

CHEM 2323

Unit 1 – General Chemistry Review

I. Atoms

- A. The Structure of the Atom
- B. Electron Configurations
- C. Lewis Dot Structures

II. Bonding

- A. Electronegativity
- B. Ionic Bonds
- C. Covalent Bonds
- D. Bond Polarity

III. Covalent Bonding

- A. Drawing Lewis Dot Structures
- B. Isomers
- C. Formal Charge

IV. Molecules

- A. Molecular Structure
- B. Dipole Moment
- C. Intermolecular Forces

V. More on Covalent Bonding and Structure

- A. Valence Bond Theory
- B. Molecular Orbital Theory
- C. Resonance

VI. Acid-Base Theories

- A. Brønsted-Lowry Theory
- B. Lewis Acid-Bases Theory

Unit 1 – General Chemistry Review

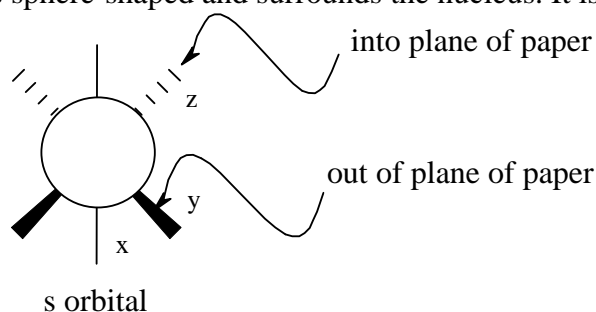
I. Atoms

A. The Structure of the Atom

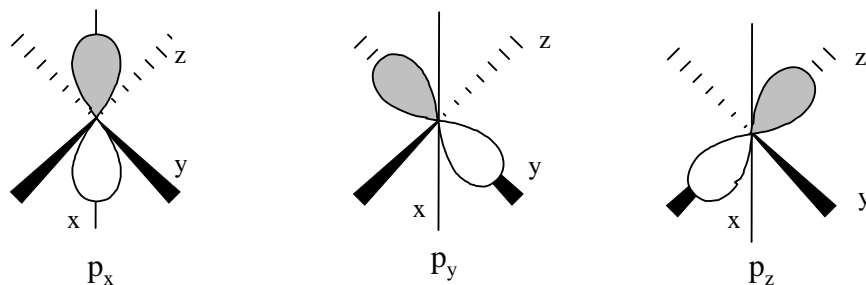
Atomic structure is currently based on quantum mechanics, a statistical theory which predicts the probability of finding an electron around the nucleus.

An electron can occupy s, p, d or f orbitals. The s and p orbitals that compose the **valence shell** are important in organic chemistry.

The **s orbital** is sphere-shaped and surrounds the nucleus. It is the lowest energy orbital.

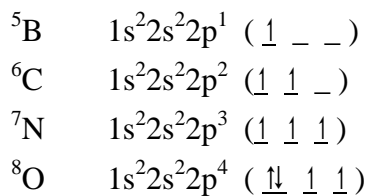


The **p orbital** is propeller-shaped. There are 3 degenerate (of the same energy) p orbitals, aligned orthogonal to each other.



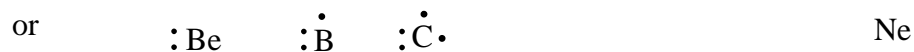
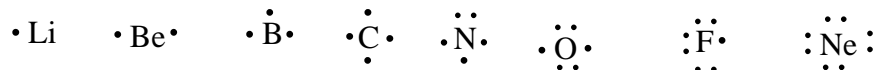
B. Electron Configurations

In ground state electron configurations, electrons occupy the lowest energy orbitals possible. In the case of degenerate orbitals, electrons occupy the orbitals singly with the same spin, unless pairing is essential (Hund's Rule).



C. Lewis Dot Structures

These are used to conveniently illustrate the valence electrons for the representative elements. The number of valence electrons is easily determined from the group number.



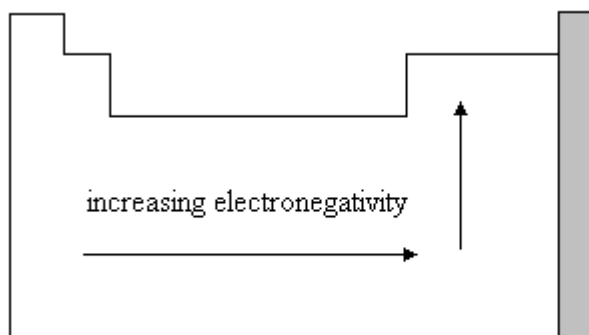
II. Bonding

Bonding between elements occurs because the compound that results is more stable than the individual atoms. Bonding is governed by the tendency of the representative elements to achieve an octet of electrons.

A. Electronegativity

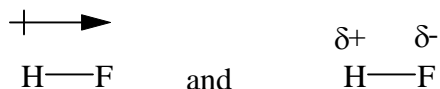
Electronegativity is the relative tendency of a covalently bonded atom to pull e^- density from the bond toward itself. It is expressed on the Pauling scale (Table 1 – handout packet).

