Chapter 10
Chemical Bonding

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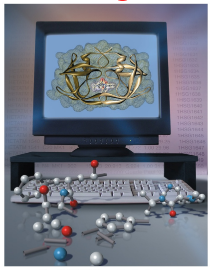
Chem 118
Introductory Chemistry



Map: Introductory Chemistry (Tro) https://chem.libretexts.org/@go/page/45050 (accessed Mar 25, 2022).

Bonding Theories

- bonding is the way atoms attach to make molecules
- an understanding of how and why atoms attach together in the manner they do is central to chemistry
- chemists have an understanding of bonding that allows them to:

1) predict the shapes of molecules and properties of substances based on the bonding within the molecules
2) design and build molecules with particular sets of chemical and physical properties

Lewis Symbols of Atoms

- also known as electron dot symbols
- use symbol of element to represent nucleus and inner electrons
- use dots around the symbol to represent valence electrons
  - ✓ put one electron on each side first, then pair
- remember that elements in the same group have the same number of valence electrons; therefore their Lewis dot symbols will look alike

Li• Be• •B• •C• •N• •O: :F: :Ne:

Lewis Bonding Theory

- atoms bond because it results in a more stable electron configuration
- atoms bond together by either transferring or sharing electrons so that all atoms obtain an outer shell with 8 electrons
  - ✓ **Octet Rule**
  - ✓ there are some exceptions to this rule – the key to remember is to try to get an electron configuration like a noble gas

Lewis Symbols of Ions

- Cations have Lewis symbols without valence electrons
  - ✓ Lost in the cation formation
- Anions have Lewis symbols with 8 valence electrons
  - ✓ Electrons gained in the formation of the anion

Li• Li^{+1} :F: $[:F:]^{-1}$

Ionic Bonds

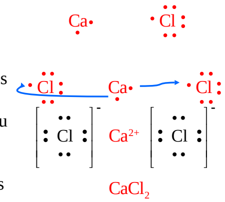
- metal to nonmetal
- metal loses electrons to form cation
- nonmetal gains electrons to form anion
- ionic bond results from + to - attraction
  - ✓ larger charge = stronger attraction
  - ✓ smaller ion = stronger attraction
- Lewis Theory allow us to predict the correct formulas of ionic compounds

Example 10.3 - Using Lewis Theory to Predict Chemical Formulas of Ionic Compounds

Predict the formula of the compound that forms between calcium and chlorine.

Draw the Lewis dot symbols of the elements

Transfer all the valance electrons from the metal to the nonmetal, adding more of each atom as you go, until all electrons are lost from the metal atoms and all nonmetal atoms have 8 electrons

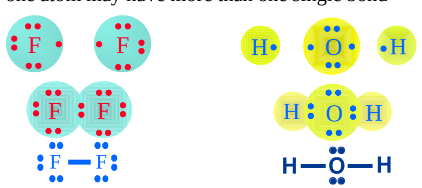


Covalent Bonds

- often found between two nonmetals
- typical of molecular species
- atoms bonded together to form molecules
  - ✓ strong attraction
- **sharing pairs of electrons** to attain octets
- molecules generally weakly attracted to each other
  - ✓ observed physical properties of molecular substance due to these attractions

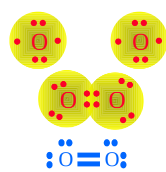
Single Covalent Bonds

- two atoms share one pair of electrons
  - ✓ 2 electrons
- one atom may have more than one single bond



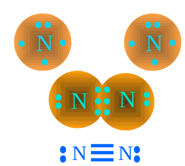
Double Covalent Bond

- two atoms sharing two pairs of electrons
  - ✓ 4 electrons
- shorter and stronger than single bond



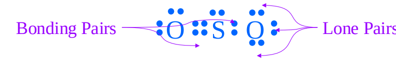
Triple Covalent Bond

- two atoms sharing 3 pairs of electrons
  - ✓ 6 electrons
- shorter and stronger than single or double bond



Bonding & Lone Pair Electrons

- Electrons that are shared by atoms are called **bonding pairs**
- Electrons that are not shared by atoms but belong to a particular atom are called **lone pairs**
  - ✓ also known as **nonbonding pairs**



