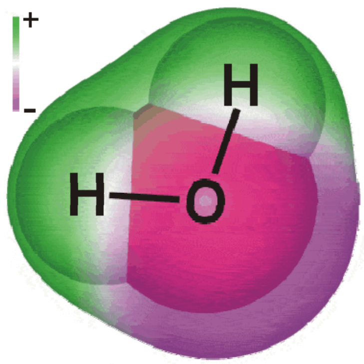
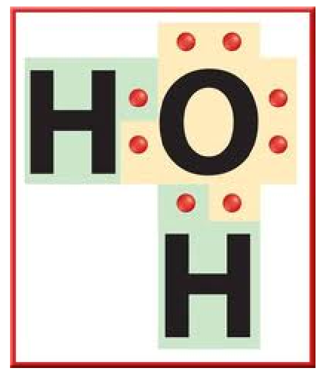
Chapter 5 Chemical Bonding:

The Covalent Bond Model





- Atoms with similar ionization energy and electronegativity DO NOT form ionic bonds.
- There is NO electron transfer!
- Electron pairs are shared to form a covalent bond.

A covalent bond is a chemical bond resulting from two ________ attracting the same shared ________.

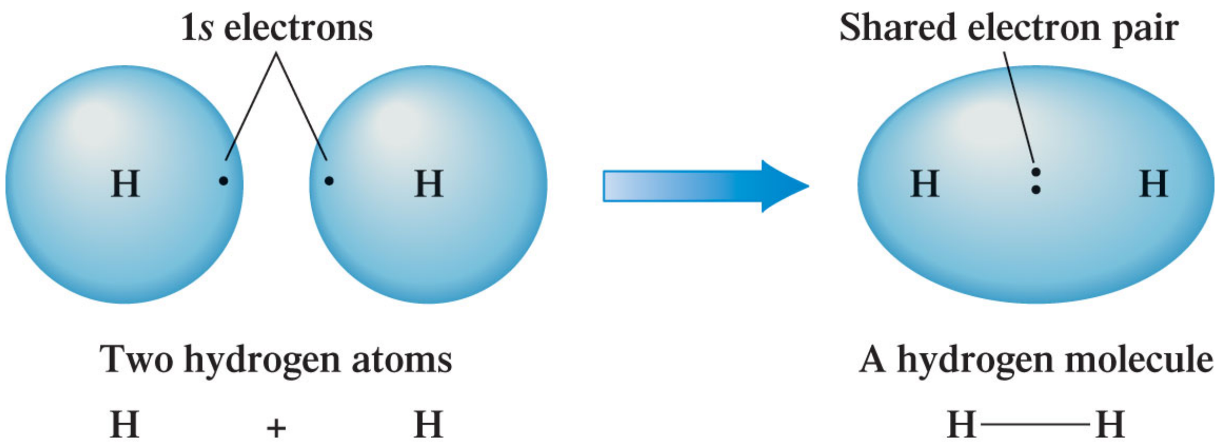


Fig 5.1 Electron sharing can occur only when electron orbitals from two different atoms overlap.

$$H\cdot + \cdot H \rightarrow H:H$$

$$H-H$$

$$H_2$$

Each atom in H_2 has the electron configuration of ____

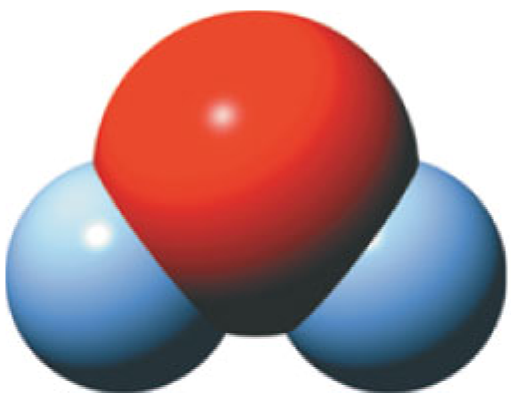
Ch 5.2 Lewis Structures for Molecular Compounds

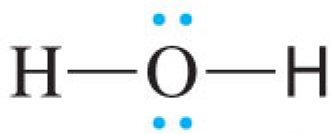
H· + ·F: → H : F : or H - F :

bonding electrons

nonbonding electrons

F, with ___ valence electrons, forms 1 covalent bond.

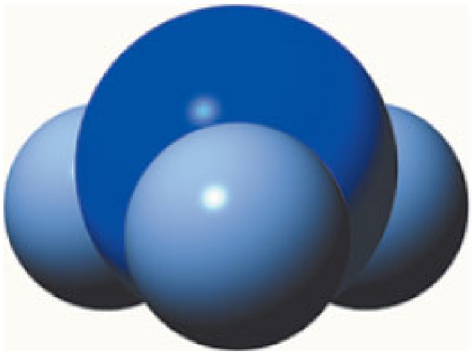


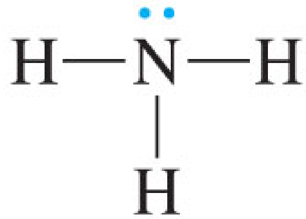


Water, H_2O

O, with ___ valence electrons, forms 2 covalent bonds.

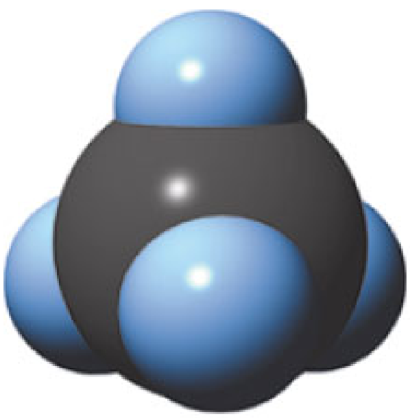
Nonbonding electrons are also called ________ ________ .