General trend:



Fluorine has the highest electronegativity, which means that when it is bonded to any other atom, it has the greater ability to pull the electron density of that bond toward itself.

There are two ways to illustrate bond polarity:

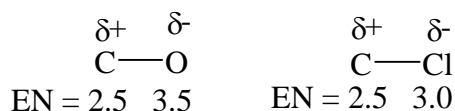


Either method is acceptable, though δ^+/δ^- is much more useful in organic chemistry.

Carbon and hydrogen have similar electronegativities, so that the C – H bond is generally considered non-polar.



Carbon, when bonded to other non-metals, is usually less electronegative, so typically the partial positive is on C.



III. Covalent Bonding

A. Drawing Lewis Dot Structures

The octet rule: The ultimate goal is for all atoms to end up with 8 e⁻s. This is a must. The only exceptions are:

- (a) H can have only 2 e⁻ (1 bond)
- (b) Be prefer to have only 4 e⁻ (2 bonds)
- (c) B and Al prefer to have only 6 e⁻ (3 bonds)

(1) Draw out a preliminary skeletal structure, attaching all atoms with single bonds.

- (a) least EN atom is central (except H)
- (b) oxygens rarely bond to each other
- (c) in oxyacids, H is attached to oxygen, not the central atom
- (d) go for symmetry (though not as important for organic compounds)
- (e) the following is a list of *preferred* # bonds (octet rule takes precedence)

—H 1 bond (a must)

$\begin{array}{c} | \\ \text{—C—} \\ | \end{array}$ 4 bonds

$\begin{array}{c} \cdot\cdot \\ \text{—N—} \\ | \end{array}$ 3 bonds

$\begin{array}{c} \cdot\cdot \\ \text{—O—} \\ \cdot\cdot \end{array}$ 2 bonds

$\begin{array}{c} \cdot\cdot \\ \text{—X:} \\ \cdot\cdot \end{array}$ 1 bond
 (X = halogens)

(2) Calculate total number of valence electrons (group #) present in the molecule. Add 1 e⁻ per negative charge. Subtract 1 e⁻ per positive charge.

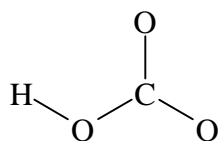
(3) Subtract # e⁻s used in the skeletal structure from the total valence electrons.

(4) Place the remaining electrons around atoms needing an octet. Begin with atoms attached to the central atom. Put lone pairs on the central atom last.

(5) If an atom does not have an octet, move a lone pair from an adjacent atom in between the two to make a double bond. Repeat until all octets are obtained.

example: Draw the Lewis dot structure for the hydrogen carbonate ion (HCO₃⁻).

rule 1: Here is the preliminary skeletal structure that fits best with this rule.



- C is least electronegative of the two choices (H excluded).
- oxygens are not bonded to each other
- this is an oxyacid, so H is attached to O
- this is fairly symmetrical
- preferred bonds is not terribly important with the inorganic compounds, also, this item is best considered after the Lewis dot is complete

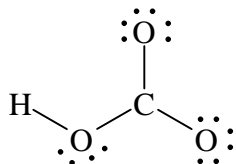
rule 2:

H	(1 x 1) =	1
C	(4 x 1) =	4
O	(6 x 3) =	18
(-)		<u>1</u>
total electrons	=	24

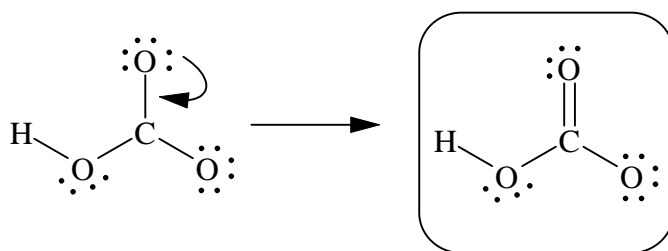
rule 3:

total electrons.....	24
- electrons used in skeletal.....	-8 (4 bonds)
remaining electrons	16

rule 4:



rule 5: Notice that in the structure above, C still doesn't have 8 electrons. So, we take a lone pair of an adjacent atom, and move that pair in between to form a double bond. There are three adjacent atoms, so just pick any one.

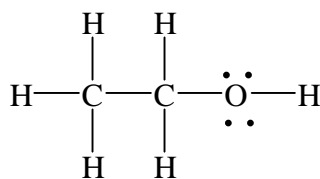


The fact that different Lewis dot structures (all correct) can be obtained by moving lone pairs from different atoms to make the double bond will be discussed in the resonance section, below.