Polyatomic Ions

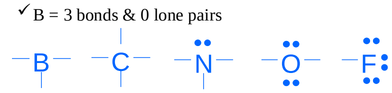
- The polyatomic ions are attracted to opposite ions by ionic bonds
  - ✓ Form crystal lattices
- Atoms in the polyatomic ion are held together by covalent bonds

Lewis Formulas of Molecules

- shows pattern of valence electron distribution in the molecule
- useful for understanding the bonding in many compounds
- allows us to predict shapes of molecules
- allows us to predict properties of molecules and how they will interact together

Lewis Structures

- some common bonding patterns
  - ✓ C = 4 bonds & 0 lone pairs
    - ➢ 4 bonds = 4 single, or 2 double, or single + triple, or 2 single + double
  - ✓ N = 3 bonds & 1 lone pair,
  - ✓ O = 2 bonds & 2 lone pairs,
  - ✓ H and halogen = 1 bond,
  - ✓ Be = 2 bonds & 0 lone pairs,
  - ✓ B = 3 bonds & 0 lone pairs



Writing Lewis Structures for Covalent Molecules

1) Attach the atoms together in a skeletal structure
   - ✓ most metallic element generally central
   - ✓ halogens and hydrogen are generally terminal
   - ✓ many molecules tend to be symmetrical
   - ✓ in oxyacids, the acid hydrogens are attached to an oxygen
2) Calculate the total number of valence electrons available for bonding
   - ✓ use group number of periodic table

Writing Lewis Structures for Covalent Molecules

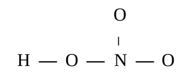
3) Attach atoms with pairs of electrons and subtract electrons used from total
   - ✓ bonding electrons
4) Add remaining electrons in pairs to complete the octets of all the atoms
   - ✓ remember H only wants 2 electrons
   - ✓ don't forget to keep subtracting from the total
   - ✓ complete octets on the terminal atoms first, then work toward central atoms

Writing Lewis Structures for Covalent Molecules

5) If there are not enough electrons to complete the octet of the central atom, bring pairs of electrons from an attached atom in to share with the central atom until it has an octet
   - ✓ try to follow common bonding patterns

Example HNO_3

1) Write skeletal structure
   - ✓ since this is an oxyacid, H on outside attached to one of the O's; N is central



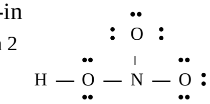
2) Count Valence Electrons and Subtract Bonding Electrons from Total

N = 5
H = 1
$O_3 = 3 \cdot 6 = 18$
Total = 24 e^-

	Electrons
Start	24
Used	8
Left	16

Example HNO_3

3) Complete Octets, outside-in
   - ✓ H is already complete with 2
     - ➢ 1 bond



4) Re-Count Electrons

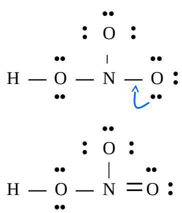
N = 5
H = 1
$O_3 = 3 \cdot 6 = 18$
Total = 24 e^-

	Electrons
Start	24
Used	8
Left	16

	Electrons
Start	16
Used	16
Left	0

Example HNO_3

5) If central atom does not have octet, bring in electron pairs from outside atoms to share
   - ✓ follow common bonding patterns if possible



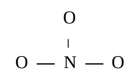
Other Examples: CO_2, H_2O, …

Writing Lewis Structures for Polyatomic Ions

- the procedure is the same, the only difference is in counting the valence electrons
- for polyatomic cations, take away one electron from the total for each positive charge
- for polyatomic anions, add one electron to the total for each negative charge

Example NO_3^-

1) Write skeletal structure
   - ✓ N is central because it is the most metallic



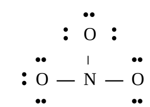
2) Count Valence Electrons and Subtract Bonding Electrons from Total

N = 5
$O_3 = 3 \cdot 6 = 18$
(-) = 1
Total = 24 e^-

	Electrons
Start	24
Used	6
Left	18

Example NO_3^-

3) Complete Octets, outside-in



4) Re-Count Electrons

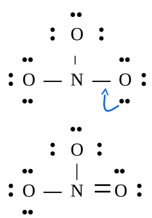
N = 5
$O_3 = 3 \cdot 6 = 18$
(-) = 1
Total = 24 e^-

