Chemistry 104 Chapter Five PowerPoint Notes

Chemical Bonding: The Covalent Bond Model Chapter 5

Chemistry 104

Professor Michael Russell

Covalent Bonds

A ***covalent bond*** is a bond formed by sharing electrons between atoms.

A ***molecule*** is a group of atoms held together by covalent bonds.

Nonmetals form covalent bonds with nonmetals. They reach the Noble Gas configuration by *sharing* an appropriate number of electrons.

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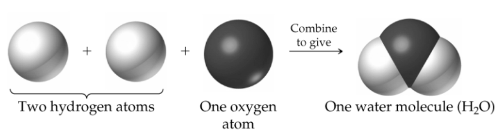
A water molecule results when two hydrogen atoms and one oxygen atom are covalently bonded:



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Test Yourself

Are these compounds bonded through ionic or covalent bonding?

PCl_5

Na_2O

SO_3

$CaSO_3$

SbAs

Nomenclature *of covalent compounds different from ionic compounds; important to know the difference*

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Naming Molecular Compounds

When two or more nonmetal elements combine they form *covalent compounds.*

The formulas of covalent compounds are written with the less electronegative (*i.e. more metal-like*) element first.

More electronegative element gets *-ide* suffix

Use Greek Prefixes to indicate number of atoms present.

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Greek Prefixes

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

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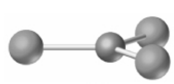
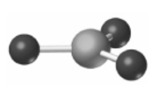
Covalent compounds and nomenclature:



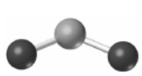
boron trichloride

BCl_3



SO_3

sulfur trioxide



NO_2

nitrogen dioxide

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Test Yourself - Covalent Bonding

Give the names for the following formulas:

N_2O_5

SO_2

OF_2

P_2O_3

NO

Give the formulas for the following names:

tetraphosphorus decaoxide

carbon dioxide

carbon monoxide

nitrogen dioxide

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Practice, practice, practice!

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When atoms come together, electrical interactions occur.
Some interactions are repulsive:

positively charged nuclei repel each other

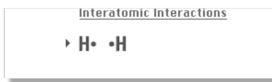
negatively charged electrons repel each other.

Other interactions are attractive:

nucleus A attracts electrons in atom B and vice-versa

If attractive forces > repulsive forces

covalent bond formed



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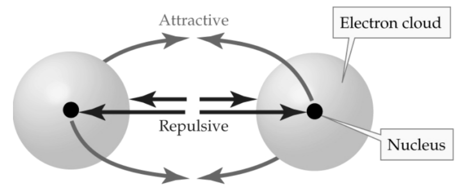
A covalent bond between two hydrogen atoms:



A covalent bond is the result of attractive and repulsive forces between atoms.

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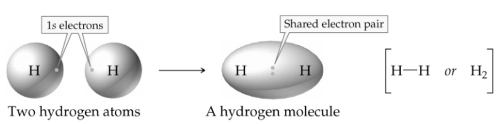
1s orbitals on two individual H atoms *bend together* and *overlap* to give egg shaped region in the *hydrogen molecule* - the covalent bond.

This shared pair of electrons in a covalent bond often represented as a line between atoms.



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A ***Bond length*** is the optimum distance between nuclei in a covalent bond.

If atoms too far apart:

attractive forces are small

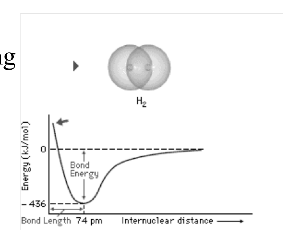
no bond forms

If atoms are too close:

repulsive forces strong

pushes atoms apart

no bond forms



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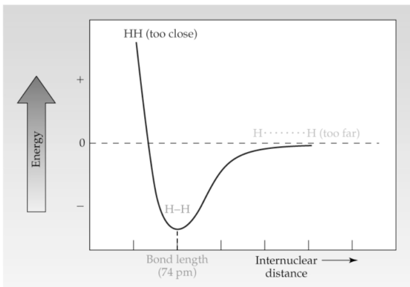

A graph of potential energy versus internuclear distance for hydrogen.

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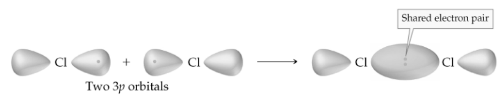
When two chlorine atoms approach each other, the unpaired 3*p* electrons are *shared* by both atoms in a covalent bond.