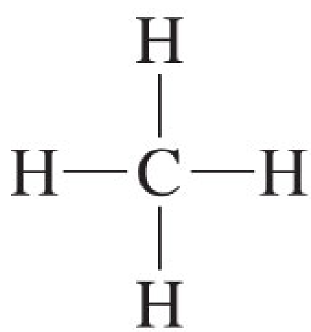




Ammonia, NH_3

N ____ valence electron 3 covalent bonds





Methane, CH_4

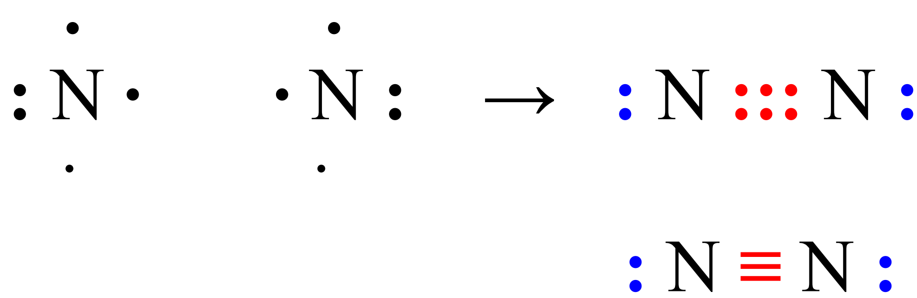
C ____ valence electron 4 covalent bonds

The number of covalent bonds formed by a nonmetallic element is directly correlated with the number of electrons it must share in order to obtain an ____________ of electrons.

Ch 5.3 Single, Double, and Triple Covalent Bonds

Some atoms must share ______ than one pair of electrons to obtain an octet of electrons.

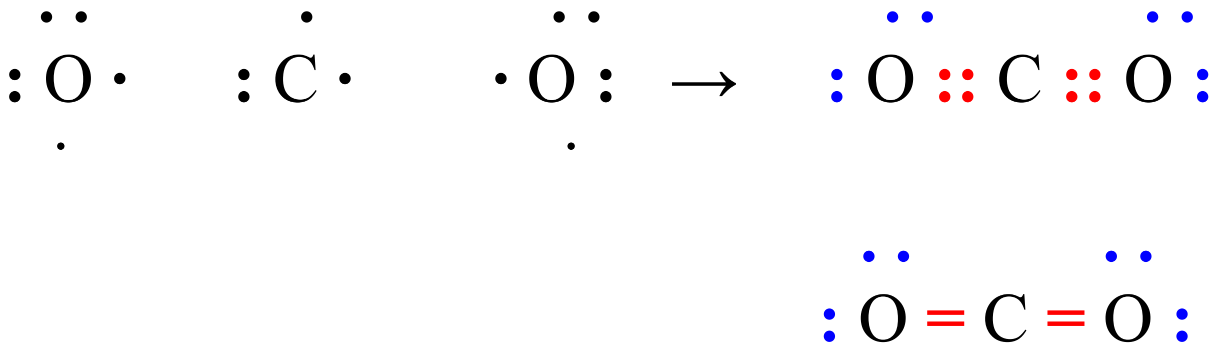
Single bond	-	2 atoms share 1 pair of e
Double bond	-	2 atoms share 2 pairs of e
Triple bond	-	2 atoms share 3 pairs of e



1 triple bond

each N in N_2 has an ______

Carbon dioxide has ____ double bonds



each atom in CO_2 has a complete octet

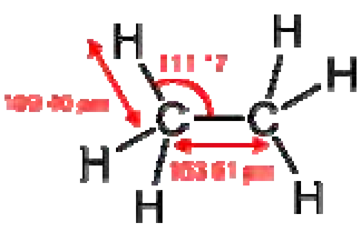
Supplemental material

Covalent bonds, bond energy, and bond length

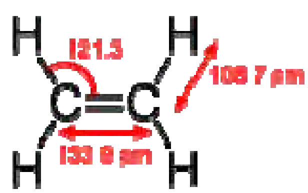
Single → Double → Triple

$$\longrightarrow$$

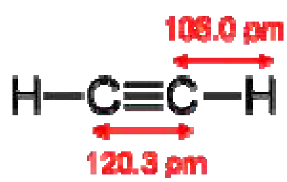
bond energy increases
bond length



ethane



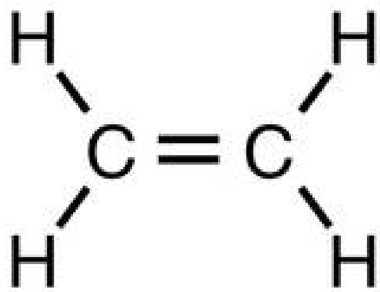
ethylene



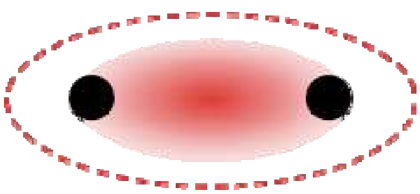
acetylene

Supplemental material: Sigma and Pi Bonds

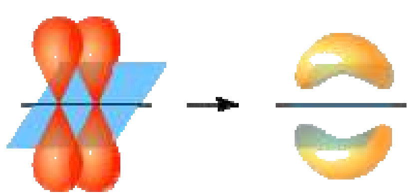
ethylene



double bond = 1 sigma (σ) bond & ___ pi (π) bond

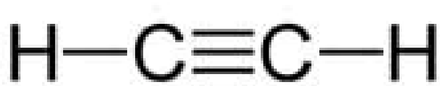


σ bond



π bond, 2 overlapping p orbitals (2 lobes/orbital) restricted motion

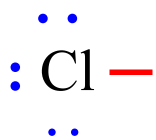
acetylene



triple bond = 1 σ bond & ___ π bonds

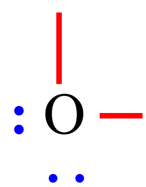
Ch 5.4 Valence Electrons and Number of Covalent Bonds Formed

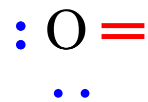
Group VIIA - 7 valence electrons - 1 covalent bond
& ____ lone pairs