	Electrons
Start	24
Used	6
Left	18

	Electrons
Start	18
Used	18
Left	0

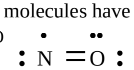
Example NO_3^-

5) If central atom does not have octet, bring in electron pairs from outside atoms to share
   - ✓ follow common bonding patterns if possible



Exceptions to the Octet Rule

- H & Li, lose one electron to form cation
  - ✓ Li now has electron configuration like He
  - ✓ H can also share or gain one electron to have configuration like He
- Be shares 2 electrons to form two single bonds
- B shares 3 electrons to form three single bonds
- expanded octets for elements in Period 3 or below
  - ✓ using empty valence *d* orbitals
- some molecules have odd numbers of electrons
  - ✓ NO

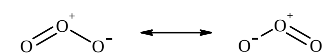


Resonance

- we can often draw more than one valid Lewis structure for a molecule or ion
- in other words, no one Lewis structure can adequately describe the actual structure of the molecule
- the actual molecule will have some characteristics of all the valid Lewis structures we can draw

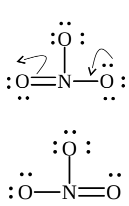
Resonance

- Lewis structures often do not accurately represent the electron distribution in a molecule
  - ✓ Lewis structures imply that O_3 has a single (147 pm) and double (121 pm) bond, but actual bond length is between, (128 pm)
- Real molecule is a *hybrid* of all possible Lewis structures
- Resonance stabilizes the molecule
  - ✓ maximum stabilization comes when resonance forms contribute equally to the hybrid



Drawing Resonance Structures

1. draw first Lewis structure that maximizes octets
2. move electron pairs from outside atoms to share with central atoms
3. if central atom 2nd row, only move in electrons if you can move out electron pairs from multiple bond

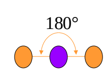


Molecular Geometry

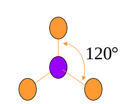
- Molecules are 3-dimensional objects
- We often describe the shape of a molecule with terms that relate to geometric figures
- These geometric figures have characteristic "corners" that indicate the positions of the surrounding atoms with the central atom in the center of the figure
- The geometric figures also have characteristic angles that we call **bond angles**

Some Geometric Figures

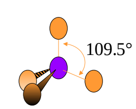
- **Linear**
  - ✓ 2 atoms on opposite sides of central atom
  - ✓ 180° bond angles



- **Trigonal Planar**
  - ✓ 3 atoms form a triangle around the central atom
  - ✓ Planar
  - ✓ 120° bond angles



- **Tetrahedral**
  - ✓ 4 surrounding atoms form a tetrahedron around the central atom
  - ✓ 109.5° bond angles



Predicting Molecular Geometry

- VSEPR Theory
  - ✓ Valence Shell Electron Pair Repulsion
- The shape around the central atom(s) can be predicted by assuming that the areas of electrons on the central atom will try to get as far from each other as possible
  - ✓ areas of negative charge will repel each other

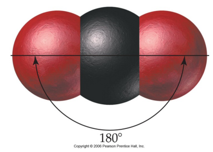
Areas of Electrons

- Each Bond counts as 1 area of electrons
  - ✓single, double or triple all count as 1 area
- Each Lone Pair counts as 1 area of electrons
  - ✓Even though lone pairs are not attached to other atoms, they do "occupy space" around the central atom
- Lone pairs take up slightly more space than bonding pairs
  - ✓Effects bond angles

Linear Shapes

- Linear
  - ✓2 areas of electrons around the central atom, both bonding
    - ➢Or two atom molecule as trivial case
  - ✓180° Bond Angles

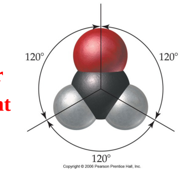
:Ö = C = Ö:



Trigonal Shapes

- Trigonal
  - ✓3 areas of electrons around the central atom
  - ✓120° bond angles
  - ✓All Bonding = **trigonal planar**
  - ✓2 Bonding + 1 Lone Pair = **bent**

