Chemical Bonding

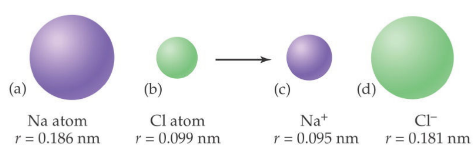
- Minimization of energy is one of the driving forces that cause chemical reactions to occur.
- This minimization refers to the potential energy caused by attractions and repulsions between charged particles in the structure of the compound.
- The term **chemical bond** describes the forces that hold atoms together.
- The **valence electrons** are the outermost electrons that are responsible for holding two or more atoms together in a chemical bonding.
- (ca 1920) The American scientist Gilbert Newton Lewis observed that many elements are more stable when they contained eight valence electrons.
- He proposed that atoms with fewer than eight valence electrons bond together to share electrons and complete their valence (outer shell) electrons.

Octet Rule

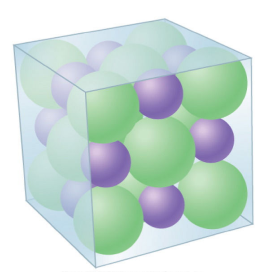
Types of Chemical Bonding

Ionic Bonding

- Electrons are *completely transferred* from one atom to another.
- They lose or gain electrons to form ions.
- The oppositely charged ions are attracted by electrostatic forces (basis of ionic bonding).
- Ionic bonds form between metals and nonmetals.



The atomic radius changes.



3-D crystal structures form.

Formation of Anions

Formation of Cations

Example:

Which of the following statements is correct regarding an ionic bond between aluminum and iodine?

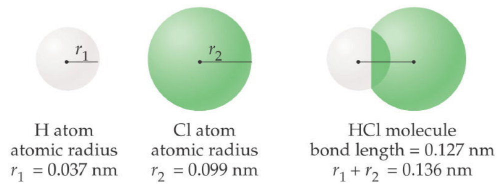
1. The aluminum atom loses electrons, and the iodine atom gains electrons.
2. The aluminum atom is larger in radius than the aluminum ion.
3. The iodine atom is smaller in radius than the iodine ion.
4. The aluminum and iodine ions form a bond by electrostatic attraction.

Answer: They are all correct!

Types of Chemical Bonding

Covalent Bonding

- Occurs when two or more elements *share* electrons.
- Covalent bonding occurs because atoms in the compound will want to gain electrons to fill their valence shells.
- Some atoms can share multiple pairs of electrons
  = multiple bonds.
- Occurs between two nonmetals.



The distance between two nuclei is called the bond length.

Bond Energy

- Energy is released when two ions are attracted to one another and form an ionic bond.
- Energy is also released when two atoms are attracted and form a covalent bond.

$$H\,(g) + Cl\,(g) \longrightarrow HCl\,(g) + \text{heat}$$

- Energy is required to break an H-Cl bond.

$$HCl\,(g) + \text{heat} \longrightarrow H\,(g) + Cl\,(g)$$

The amount of energy required to *break* a covalent bond between two atoms is called the **bond energy**.

Example:

Which of the following statements is correct regarding a covalent bond between a hydrogen atom and a sulfur atom in H_2S?

1. Valence electrons are shared by a hydrogen atom and a sulfur atom.
2. Bonding electrons are distributed over the entire H_2S molecule.
3. The H-S bond length is less than the sum of the two atomic radii.
4. Breaking the H-S bond requires energy.

Answer: They are all correct!

Types of Chemical Bonding

Metallic Bonds

- A metallic bond occurs because of the attractive forces between the positively charged metal ions and the negatively charged valence electrons that move among them.
- Effectively, a metal consists of metal ions surrounded by a freely moving "sea of electrons".
- This gives metals their characteristic traits:
- They are able to conduct electricity and they are maleable.

Lewis-Dot Structures

The symbol of the element is surrounded by the number of dots that match the number of valence electrons.

S: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$ or S: [Ne] $3s^2\ 3p^4$

So S has 6 valence electrons

Electron Dot Formulas of Molecules

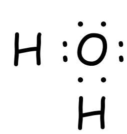
Guidelines for Drawing Electron Dot Formulas of Molecules

1. Calculate the total number of valence electrons.
2. Divide the total number of valence electrons by two to find the number of electron pairs.
3. Surround the central atom with four electron pairs. Use the remaining electron pairs to complete an octet around each atom (except H).
4. The electron pairs shared by atoms are called bonding electrons. The electron pairs that complete an octet but are not shared are called lone pairs (nonbonding pairs).
5. You can move a lone pair to complete an octet.