1 single bond

Group VIA - 6 valence electrons - 2 covalent bonds
& ____ lone pairs

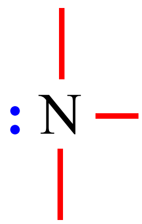




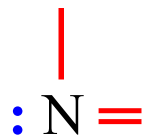
2 single bonds

1 double bond

Group VA - 5 valence electrons - 3 covalent bonds
& ______ lone pair



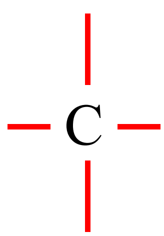
3 single



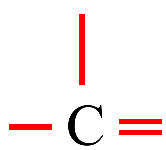
1 single
& 1 double

1 triple

Group IV - 4 valence electrons - 4 covalent bonds
& ______ lone pairs

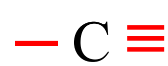


4 single



2 single
& 1 double

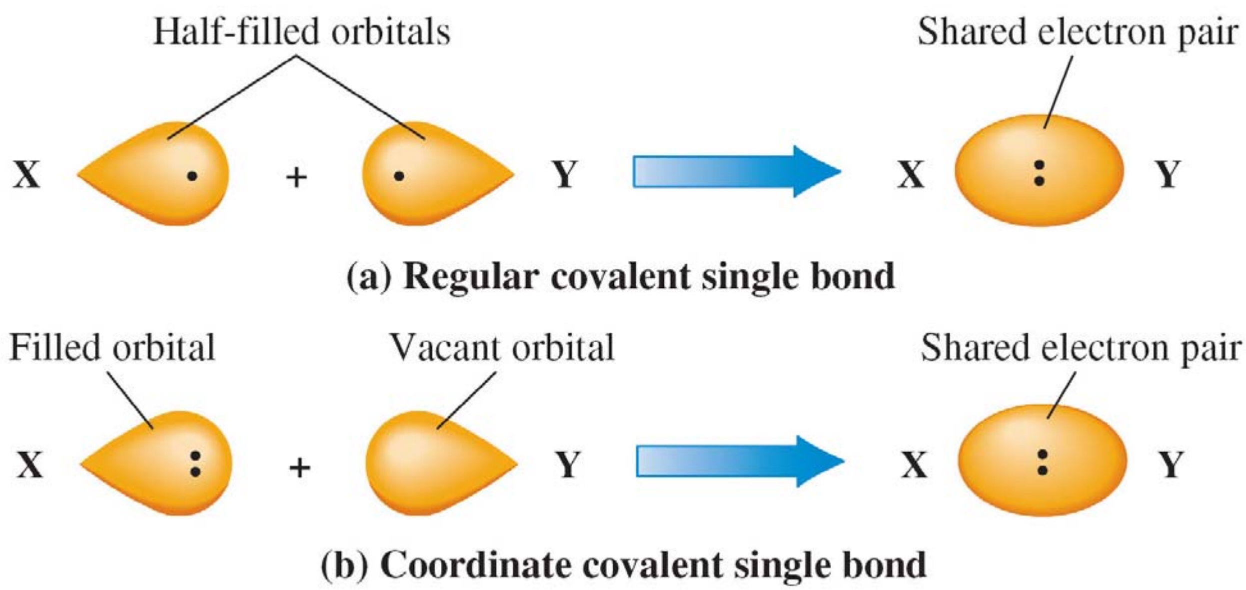
2 double



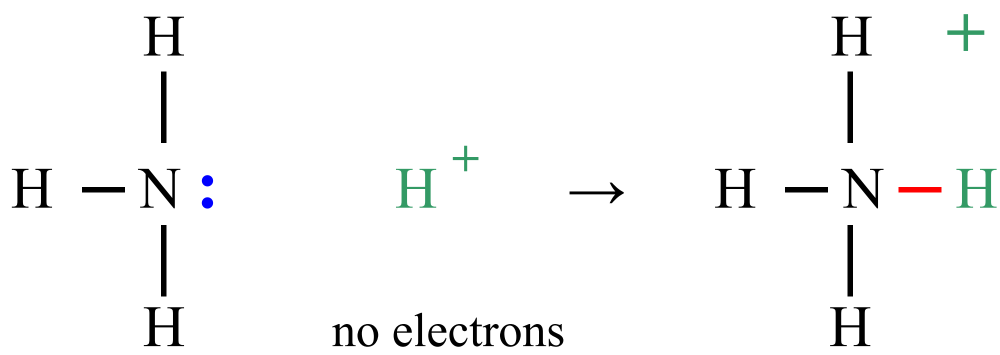
1 single
& __________

Ch 5.5 Coordinate Covalent Bonds

In a coordinate covalent bond ______ electrons of a shared pair come ____________ of the two atoms in the bond.



Ammonia shares its ________ with the hydrogen ion.



8 electrons

1 coordinate
covalent bond

$$NH_3 \quad + \quad H^+ \quad \rightarrow \quad NH_4^+$$

ammonia ________ ion

Ch 5.6 Systematic Procedures for Drawing Lewis Structures

Bottom Line

- All valence electrons must be shown.
- All atoms must have an octet of electrons.
  (common exceptions include H)

Strategy

- 6 steps (slightly modified here)

Example A

- Ammonia ________

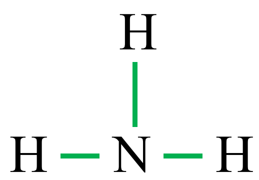
Step 1. Sum up all valence e (adjust for charge if necessary)

N in VA = 5 valence e
H in IA = 1 valence e

1 x 5e + 3 x 1e = ____ valence e (there is no charge)

Step 2. Draw skeleton structure and connect atoms by ________ covalent bonds.

NH_3 the central atom is often written first

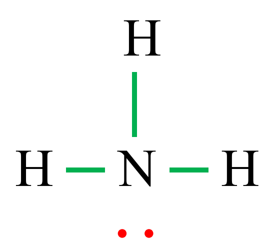


Step 3. Subtract the number of e used in skeletal structure bonds from valence e of step 1.

8 valence e from step 1
\- 6 e used in step 2 (2 e per bond)
2 e left

Step 4. Count number of e needed in skeletal structure (step 2) to give each atom an octet (2 e for H). If that number = e left in step 3, distribute them.

- all H are ______
- N needs 2 e for octet
- 2 e left in step 3
- add as lone pair on _____



Example B Hydrogen cyanide HCN (written in order bonded)

Step 1 total valence e = 1 + 4 + 5 = 10 e (no charge)

Step 2 H — C — N

Step 3 10 e available
- __ e used in step 2
6 e left

Step 4 H has 2 e = complete
C in step 2 needs 4 e
N in step 2 needs 6 e
need 10 e to complete all octets

We are short ____ e

Step 5 For each 2 e short, share one more pair by forming multiple bonds. Add 2 more bonds.

