Chemistry 6A Fall 2007

Dr. J. A. Mack

Wednesday

10/3/07

Exam 1: Friday 10/5/07 (here in lecture)

What will be covered on the exam?

What do I need to bring?

- Chapter 1-3 (all)
- Chapter 4: (4.1-4.5 and 4.10)
- Any thing from lab as well

Bring a Pencil, Eraser, Calculator and scamtron form 882

YOU NEED TO KNOW YOUR LAB SECTION NUMBER!

How should I prepare for the exam…

1. Get some sleep the night before.
2. Go over your quizzes.
3. look at your HW
4. look over additional HW problems
5. Focus on what you know first

What should I not do…

1. Put off studying until Thursday night
2. Party Thursday night! ***(there will be plenty of time for that later)***
3. Snarf down 4 doughnuts and 3 red-bulls right before the exam!

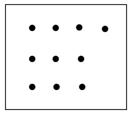
N_2 Each nitrogen atom needs 8 electrons to complete an octet…
But each nitrogen has only 5 valence electrons!
As a result, the electrons must be shared.

There are 10 electrons available

:N ≡ N:

Triple bond!

:N:::N:



N_2 needs 16 electrons (2×8)

Therefore 6 electrons must be shared (16–10)

The rest are lone pairs that complete the octet.

How do we know which atom is in the center of a molecule?

In general, the atom that is the lowest and to the left on the periodic table will be in the center.

Most often, the atom with the lowest *electronegativity* will be in the center.

TABLE 4.4 Electronegativities for the common representative elements

Increasing electronegativity →

Decreasing electronegativity ↓

		H 2.1					
Li 1.0	Be 1.5		B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2		Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0		Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0		In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9						

S would be favored in the center over O

As one moves from left to right and bottom to top on the periodic table, the *ELECTRONEGATIVITY* of an atom decreases.

SO_2

sulfur dioxide

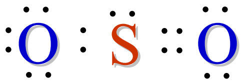
Sulfur and each oxygen have 6 electrons: (6×3 = 18 electrons available)

Each atom needs 8 electrons: (3×8 = 24)

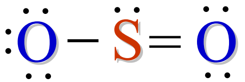
This means that 6 electrons must be shared. (24–18 = 6)

Place the S-atom in the center, each O-atom on either side.

add in the shared electrons

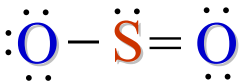


add in the rest of the electrons as lone pairs to complete the octets

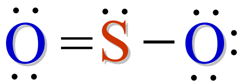


Dr. Mack.....

Does it matter where the double bond in SO_2 is?



on the right…



or on the left…

it is the same molecule

Multiple Bonds:

When 2 electrons are shared between atoms a *single* bond exists

H–Ö–H water has only single bonds

When 4 electrons are shared between atoms a *double* bond exists

Ö=Ö an oxygen molecule has a double bond

When 6 electrons are shared between atoms a *triple* bond exists

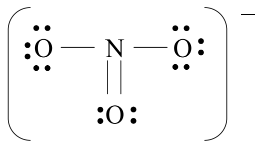
:C≡O: carbon monoxide has a triple bond

Lewis Dot structures for Polyatomic Anions:

NO_3^- Nitrate anion one extra for the charge

	N		3O			
	5	+	3(6)	+ 1	=	24 electrons available
	8	+	3(8)		=	32 electrons needed
	32	–	24		=	8 electrons shared
					=	4 bonds (8 ÷ 2)
					=	8 lone pairs (24 ÷ 2) = 12 - 4

N goes in the center as it is to the right of O



brackets indicate charge

Molecular Geometry and Bonding Theories (4.8)

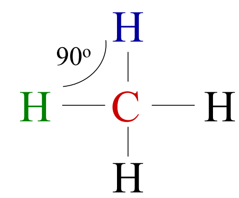
Lewis structures tell us how atoms are connected in a molecule:

bonds *(bp)* & lone pairs (***lp***) etc…

The 3–D shape of a molecule is however, determined by its bond angles.

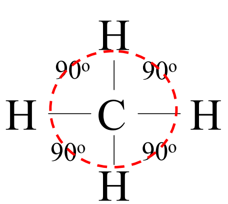
Consider methane, CH_4:

The Lewis structure suggests that the H–C–H bond angles are 90° & that the molecule is flat.