Each chlorine atom in the Cl_2 molecule has 7 electrons in its own valence shell, and sharing one more gives each valence shell an octet.



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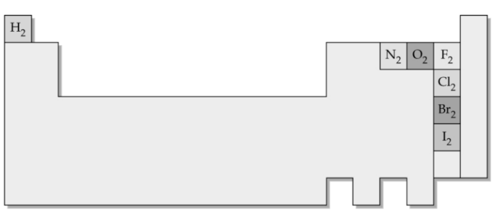
In addition to H_2 and Cl_2, five other elements *always* exist as diatomic molecule.



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Elements that Exist as Diatomic Molecules

Have

No

Fear

Of

Ice

Clear

Brew

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Elements that Exist as Diatomic Molecules

Have - *Hydrogen,* H_2

No - *Nitrogen,* N_2

Fear - *Fluorine,* F_2

Of - *Oxygen,* O_2

Ice - *Iodine,* I_2

Clear - *Chlorine,* Cl_2

Brew - *Bromine,* Br_2

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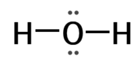
Covalent Bonds and the Periodic Table

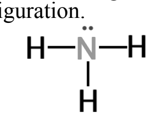
Covalent bonds can form between unlike atoms as well, making possible a vast number of molecular compounds.

Water, H_2O, consists of two hydrogen atoms joined by covalent bonds to one oxygen atom.

Ammonia, NH_3, consists of three hydrogen atoms joined by covalent bonds to one nitrogen atom.

In most covalent molecules, each atom shares enough electrons to achieve a noble gas configuration.





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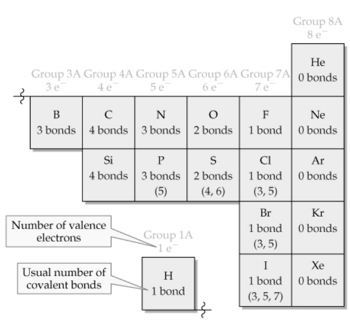

Fig 5.4 For P, S, Cl, and other elements in the third period and below, the number of covalent bonds may vary, as indicated by the numbers shown in parentheses.

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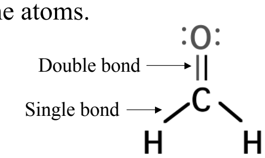
Multiple Covalent Bonds

Single bond: A bond formed by sharing two electrons or one pair – represented by a single line between the atoms.

Double bond: A bond formed by sharing four electrons or two pairs – represented by two lines (=) between the atoms.



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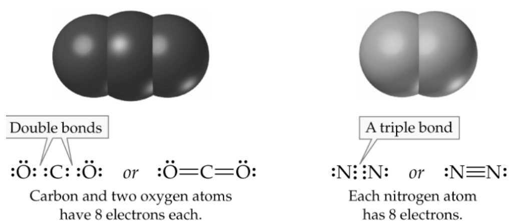
Triple bond: A bond formed by sharing six electrons or three pairs– represented by three lines (≡) between the atoms.



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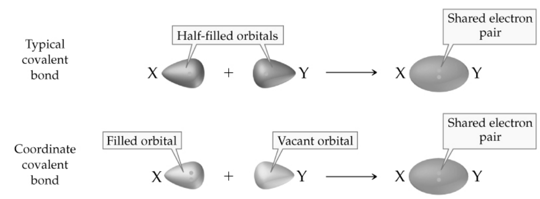
Coordinate Covalent Bonds

Coordinate Covalent Bond: The covalent bond that forms when both electrons are donated by the same atom.



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Molecular Formulas and Lewis Structures

Molecular Formula: A formula that shows the number and kind of atoms in a molecule

Structural formula: Molecular representation that shows the connections among atoms by using lines to represent covalent bonds

Example for water:

H_2O = molecular formula

H-O-H = structural formula

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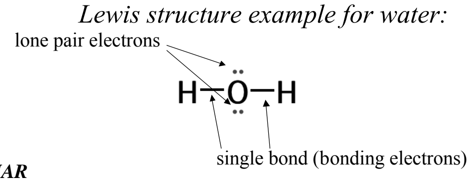
Lewis structure: Molecular representation showing both the connections among atoms *and* the locations of lone pair valence electrons.

A ***lone pair*** is a pair of electrons not used for bonding.