H — C ≡ N

Step 6 Distribute remaining e to complete octets.

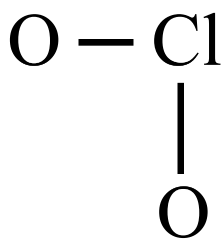
10 e available (step 1)
- __ e used (step 5)
2 e remain

H — C ≡ N :

Example C (drill problem) Chlorite ion ClO_2^-

valence e : $7 + 2(6) + 1 = 20$ e

$$\begin{array}{r} - \ 4 \text{ e used} \\ \hline 16 \text{ e left} \end{array}$$



complete this structure

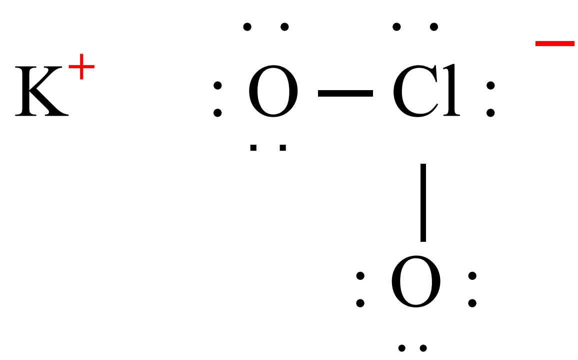
O needs ______

Cl needs ______

______ e needed

Ch 5.7 Bonding in Compounds with Polyatomic Ions Present

Both ionic and covalent bonds are present.



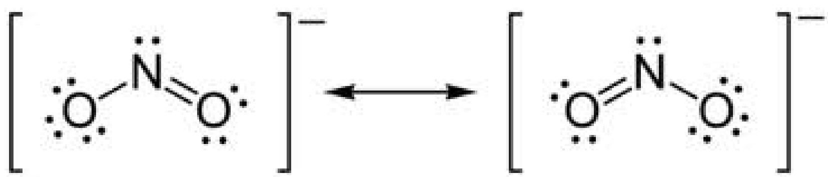
Supplemental Material

Resonance Structures

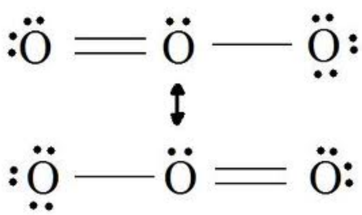
A single Lewis structure does not always adequately represent a substance and the concept of resonance is used to describe the bonding in such molecules.

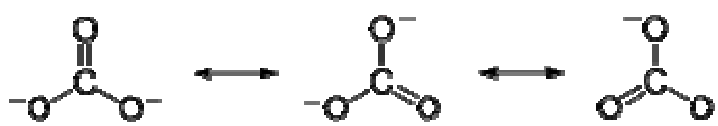
Resonance structures are two or more Lewis structures that represent the same ion or molecule equally well.

Examples:



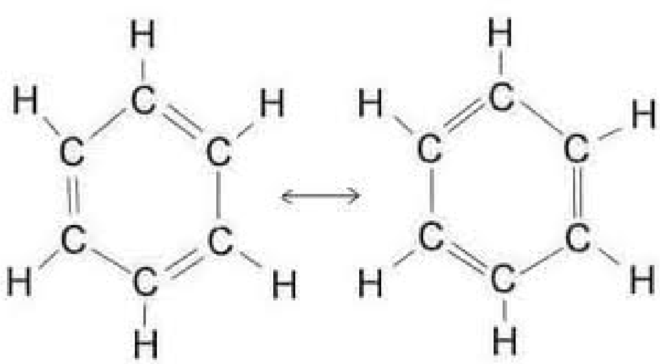
______ ion NO_2^-





ozone O_3

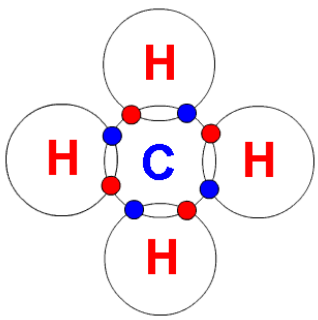
______ ion CO_3^{2-}

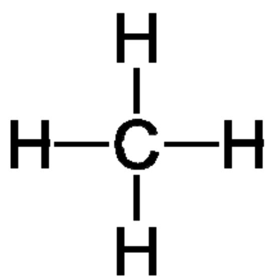


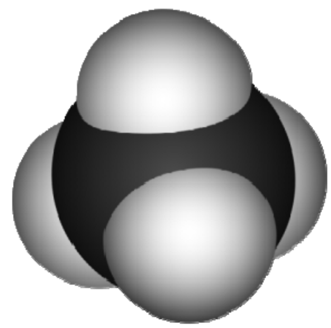
benzene

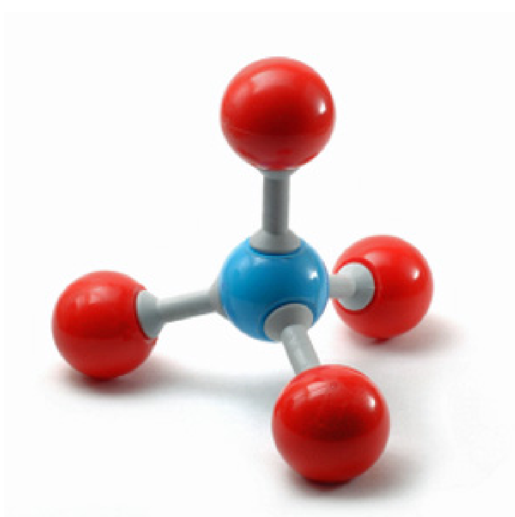
Ch 5.8 Molecular Geometry is the 3-D arrangement of atoms within a molecule