Experimentally however, the H–C–H bond angles are found to be 109.5°

If this is the case, the molecule cannot be planar…

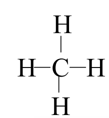


90°+ 90° + 90° +90° = 360°

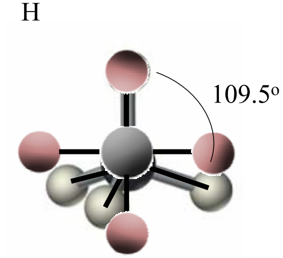
109.5° + 109.5° + 109.5° + 109.5° = 438.0°

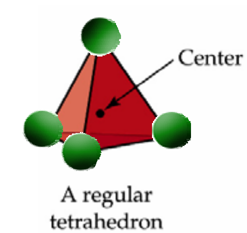
The methane molecule must be 3-D since the sum of the angles is greater than 360 degrees!

To accommodate the 109.5° bond angles, the atoms adopt a new 3-D geometry:



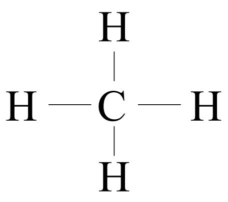
The H-atoms fit at the corners of a regular *tetrahedron* shape with the carbon at the center.



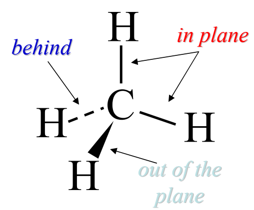


The molecular geometry of CH_4 is said to be "***tetrahedral***"

Drawing a Lewis structure in 3-D perspective:



becomes…



The solid lines indicate that the bond is in the plane of the screen, the dashed indicates that the bond is behind the screen plane and the triangle indicates that the bond is coming out of the screen plane.

Molecular Shapes: Valence Shell Electron Pair Repulsion

A molecule can be described in terms of the distribution of the bonding atoms about the central atom:

Molecular Geometry (MG)

A molecule can be described in terms of the distribution of the bonding pair electrons (bp) and lone pair electrons (lp) about the central atom:

Electronic Geometry (EG)

Molecular Shapes: Valence Shell Electron Pair Repulsion

In order to predict molecular shape, we assume the valence electrons of each atom in the molecule repel one another.

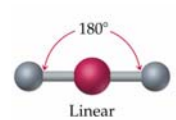
When this occurs, the molecule adopts a 3D geometry that minimizes this repulsion where:

This process is known as:

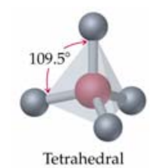
Valence Shell Electron Pair Repulsion theory.

(VSEPR)

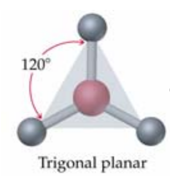
There are three fundamental shapes that describe the *electronic geometries* for molecular shape:



2 pairs of electrons



4 pairs of electrons

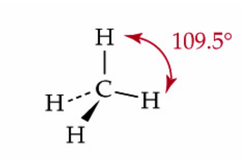


(triangular)

3 pairs of electrons

balloons time!

4 bonding pairs of electrons about the central atom, no lone pairs:

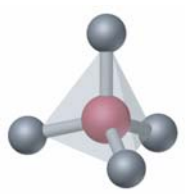


Methane: CH_4

4 groups of electrons (all bonding)

Electronic Geometry: ***Tetrahedral***

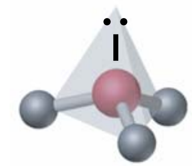
Molecular Geometry: ***Tetrahedral***



Tetrahedral

When there are no lone pairs of electrons in a molecule, the *molecular* and *electronic* geometries are the same!

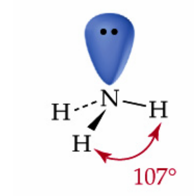
4 total pairs of electrons about the central atom, three bonding, one lone pair:



Trigonal pyramidal

Ammonia: NH_3

Molecular Geometry (MG):
Trigonal pyramidal
Electronic Geometry (EG)
Tetrahedral



4 groups of electrons (3 bp, 1 lp)

Electronic Geometry: ***Tetrahedral***

Molecular Geometry: ***Trigonal Pyramidal***