## <u>Introduction and Review</u>

**Reading**: Wade chapter 1, sections 1-1- 1-8
**Study Problems**: 1-21, 1-23, 1-25, 1-26, 1-34, 1-35

**Key Concepts and Skills**:

- Draw and interpret Lewis, Condensed, and line angle structural formulas. Calculate and show which atoms bear formal charges
- Be able to write electron configurations for elements in the first three rows of the periodic table.

**Lecture Topics:**
**I. Review**

a. Basic structure of atoms
b. Electron configurations of the elements. The **Aufbau principle** allows one to build up the ground-state electronic configurations of atoms in the periodic table. Quantum numbers used to describe energy of electrons in an atom: principle (n=1,2,3… the *shell*), angular momentum [(l=n-1), the *subshell*, and $m_l$ (-l..0..l ) the *orbitals*], and spin ($m_s$ ±1/2). Note the Pauli exclusion principle and Hund's rule for orbital filling.

| | |
|---|---|
| $1s^2$ | 2 electrons |
| $2s^2\ 2p^6$ | 8 electrons (octet rule) |
| $3s^2\ 3p^6\ 3d^{10}$ | 18 electrons |

c. The closer an electron is to the nucleus, the lower in energy it is; electrons in s orbitals are closer to the nucleus and thus lower in energy than electrons in p-orbitals.
d. Know the electron density distributions and shape of s orbitals and p orbitals.

We are concerned with the valence electrons – those in the outer shell – which participate in bonding.

II. Types of bonding: Ionic and Covalent
Lewis: a filled shell of electrons (noble gas configuration) is especially stable.
How to obtain a filled shell:

a. Ionic bonding: involves a transfer of electrons between atoms, resulting in ions of opposite charge that electrostatically attract each other. Usually results from the interaction of group IA, IIA metals with non-metals.
b. Covalent bonding: electrons shared rather than transferred – the most common type of bonding in organic compounds

Representations: Lewis structures, condensed formulas, line-angle formulas

a. in Lewis structures, each valence electron is symbolized by a dot; a bonding pair of electrons is symbolized by a pair of dots or a dash. Lewis structures are arranged so that all atoms have a noble gas configuration. Multiple bonds result from the sharing of two or more pairs of electrons between two atoms; lone pairs on heteroatoms (N, O, halogens) are symbolized by a pair of dots.

b. Condensed formulas are written <u>without</u> showing all the individual bonds.
c. For line angle formulas (skeletal or stick-figure), bonds are represented by lines. The junction of two lines is a carbon atom. N, O, halogens are shown, H is omitted, except when bonded to a heteroatom.

Example:

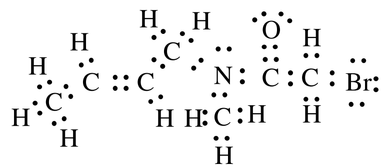
Condensed Structure

$CH_3CHCHCH_2N(CH_3)COCH_2Br$

or

$CH_3CH{=}CHCH_2N(CH_3)COCH_2Br$

Lewis structures



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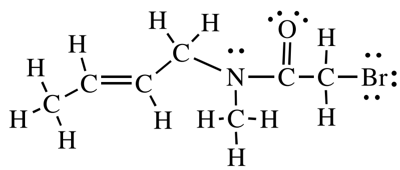
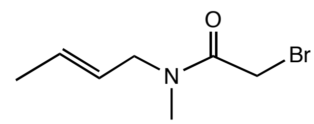


Line-Angle formula



•An Atom's valence is the number of bonds it usually forms; memorize: neutral carbon is always four valent, neutral nitrogen is always three valent, neutral oxygen is always two valent, neutral halogens and hydrogen are always one valent. See table below for positive and negatively charged variants of each of these elements.

• *Non-polar covalent bonds:* equal sharing of electrons between two of the same atoms or atoms of similar electronegativity : H–H or C–H

•*Polar covalent bonds:* bond between different elements in which electrons are attracted more strongly to one of two nuclei. Consult periodic table and Pauling electronegativities to determine whether a given bond will be polar and the direction and magnitude of its dipole moment. Examples: N–B, N–O, C–B, C–Br.

•Formal charges allow us to determine which atoms bear most of the charge in neutral/charged molecules. FC = group # - [non-bonded $e^-$s + 1/2 bonding $e^-$s]

•***Never*** exceed eight valence electrons for second row elements. If you find more than eight total electrons around an atom, ionic bonding may also be involved. Example: compute formal charges on all atoms and draw a Lewis structure for $CH_3NH_3Cl$

# Bonding patterns for organic molecules

| | positive | neutral | negative |
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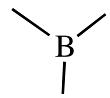
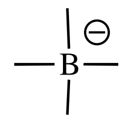
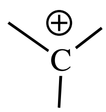
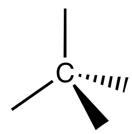
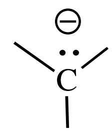
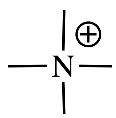
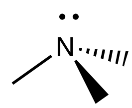
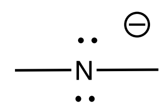
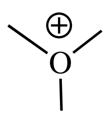
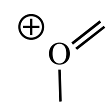
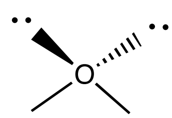
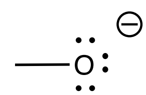
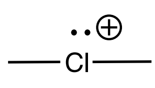
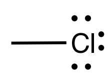
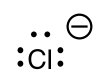
| B | |  |  |
| C |  |  |  |
| N |  |  |  |
| O |  or  |  |  |
| Cl |  |  |  |

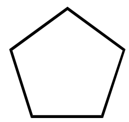
Additional Problems for practice:

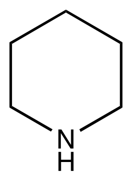
1. Write Lewis (electron-dot) structures for these molecules:
(a) $AlH_3$ (b)$CH_3SCH_3$ (c)$CH_2{=}CHCl$ (d)$CH_2{=}CHCH{=}CH_2$

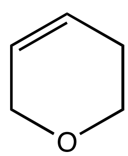
(e) $CH_3CN$

2. Give the ground state electronic configurations of the following elements. (a) sodium (b)aluminum (c)silicon (d) calcium

3.Draw Lewis Structures that correspond to each of the following line-angle formulas:

a. 

b. 

c. 

4. Draw Lewis structures that correspond to each of the following condensed formulas:

(a) $CH_3CH_2CH_2Br$
(b) $CH_3NH_2$
(c) $CH_3CH_2OCH_2CH_3$

5.Draw line-angle formulas that correspond to each of the following condensed formulas.
(a) $CH_3COCH_3$
(b) $CH_3COOH$
(c) $CH_3CH_2CHO$
(d) $CH_3CH(OH)CH_3$