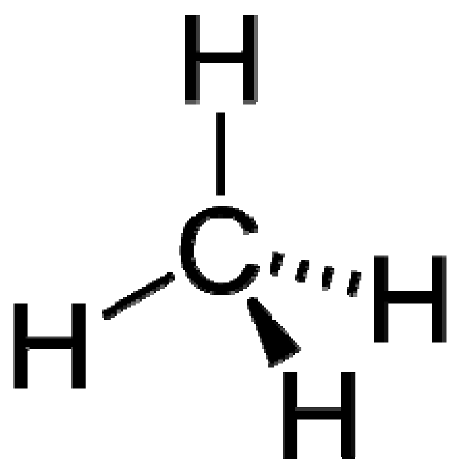
Molecular models of methane





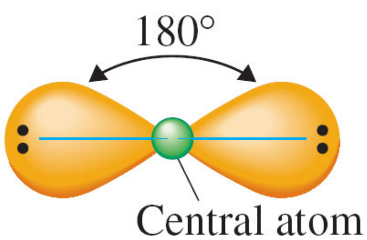






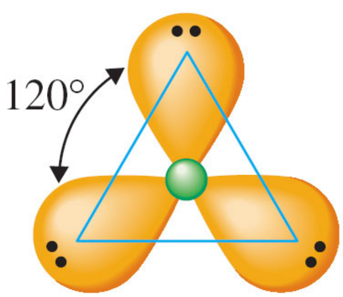
We can predict the molecular geometry using a molecule's Lewis structure and the Valence Shell Electron-Pair Repulsion theory (____________ theory)

- Electrons repel each other
- Electrons tend to be as far apart as possible
- Electron pairs control ____________



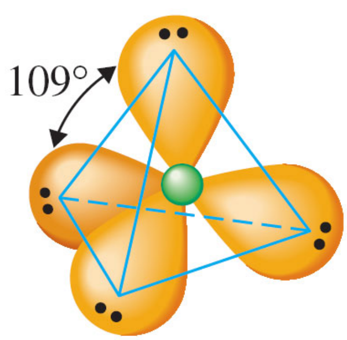
(a) Linear

2 electron pairs or VSEPR groups



(b) Trigonal planar

3 electron pairs or VSEPR groups



(c) Tetrahedral

4 electron pairs or VSEPR groups

Figure 5.8

Steps in predicting molecular geometry

- draw a Lewis structure
- determine the number of VSEPR groups
  bonding and nonbonding pairs count equally
  single, double, and triple bonds count equally as one group
- predict which arrangement of VSEPR groups minimizes repulsion

A. Molecules with two VSEPR groups are linear

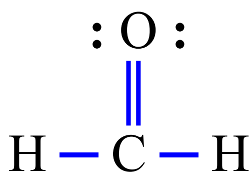
$$:\ddot{O}=C=\ddot{O}:$$

2 VSEPR groups around the central atom

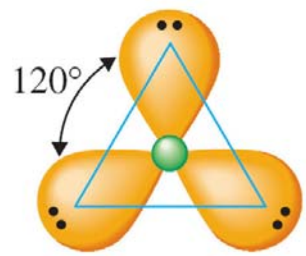
bond angle ______

B. Molecules with three VSEPR groups are either trigonal planar or angular

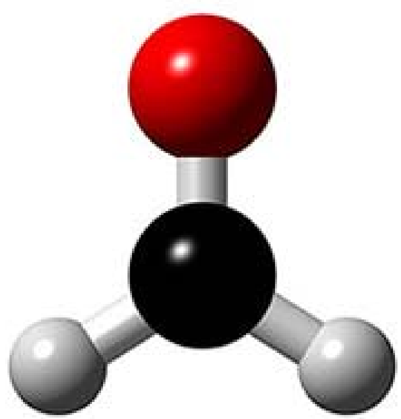
H_2CO (formaldehyde)



3 VSEPR groups



(b) Trigonal planar

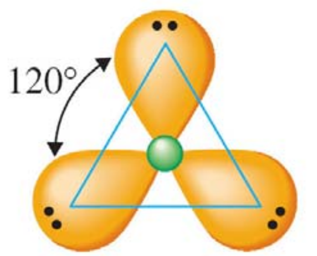


Trigonal planar

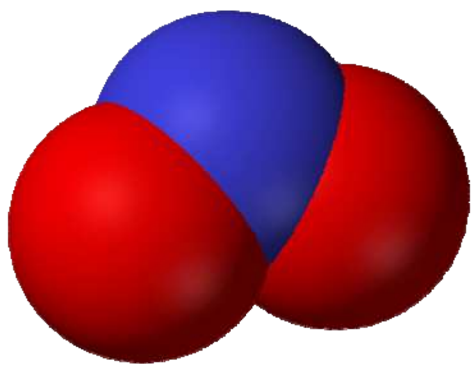
bond angle ~______

NO_2^- nitrite ion $\left[:\ddot{O}=\ddot{N}-\underset{\cdot\cdot}{\ddot{O}}:\right]^-$

_____ VSEPR groups around N
2 bonding, 1 nonbonding



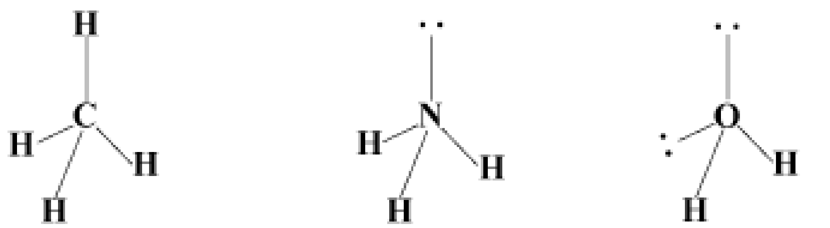
(b) Trigonal planar



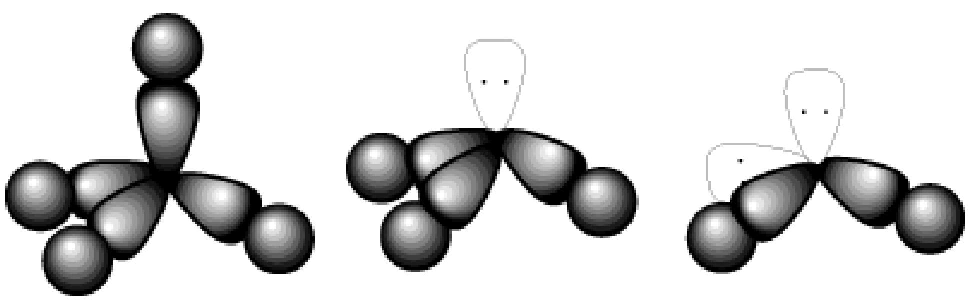
bent or angular shape
bond angle ~120°

C. Molecules with four VSEPR groups have three possible geometries:

- Tetrahedral (no nonbonding e pairs)
- Trigonal pyramidal (____ nonbonding e pair)
- Angular or bent (_____ nonbonding e pairs)