H_2O

O = 6 VE
H = 1 VE x 2 = 2 VE
Total Valence Electrons = 8

Number of pairs = 8/2 = 4



Examples:

- $H\mathbf{C}N$
- $\mathbf{C}HCl_3$
- $\mathbf{C}O_2$
- $\mathbf{N}H_3$
- $\mathbf{S}O_3$

The central atom is in bold.

Electron Dot Formulas of Polyatomic Ions

Guidelines for Drawing Electron Dot Formulas of Molecules

1. Calculate the total number of valence electrons. Add electrons for anions and subtract electrons for cations.
2. Divide the total number of valence electrons by two to find the number of electron pairs.
3. Surround the central atom with four electron pairs. Use the remaining electron pairs to complete an octet around each atom (except H).
4. The electron pairs shared by atoms are called bonding electrons. The electron pairs that complete an octet but are not shared are called lone pairs (nonbonding pairs).
5. You can move a lone pair to complete an octet.

NH_4^+

N = 5 VE
H = 1 VE x 4 = 4 VE
(+1 charge) = -1 VE
Total Valence Electrons = 8 VE

Number of electron pairs = 8/2 = 4

Examples:

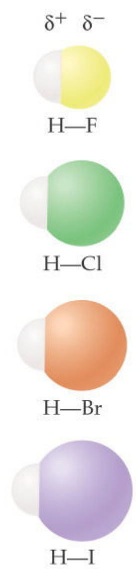
- **Br**O_3^-
- **S**O_4^{2-}
- **C**N^-
- **P**O_4^{3-}

The central atom is in bold.

Polar and Nonpolar Covalent Bonds

- A bond in which bonding electrons are shared equally is a **nonpolar bond**.

- A bond in which bonding electrons are shared unequally is a **polar bond**.



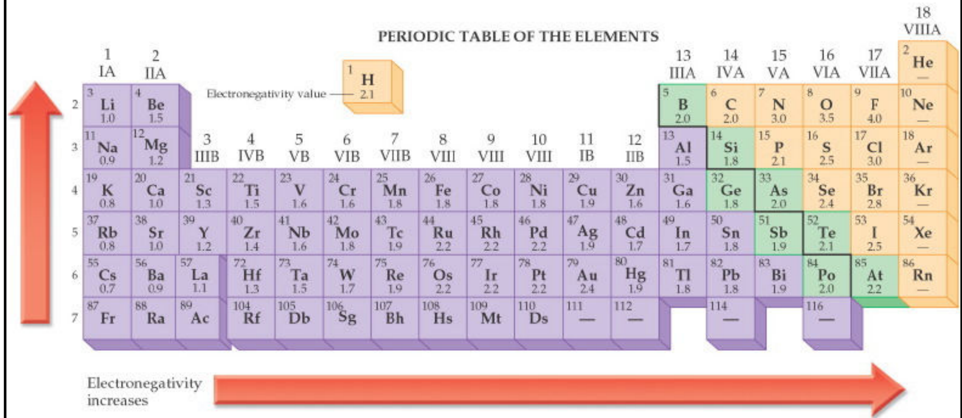
Symbols used to indicate polarity:

δ^+ = less electronegative atom

δ^- = more electronegative atom

$+\!\!\longrightarrow$ points toward more EN atom

The **electronegativity** of an element is the ability of an atom of that element in a molecule to attract bonding electron pairs to itself.



Nonpolar covalent bonds can occur between two atoms with equal EN values.

- Diatomic molecules have nonpolar covalent bonds.

  Examples:
  H_2, O_2, N_2, Cl_2, F_2, I_2, Br_2

- Nonpolar covalent bonds can also occur between *different* atoms which have identical EN values.

  Examples:
  N-Cl, C-S, Si-Ge, etc.

Examples:

1. For each of the bonds below:
   (a) Use delta notation to indicate the difference in EN, and
   (b) Use an arrow to indicate polarity.

C-F   C-O   C-Cl   O-H

2. Identify the bonds by circling one for each:

(a) The C-O bonds in CO_2	ionic	polar covalent	nonpolar covalent
(b) The Cl-Cl bonds in Cl_2	ionic	polar covalent	nonpolar covalent
(c) The Na-O bonds in Na_2O	ionic	polar covalent	nonpolar covalent
(d) The C-C bonds in C_4H_{10}	ionic	polar covalent	nonpolar covalent

Valence Shell Electron Pair Repulsion Theory
VSEPR (pronounced "vesper")

- Electron pairs surrounding an atom tend to repel each other and the shape of the molecule is the result of this electron pair repulsion.
- Electron Pair Geometry – indicates the arrangement of bonding and lone pair electrons around the central atom.
- Molecular Shape – indicates the arrangement of atoms around the central atom as a result of electron pair repulsion.
- A bond angle is formed by any two atoms bonded to the central atom.

