangement

3 rules to predict g-sec:

Rule 1: The lowest energy levels are filled first. Use the Aufbau system.

Rule 2: Only two electrons can occupy an orbital, and they must be of opposite spin.

Rule 3: If two or more empty orbitals of equal energy are available, one electron is placed in each with their spins parallel until all are half-full.

Try example from class picks and problems 1.1 and 1.2

Shorthand g-sec notations: Sodium – $1s^2\ 2s^2\ 2p^6\ 3s^1$

1. Complete atomic structure problems
2. Complete electron configuration problems
3. Page 31: Problems 1.27, 1.28 (page 31)
4. Bring in an advertisement for an organic compound.
5. Cover textbook.

1.3 Development of Chemical Bonding Theory

Refer to *Brief History of Organic Chemistry's Origin* at the beginning of the notes

3D Diagrams:

Solid lines – are in the plane of the paper
Heavy wedged lines – comes out of the plane (paper)
Dashed lines – goes into the plane (paper)

Van't Hoff Practice

1. Draw the molecular model for CH_4, CF_4, CCl_4, CH_2Cl_2, CF_2Cl_2, and $CHFCl_2$ using solid, wedged, and dashed lines to show its tetrahedral geometry
2. Complete Problem 1.3

1.4 The Nature of Chemical Bonds: Ionic Bonds

Ions – atoms that have gained some negative electrons (becomes -) or have lost some negative electrons (becomes +)

Lewis Diagrams for Ionic Bonds

1.5 The Nature of Chemical Bonds: Covalent Bonds

Molecules – 2 or more atoms that share electrons

Lewis Diagrams for Covalent Bonds

1.6 Formation of Covalent Bonds

Bond strength –

Bond length –

Sigma (σ) bonds – bonds that are "cylindrically symmetrical"

Draw Figures 1.6, 1.7

1.7 Hybridization: sp3 Orbitals and the Structure of Methane

Sp3 hybrids –
Bond angle –
Draw Figures 1.8, 1.9

1.8 The Structure of Ethane

Draw Figure 1.10

1.9 Hybridization: sp2 Orbitals and the Structure of Ethylene

Sp2 hybrids –

Pi (π) bond –

Draw Figures 1.11, 1.12, 1.13

1.10 Hybridization: sp Orbitals and the Structure of Acetylene

Sp hybrids –

Draw Figures 1.14, 1.15

Summary of Hybridization

Hybrid Type	# of sp^n orbitals	# of p orbitals	2p ↑↓ ↑↓ ↑↓ 2s ↑↓
sp^3	4	0	2sp^3 ↑– ↑– ↑– ↑–
sp^2	3	1	2p ↑– 2sp^2 ↑– ↑– ↑–
sp	2	2	2p ↑– ↑– 2sp ↑– ↑–

Type	C-C Angle	C-H Length	C-H Strength	C-C Length	C-C Strength
sp^3	109.5	1.09	438	1.54	376
sp^3	109.5		420 (ethane)		
sp^2	120	1.076		1.33	611
sp	180	1.06		1.20	835

1.11 Bond Polarity and Electronegativity

Polar covalent bonds –
Electronegativity –
Partial positive charge (δ^+)
Partial positive charge (δ^-)
Inductive effect –
Draw Figure 1.16