

# Chemistry Unit 5

# Types of Chemical Bonds

- A chemical bond forms when 2 or more atoms rearrange valence electrons to increase stability.
- There are three types of bonds, ionic, covalent, and metallic bonds.

# Ionic Bonds

- Ionic Bonds form when valence electrons are transferred from one atom to another
  - Cation- atom loses electrons and become +
  - Anion- atom gains electrons and become –
- In ionic compounds the ions are arranged in a crystal lattice with strong forces holding the ions together.

# Ionic Bond Properties

- High melting and boiling points
- Hard- not easily crushed
- Conduct electricity when melted or dissolved because the ions are free to move.

# Covalent Bonds

- Covalent bond forms when electrons are shared, forming molecules or compounds.
- Covalent bonds have weaker forces holding the molecules together.

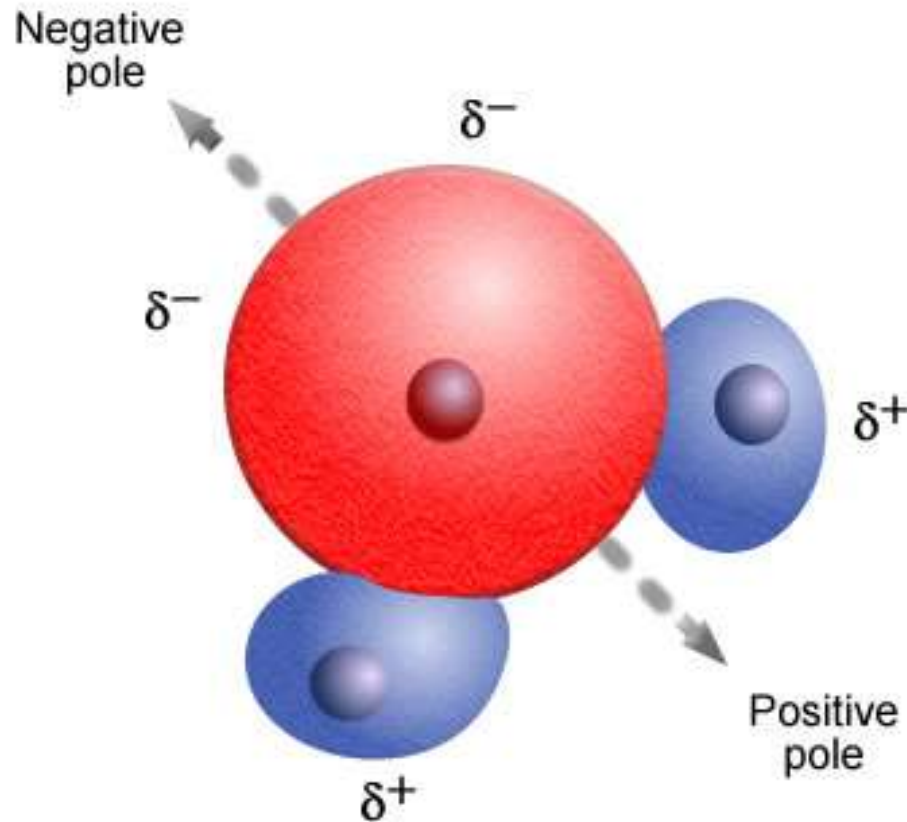
# Covalent Bond Properties

- Lower melting and boiling points
- Many covalent compounds are volatile (evaporate easily) liquids or gases.
- Softer- easier to crush
- Are not conductors of electricity.

# Electronegativity

- Electronegativity is a property that tells how strong an atom's attraction is for its electrons.
- Since oxygen has a higher electronegativity than hydrogen, oxygen holds onto shared electrons more, giving the oxygen a slightly negative charge and the hydrogen a partial positive charge.

# Water Molecule



Water molecule

# Two types of Covalent Bonds

- Polar covalent bonds form when electrons are shared unequally, creating partially charged ends or poles.
- Nonpolar covalent bonds form when electrons are shared equally because atoms have the same electronegativities.

# Electronegativity Chart

Electronegativity difference	Type of Bond
Greater than or equal to 1.7	Ionic Bond
Between 1.7 and 0.3	Polar Covalent Bond
Less than or equal to 0.3	Nonpolar Covalent Bond

# Practice

- Mg and F?

- S and O

# More Practice

- Li and Cl?
- C and O?
- Na and Cl?
- Cl and Br?
- S and H?