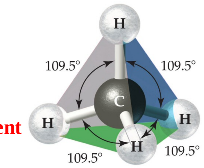
H — C = Ö:
|
H



Tetrahedral Shapes

- Tetrahedral
  - ✓4 areas of electrons around the central atom
  - ✓109.5° bond angles
  - ✓All Bonding = **tetrahedral**
  - ✓3 Bonding + 1 Lone Pair = **trigonal pyramid**
  - ✓2 Bonding + 2 Lone Pair = **bent**

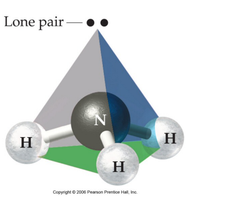
H
|
H — C — H
|
H

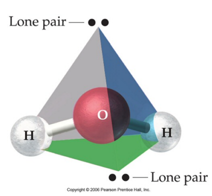


Tetrahedral Derivatives

H — N̈ — H
|
H

H — Ö — H

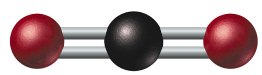




Molecular Geometry: Linear

- Electron Groups Around Central Atom = 2
- Bonding Groups = 2
- Lone Pairs = 0
- Electron Geometry = Linear
- Angle between Electron Groups = 180°

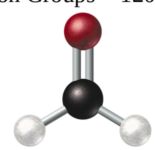
:Ö=C=Ö:



Molecular Geometry: Trigonal Planar

- Electron Groups Around Central Atom = 3
- Bonding Groups = 3
- Lone Pairs = 0
- Electron Geometry = Trigonal Planar
- Angle between Electron Groups = 120°

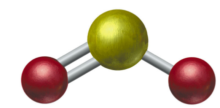
Ö:
||
H—C—H



Molecular Geometry: Bent

- Electron Groups Around Central Atom = 3
- Bonding Groups = 2
- Lone Pairs = 1
- Electron Geometry = Trigonal Planar
- Angle between Electron Groups = 120°

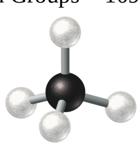
:Ö=S̈—Ö:



Molecular Geometry: Tetrahedral

- Electron Groups Around Central Atom = 4
- Bonding Groups = 4
- Lone Pairs = 0
- Electron Geometry = Tetrahedral
- Angle between Electron Groups = 109.5°

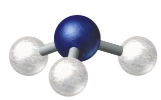
H
|
H—C—H
|
H



Molecular Geometry: Trigonal Pyramid

- Electron Groups Around Central Atom = 4
- Bonding Groups = 3
- Lone Pairs = 1
- Electron Geometry = Tetrahedral
- Angle between Electron Groups = 109.5°

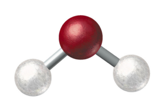
H—N̈—H
|
H



Molecular Geometry: Bent

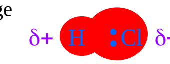
- Electron Groups Around Central Atom = 4
- Bonding Groups = 2
- Lone Pairs = 2
- Electron Geometry = Tetrahedral
- Angle between Electron Groups = 109.5°

H—Ö—H



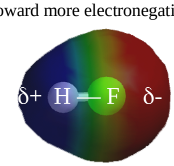
Bond Polarity

- bonding between unlike atoms results in unequal sharing of the electrons
  - ✓one atom pulls the electrons in the bond closer to its side
  - ✓one end of the bond has larger electron density than the other
- the result is **bond polarity**
  - ✓the end with the larger electron density gets a partial negative charge and the end that is electron deficient gets a partial positive charge



Electronegativity

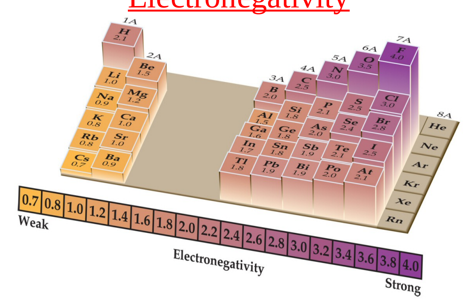
- measure of the pull an atom has on bonding electrons
- increases across period (left to right)
- decreases down group (top to bottom)
- larger difference in electronegativity, more polar the bond
  - ✓negative end toward more electronegative atom