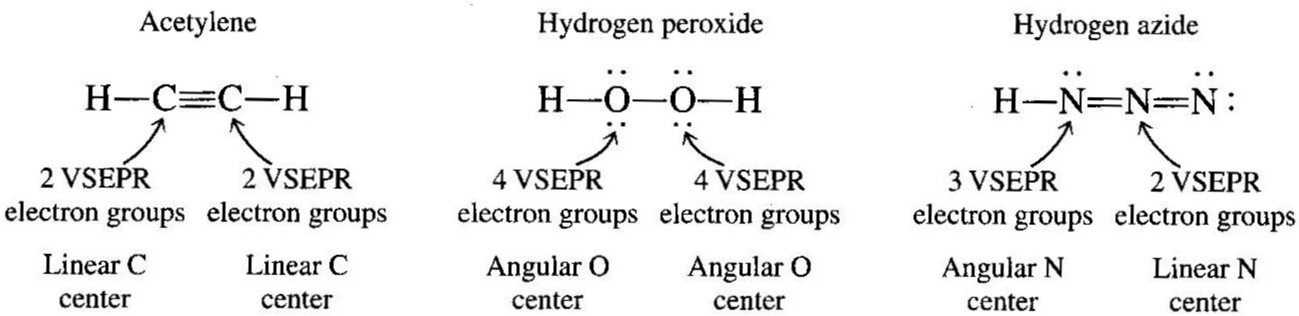
methane ammonia water

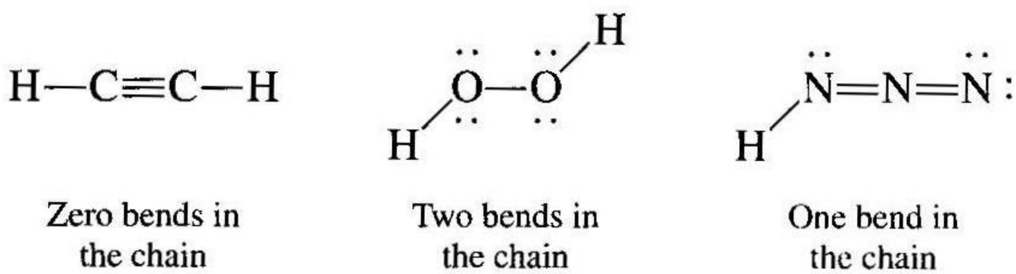


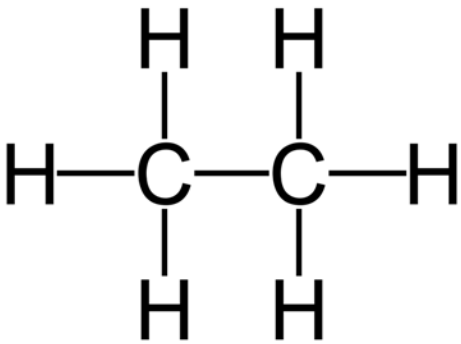
tetrahedral trigonal pyramidal bent

109.5^{o} 107^{o} 104.5^{o}

Molecules with more than one central atom



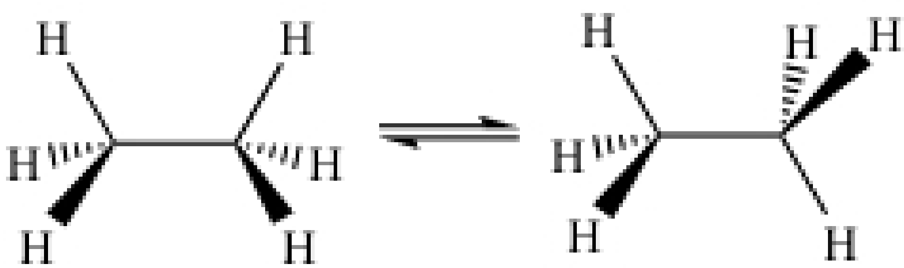


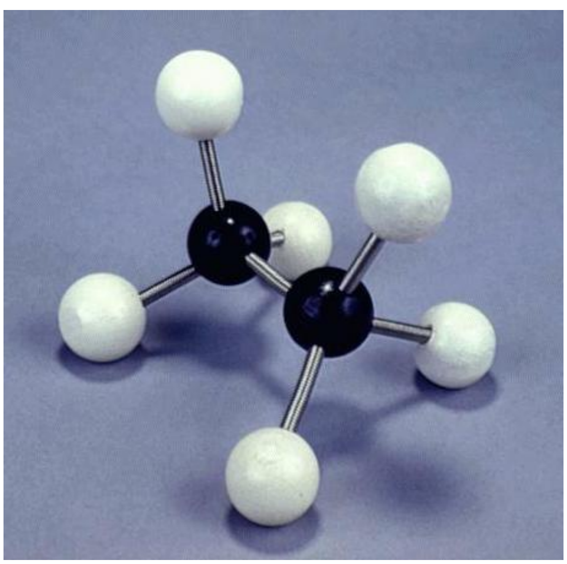


Ethane

4 VSEPR groups for ______ carbon atom

Tetrahedral arrangement around each carbon atom.