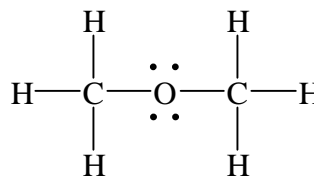
B. Isomers

Isomers are different compounds with the same molecular formula. There are many types of isomers. Isomers that have the atoms attached in different orders are called **constitutional** or **structural isomers**.

e.g., C_2H_6O



ethanol



dimethyl ether

C. Formal Charge

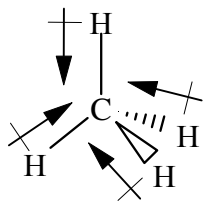
Formal charge compares the number of valence electrons for an atom to the number of electrons nearest that atom in a Lewis dot structure. Formal charge is very useful in organic chemistry; it is used to judge the quality of resonance structures and it helps with understanding how organic reactions occur.

tetrahedral		109.5°
trigonal pyramidal		107°
angular		104.5°

Notice that bond angles decrease as lone pairs are added to the central atom. This is because lone pairs take up more space than atoms of any kind.

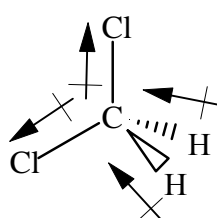
B. Dipole Moment

Dipole moment (μ) is expressed in Debye (D), and is a measure of the net molecular polarity. It is the result of the sum of all individual bond polarities. Non-polar molecules have no dipole moment. More polar molecules have higher dipole moments.



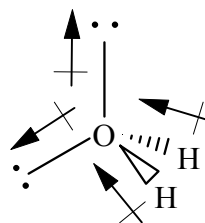
all arrows cancel

$$\mu = 0 \text{ D}$$



arrows don't
cancel

$$\mu = 1.62 \text{ D}$$



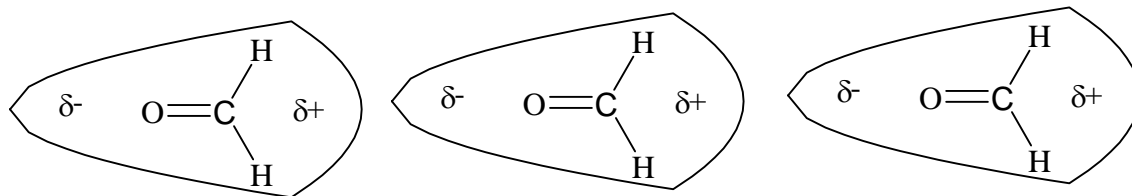
electron pairs contribute
to dipole moment

$$\mu = 1.85 \text{ D}$$

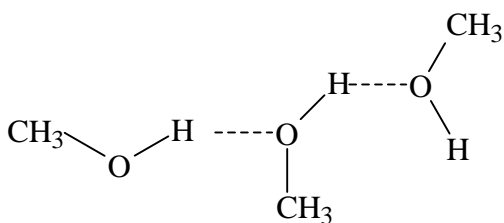
C. Intermolecular Forces

Intermolecular forces are the forces that hold molecules together. All intermolecular forces experienced between neutral molecules are generally referred to as **van der Waals forces**. These may be broken down into three categories.

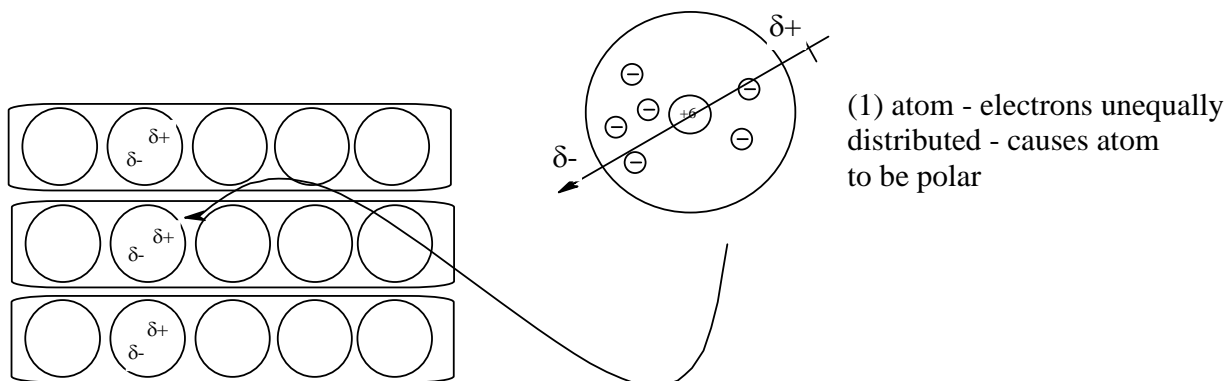
(1) **Dipole-dipole forces** occur between polar covalent molecules. The electronegative end of one molecule is attracted to the electropositive end of another molecule.



(2) **Hydrogen bonding** is a special, stronger version of the dipole-dipole force. It occurs in molecules with H – N, H – O, or H – F bonds. The H in one molecule is attracted to the N, O, or F in another molecule.



(3) **London or dispersion forces** are temporarily induced dipoles caused by instantaneous, unequal distribution of electrons about the nucleus. London forces occur in all molecules, but are only apparent in non-polar covalent molecules. In these, none of the other forces exist.



(2) molecules - polar atom in one molecule has a slightly polarizing effect on the nearest atom in another molecule