Lewis structure example for water:



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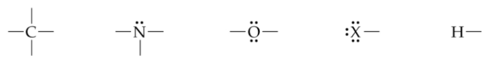
Drawing Lewis Structure

To draw a Lewis structure, you need to know the connections among atoms.

Knowing *common bonding patterns* simplifies writing Lewis structure.



Carbon 4 bonds | Nitrogen 3 bonds | Oxygen 2 bonds | Halogen 1 bond | Hydrogen 1 bond

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Building a Lewis Dot Structure

Ammonia, NH_3

1. Decide on central atom; never H.

Central atom is atom of lowest affinity for electrons (usually first atom listed)

Therefore, N is central

2. Count valence electrons

H = 1 and N = 5

Total = (3 x 1) + 5

= 8 electrons *or* 4 pairs

Pairs can be lone or bonding

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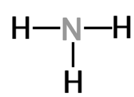
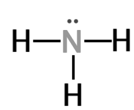
Building a Lewis Dot Structure

3. Form a covalent bond between the central atom and surrounding atoms.



4. Remaining electrons form LONE PAIRS to complete octet as needed.



3 BOND PAIRS and 1 LONE PAIR.

Note that N has an octet (8 electrons), while H shares only 2 electrons

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<u>Lewis structure rules</u>:

1: Decide on a central atom and find the total number of valence electrons in molecule or ion

2: Draw a line between each pair of connected atoms to represent a covalent bond

3: Add lone pairs so that each *peripheral* atom (except H) gets an octet

4: Place all remaining electrons on the central atom

5: If central atom does not have an octet, take lone pair(s) from neighboring atom(s) and form multiple bond(s) to the central atom

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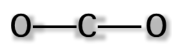
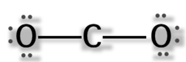
Carbon Dioxide, CO_2

1. Central atom = _______
2. Valence electrons = __ or __ pairs
3. Form single bonds.



This leaves 6 pairs.

4. Place lone pairs on outer atoms.



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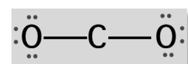
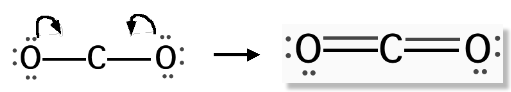
Carbon Dioxide, CO_2

4. Place lone pairs on outer atoms.



5. So that C has an octet, we shall form DOUBLE BONDS between C and O.



Practice, practice, practice!

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Shape of Molecules

Molecular shapes can be predicted by noting how many bonds and electron pairs surround individual atoms and applying ***valence-shell electron-pair repulsion*** *(VSEPR)* theory.

Basic idea of VSEPR: negatively charged electron clouds in bonds and lone pairs repel each other, keeping them as far apart as possible

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VSEPR

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VSEPR Rules

To apply VSEPR theory:

1: Draw the Lewis structure of the molecule and identify the central atom

2: Count the number of electron charge clouds (lone *and* bonding pairs) surrounding the central atom.

3: Predict molecular shape by assuming that clouds orient so they are as far away from one another as possible.

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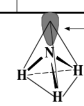
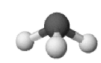
VSEPR Shape Predictor Table - Table 5.1

Clouds	Bonds	Lone Pairs	Electron Pair Geom.	Molecular Shape	Angles	Notes
2	2	0	linear	linear	180	
3	3	0	planar triangular	planar triangular	120	
3	2	1	planar triangular	bent	120	*always polar*
4	4	0	tetrahedral	tetrahedral	109	
4	3	1	tetrahedral	pyramidal	109	*always polar*
4	2	2	tetrahedral	bent	109	*always polar*



tetrahedral EPG, pyramidal Molecular shape



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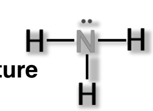
Structure Determination by VSEPR

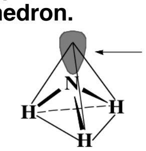
Ammonia, NH_3



1. Draw Lewis electron dot structure

2. Count Electron charge clouds (bonds and lone pairs) = 4

3. The 4 electron pairs are at the corners of a tetrahedron.



lone pair of electrons in tetrahedral position

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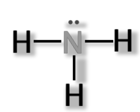
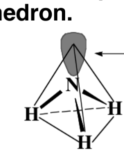
Structure Determination by VSEPR

Ammonia, NH_3

There are 4 electron pairs at the corners of a tetrahedron.





lone pair of electrons in tetrahedral position

The ELECTRON PAIR GEOMETRY is tetrahedral.