Electronegativity

2.1																	
1.0	1.5											2.0	2.5	3.0	3.5	4.0	
0.9	1.2											1.5	1.8	2.1	2.5	3.0	
0.8	1.0	1.3	1.5	1.6	1.6	1.5	1.8	1.8	1.8	1.9	1.6	1.6	1.8	2.0	2.4	2.8	
0.8	1.0	1.2	1.4	1.6	1.8	1.9	2.2	2.2	2.2	1.9	1.7	1.7	1.8	1.9	2.1	2.5	
0.7	0.9	1.1	1.3	1.5	1.7	1.9	2.2	2.2	2.2	2.4	1.9	1.8	1.8	1.9	2.0	2.2	
0.7	0.9	1.1															

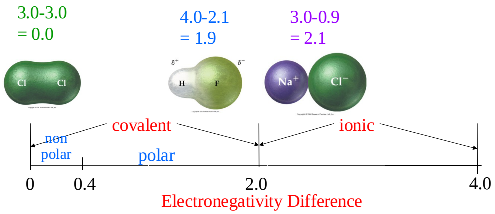
Electronegativity



Electronegativity & Bond Polarity

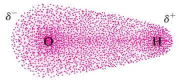
- If difference in electronegativity between bonded atoms is 0, the bond is **pure covalent**
  - ✓equal sharing
- If difference in electronegativity between bonded atoms is 0.1 to 0.3, the bond is **nonpolar covalent**
- If difference in electronegativity between bonded atoms 0.4 to 1.9, the bond is **polar covalent**
- If difference in electronegativity between bonded atoms larger than or equal to 2.0, the bond is **ionic**

Bond Polarity



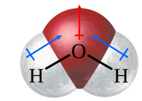
Dipole Moments

- a dipole is a material with positively and negatively charged ends
- polar bonds or molecules have one end slightly positive, δ+; and the other slightly negative, δ-
  - ✓not "full" charges, come from nonsymmetrical electron distribution
- Dipole Moment, μ, is a measure of the size of the polarity
  - ✓measured in Debyes, D

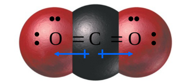


Polarity of Molecules

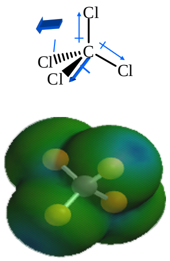
- in order for a molecule to be polar it must
  1) have polar bonds
     - electronegativity difference - theory
     - bond dipole moments - measured
  2) have an unsymmetrical shape
     - vector addition
- polarity effects the intermolecular forces of attraction



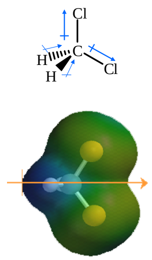
polar bonds,
and unsymmetrical
shape causes molecule
to be polar



polar bonds,
but nonpolar molecule
because pulls cancel

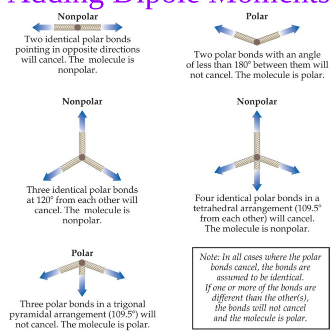


CCl_4
μ = 0.0 D



CH_2Cl_2
μ = 2.0 D

Adding Dipole Moments

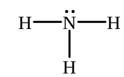


Example 10.11:
Determining if a Molecule is Polar

Example:
Determine if NH_3 is Polar.

Information
Given: NH_3
Find: if Polar
SM: formula → Lewis → Polarity & Shape → Molecule Polarity

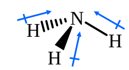
- Check.



N = 5
H = 3 · 1
total NH_3 = 8

bonding = 3 · 2 e^-
lone pairs = 1 · 2 e^-
total NH_3 = 8 e^-

bonds = polar
shape = trigonal pyramid



molecule = polar

The Lewis structure is correct. The bonds are polar and the shape is unsymmetrical, so it should be polar.