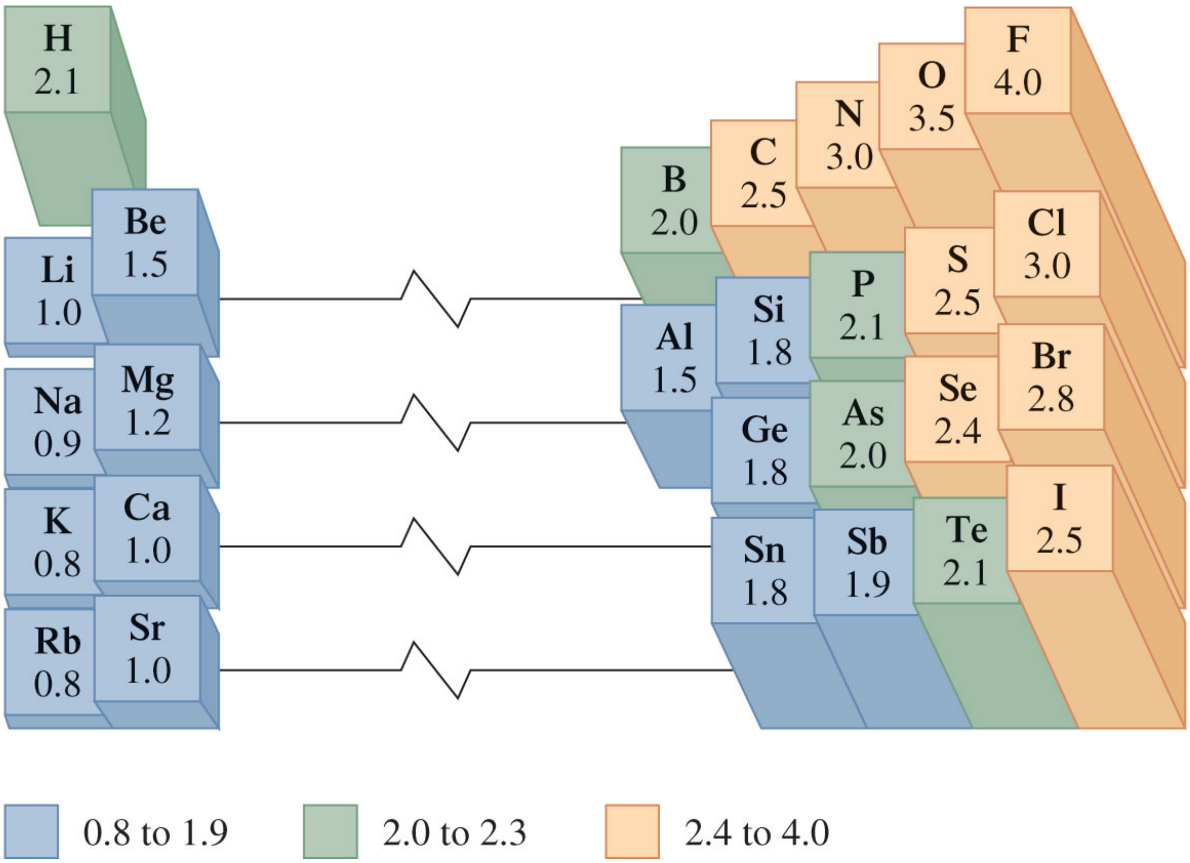




ball-and-stick model of

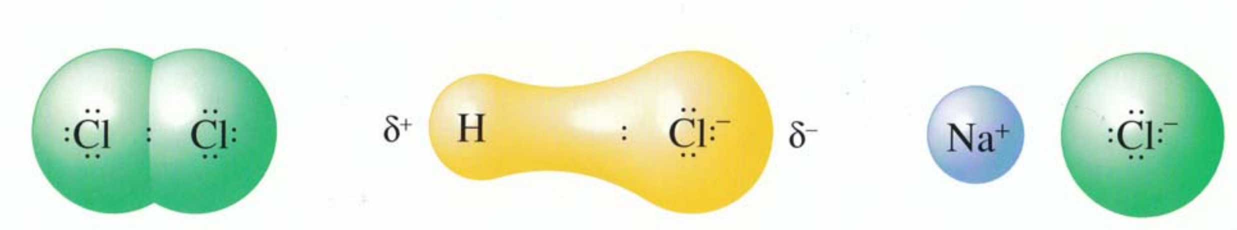
Ethane

Ch 5.9 Electronegativity



Electronegativity is a measure of the relative attraction that an atom has for the __________ electrons in a bond.

Ch 5.10 Bond Polarity



pure covalent polar covalent ionic

Differences in Electronegativity

__________ covalent bonds (similar values) ≤ 0.4

$0.4 <$ __________ covalent bonds < 1.5

$1.5 <$ "Borderline area" < 2.0 (ionic or polar covalent)

__________ bonds > 2.0

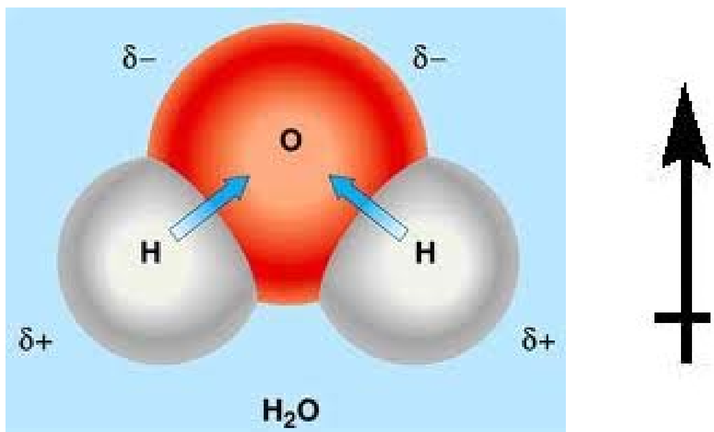
Depicting bond polarity

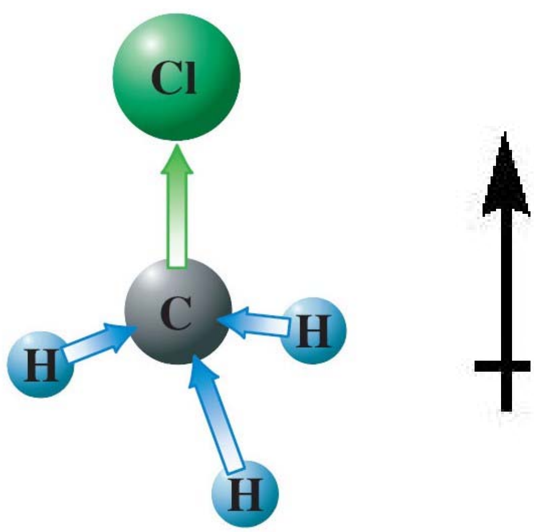
$\overset{\delta+}{\mathrm{H}} - \overset{\delta-}{\mathrm{F}}$ or $\overset{+\!\!\longrightarrow}{\mathrm{H - F}}$