# Metallic Bond

- Metallic bonds are formed when electrons are delocalized or creates a sea of electrons between two metals.
- Properties of Metals
  - Conduct Electricity
  - Malleable
  - Ductile
  - Luster

# Oxygen

- Symbol = O
- Atomic number = 8
- Protons = 8
- Electrons = 8
- Electron distribution =  $1s^2 2s^2 2p^4$
- Valence electrons = 6

# Electron Dot Diagrams

- Electron Dot Diagrams are atom's symbol surrounded by dots to represent its valence electrons
- Example
  - OLi

# More Practice

● Mg

● He

● Kr

● Al

● Br

● K

● N

● Si

# Lewis Structure

- Lewis Structure- diagram representing the arrangement of valence electrons in a molecule.
- Most atoms need 8 valence electrons to become stable. The exceptions are H and He which only need 2 valence electrons to be stable.

# Lewis Structure for H<sub>2</sub>

- H-H
- Shared pair of electrons is where 2 electrons belonging to both atoms are represented by a line between the symbols.

# Lewis Structure for Cl<sub>2</sub>

- Each Cl atom has 7 electrons, giving a total of 14 valence electrons to work with.
- Cl—Cl
- Have one shared pair
- An Unshared pair are electrons belonging to only one atom and are represented by 2 dots.

# Lewis Structure for HCl

- H—Cl
- When more than two atoms bond, you must determine which is central.
  - The central atom is often the atom with the smallest electronegativity
  - Most commonly C
  - And Never H

# Lewis Structure for CH<sub>3</sub>I

- Total number of electrons is \_\_\_\_\_.
- What is the central atom?

# Episode Unit 5problem set

# Lewis Structure for C<sub>2</sub>H<sub>4</sub>

- Total number of valence electrons

# Type of bond Chart

Type of Bond	Pairs of electrons shared
Single	1
Double	2
Triple	3

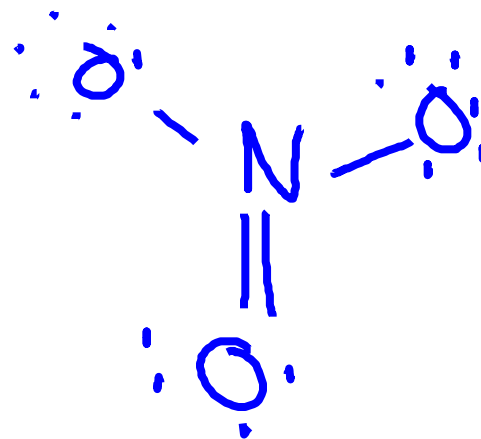
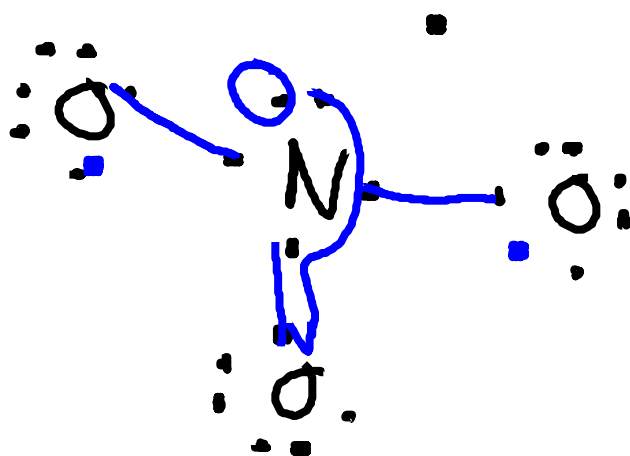
# Problem Set 3

# Polyatomic Ion

- Polyatomic ion- covalently bonded group of ions with a charge



$$5 + 6 + 6 + 6 + 1 = 24 \text{ electrons}$$



# Geometric Shape

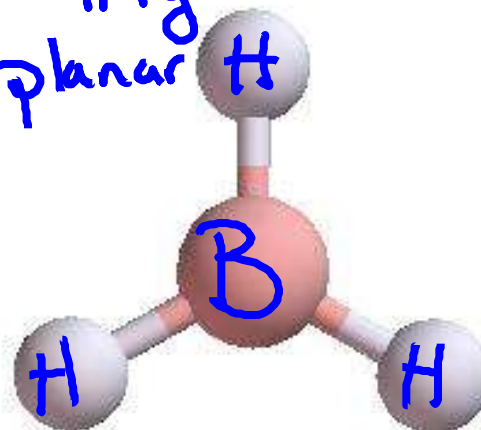
- VSEPR-valence shell electron-pair repulsion.
- Since electrons repel each other, electron pairs will be as far apart as possible