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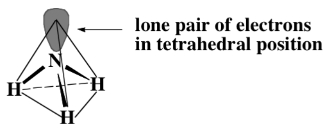
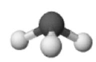
Structure Determination by VSEPR

Ammonia, NH_3

The electron pair geometry is tetrahedral.





The MOLECULAR GEOMETRY — the positions of the atoms — is PYRAMIDAL.

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Test Yourself

Describe the Lewis structure, electron pair geometry and molecular shape of methane, CH_4.

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Polar Covalent Bonds and Electronegativity

Electrons in a covalent bond occupy the region between the bonded atoms.

If atoms in bond identical (H_2, Cl_2, etc.) electrons are attracted equally to both atoms and are shared equally.

If atoms in bond different (HCl, HF, etc.) electrons may be attracted more strongly by one atom than by the other and are shared unequally.

Such bonds are known as ***polar covalent bonds.***

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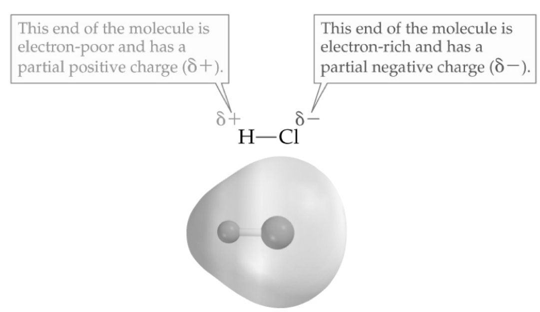

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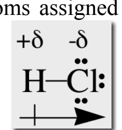
In HCl, electrons spend more time near Cl than H. Although molecule is overall neutral, the chlorine is more negative than the hydrogen, resulting in partial charges on the atoms.

Partial charges represented by placing δ- on the more negative atom and δ+ on the more positive atom.

Ability of an atom to attract electrons is called the atom's ***electronegativity***.

Fluorine, the most electronegative element, assigned a value of 4, and less electronegative atoms assigned lower values



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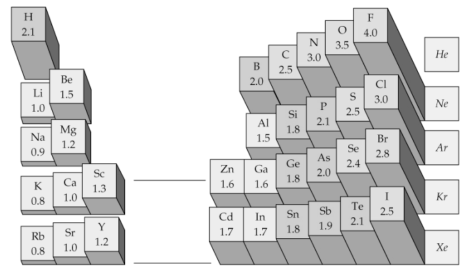

Electronegativities and the periodic table

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Polar Molecules

Entire molecule can be polar *if* electrons are attracted more strongly to one part of the molecule than another.

Molecule's polarity is due to the sum of all individual bond polarities *and* lone-pair contributions in the molecule.

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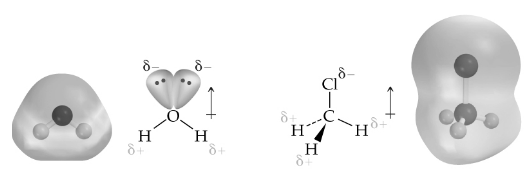
Molecular polarity is represented by an arrow pointing at the negative end and is crossed at the positive end to resemble a positive sign.



Water, H_2O Chloromethane, CH_3Cl

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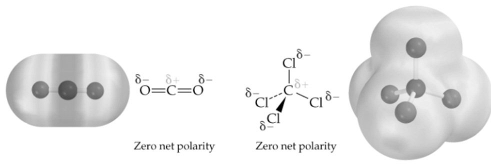
Molecular polarity depends on the shape of the molecule as well as the presence of polar covalent bonds and lone-pairs



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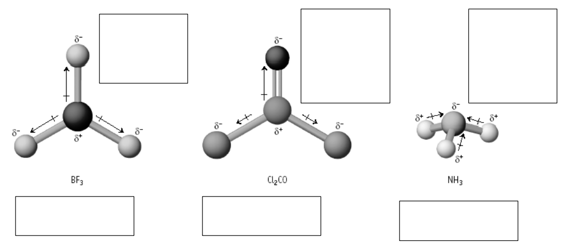
Test Yourself

Are BF_3, Cl_2CO, and NH_3 polar or nonpolar?



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End of Chapter 5

To review and study for Chapter 5, look at the "Concepts to Remember" at the end of Chapter Five