Table 12.1 Summary of VSEPR Theory

Bonding/Nonbonding Electron Pairs	Electron Pair Geometry	Molecular Shape	Bond Angle	Example Molecule
4 / 0	tetrahedral	tetrahedral	109.5°	CH_4
3 / 1	tetrahedral	trigonal pyramidal	107°	NH_3
2 / 2	tetrahedral	bent	104.5°	H_2O
4 / 0*	linear	linear	180°	CO_2
3 / 0**	trigonal planar	trigonal planar	120°	CH_2O

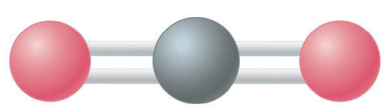
* Carbon dioxide, CO_2, has two double bonds. The three atoms form a "straight line," with carbon in the center; thus, the molecular shape is termed linear.

** Formaldehyde, CH_2O, has one double bond. The four atoms form a "flat triangle" with carbon in the center; thus, the molecular shape is termed trigonal planar.

A = Central Atom **B = Outer Atom** **E = Lone Pair**

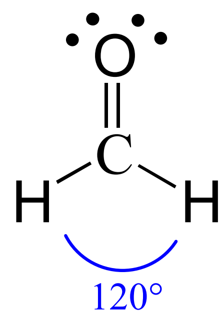
Linear – AB_2

O::C::O



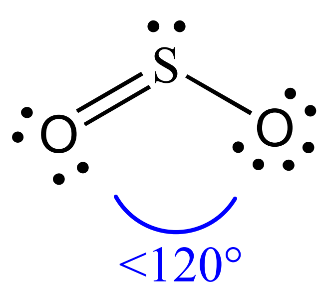
Trigonal Planar – AB_3

Formaldehyde, CH_2O

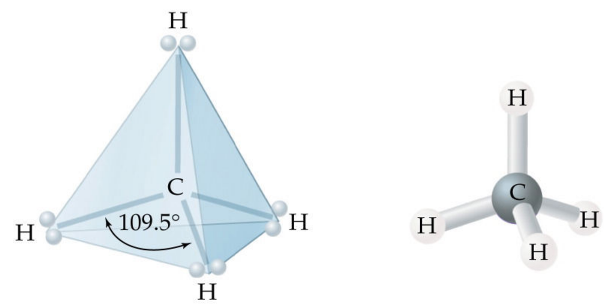


Trigonal Planar – Bent – AB_2E

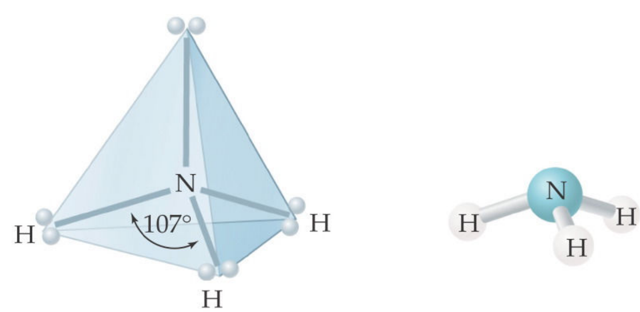
SO_2



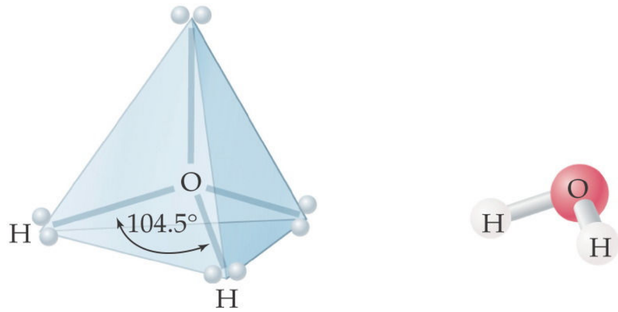
Tetrahedral – AB_4



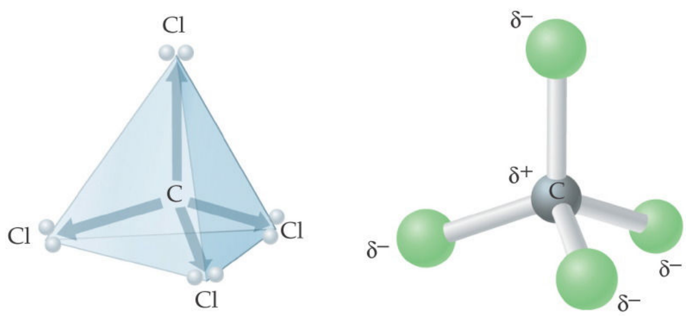
Tetrahedral – Trigonal Pyramidal – AB_3E



Tetrahedral – Bent – AB_2E_2



Nonpolar Molecules with Polar Bonds



Dipoles cancel each other out, therefore a nonpolar molecule.