Ch 5.11 Molecular Polarity depends on two factors:

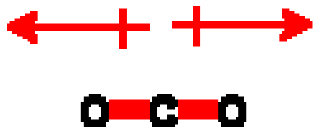
- Bond polarity
- Molecular ____________

A. Polar bond + unsymmetrical distribution of electronic charge = ____________ molecule

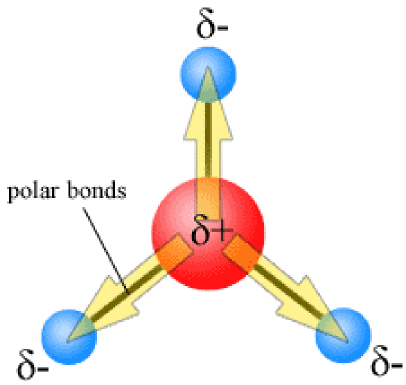




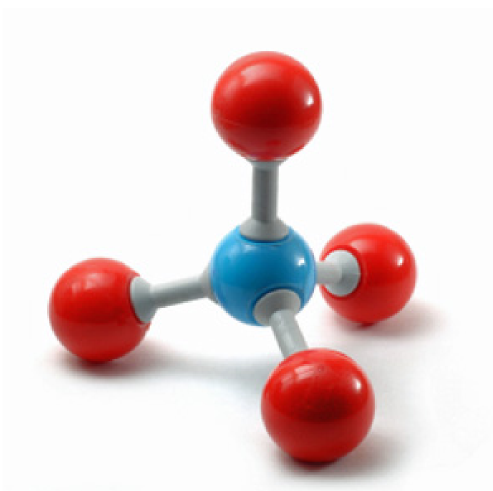
B. Polar bond + symmetrical distribution of electronic charge = ____________ molecule



Polarity cancels



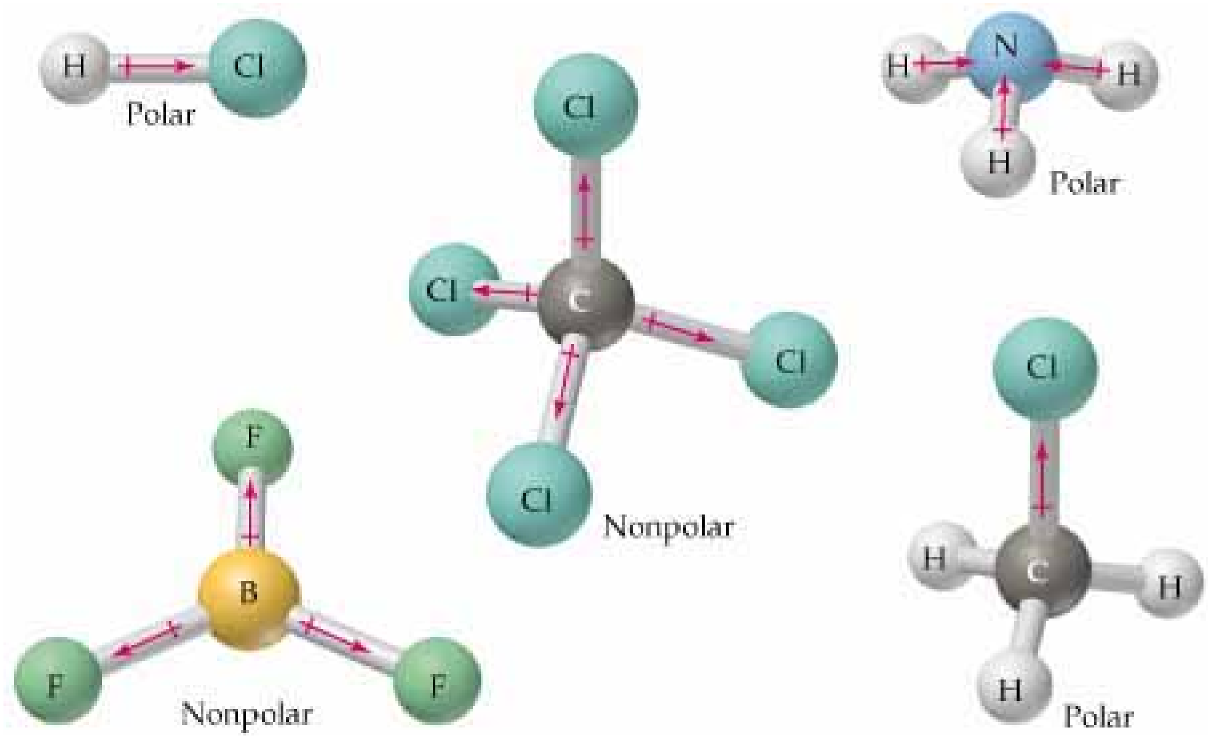
Polarity cancels in a trigonal planar molecule with 3 identical atoms or groups attached.



Polarity cancels in a tetrahedral molecule with 4 identical atoms or groups attached.

Examples of polar and nonpolar molecules
(all contain __________ bonds)

NH$_3$ is trigonal
pyramidal, not planar



Ch 5.12 Naming Binary Molecular Compounds

Different compounds exist for most pairs of ________________.
Examples of N-O compounds: NO NO_2 N_2O_3 N_2O_4 N_2O_5

$$N_2O_3$$

dinitrogen trioxide

1. prefix + full name of least electronegative nonmetal
2. prefix + stem name of more electronegative nonmetal + suffix of "ide"

Prefix	Number
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9
deca-	10

Table 5.1 Common numerical prefixes for 1 -10

Table 5.2 Accepted Common Names

Compound Formula	Accepted Common Name	Memorize?
H_2O	water	Yes
H_2O_2	hydrogen peroxide	Yes
NH_3	ammonia	Yes
N_2H_4	hydrazine	No
CH_4	methane	Yes
C_2H_6	ethane	
PH_3	phosphine	
AsH_3	arsine	