Shape	Number of atoms on Central atom	Number of unshared pairs of electrons	Example
Linear			H--Cl
Linear	2	0	O=C=O
Trigonal Planar	3	0	H <sub>2</sub> CO
Tetrahedral	4	0	CH <sub>4</sub>
Bent	2	2	H <sub>2</sub> O
Trigonal Pyramidal	3	1	NH <sub>3</sub>

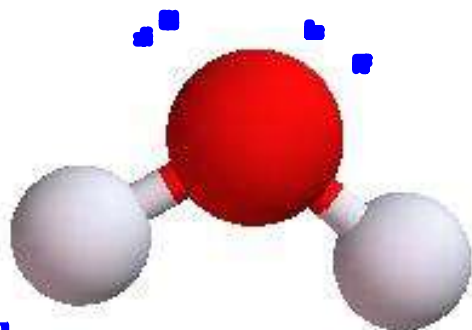
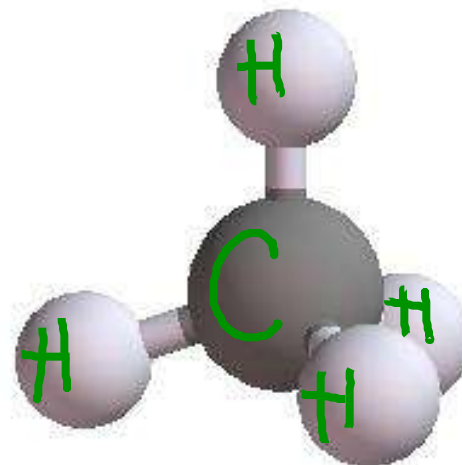
linear



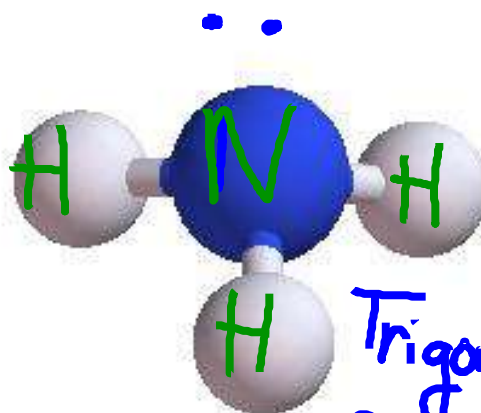
Trigonal  
planar



Tetrahedral



bent



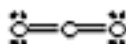


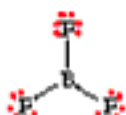
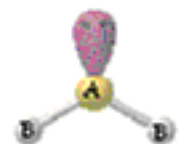
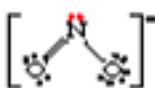


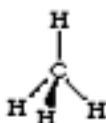
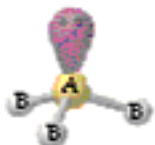
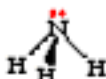
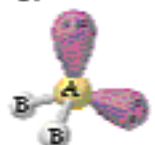



Trigonal  
pyramidal

# Polar Molecules

- Polar Molecules must contain at least one polar bond and shaped so that there is a + and – end.
- water

**TABLE 9.2 Electron-Pair Geometries and Molecular Shapes for Molecules with Two, Three, and Four Electron Domains About the Central Atom**

Total Electron Domains	Electron-Domain Geometry	Bonding Domains	Nonbonding Domains	Molecular Geometry	Example
2 pairs	 Linear	2	0	 Linear	
3 pairs	 Trigonal planar	3	0	 Trigonal planar	
		2	1	 Bent	
4 pairs	 Tetrahedral	4	0	 Tetrahedral	
		3	1	 Trigonal pyramidal	
		2	2	 Bent	

# Non-polar Molecule

- Non-polar molecule contains only nonpolar bonds or contains polar bonds, but has no charged ends.

# Intermolecular Forces

- Forces of attraction between molecules.
- Are weaker than covalent and ionic bonds.

# Types of Intermolecular Forces

- Dipole-dipole forces: force of attraction between + end of one molecule and – end of another molecule
  - The strongest intermolecular force.

# Types of Intermolecular Forces

- Hydrogen Bonding: occurs in molecules with H-F, H-O, H-N
- Large + charge on H is attracted to an unshared pair of electrons on a neighboring molecule

# Types of Intermolecular Forces

- London Dispersion Forces: weak intermolecular forces resulting from constant motion of electrons.
- The only type of intermolecular force between nonpolar molecules.
- Induces a temporary polar end on a molecule.