CHAPTER 6:
CHEMICAL NAMES AND FORMULAS: MOLECULAR COMPOUNDS

CHAPTER 16: COVALENT BONDING

6.1 – Introduction to Chemical Bonding

A **chemical bond** is a mutual electrical attraction between the nuclei and valence electrons of different atoms that binds the atoms together. Valence electrons determine the chemical properties of an element.

- **Valence electrons:** electrons in the highest occupied energy level of an element's atoms. ("s" and "p" sublevel)
- **How do you determine the number of valence electrons?**
  - The group number is the number of valence electrons
  - The electron configuration diagram

6.1 – Introduction to Chemical Bonding

In forming compounds, atoms tend to achieve the electron configuration of a Noble gas.

Stable Octet: 8 valence electrons, or a Noble gas configuration

Octet rule: atoms will lose, gain, or share electrons to achieve a stable octet of electrons.

6.1 – Introduction to Chemical Bonding

In forming compounds, atoms tend to achieve the electron configuration of a Noble gas.

There are two types of bonding:

- **Ionic bonding** is bonding that results from the electrical attraction between anions and cations
- **Covalent bonding** results from the sharing of electron pairs between two atoms

Covalent Bonds

Covalent bond: A bond between two **nonmetals** in which they **share electrons** to form a stable octet.

Atoms can share 2 (single bond), 4 (double bond), or 6 (triple bond) electrons.

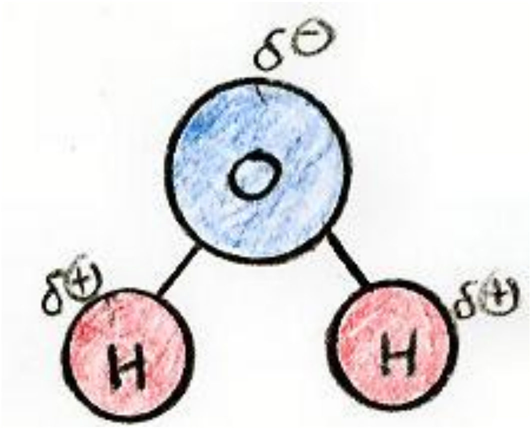
Ionic or Covalent?

Bonding between atoms of different elements is rarely purely ionic or covalent.

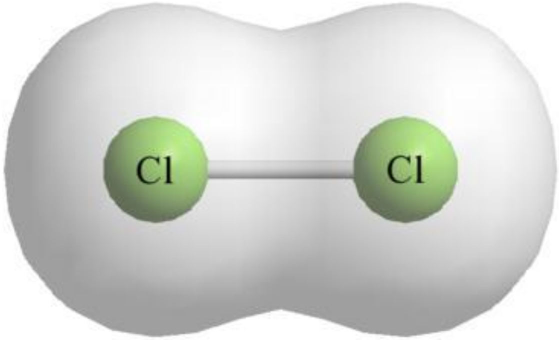
- The degree of ionic or covalent bonding is determined by the differences in the electronegativity of the elements.
  - Polar covalent bonds
  - Nonpolar covalent bonds
  - Ionic bonds

Covalent Bonding

<u>Polar Covalent</u>: The difference in the elements electronegativities is 0.4-1.7



<u>Non-Polar Covalent</u>: slight to no difference in elements electronegativities (0-0.4)



6.2: Chemical Formulas

Shows what is found in a chemical compound:

- Types of atoms
- Numbers of atoms

Monatomic elements are represented by their atomic symbols: Helium: He

If more than one atom is present, the number of atoms is represented with a subscript.

- **Example:** Hydrogen H_2

Molecular Compounds

Molecules – a neutral group of atoms held together by covalent bonds

- Consist of nonmetal-nonmetal bonds

- **Molecular compound –** a chemical compound whose simplest units are molecules.
- **Molecular formula-** shows the types and numbers of atoms combined in a single molecule of a molecular compound.

Diatomic Molecules

Diatomic Elements – a group of elements that naturally exist as two atoms covalently bonded together

Hydrogen H_2

Fluorine F_2

Oxygen O_2

Nitrogen N_2

Chlorine Cl_2

Bromine Br_2

Iodine I_2

Molecular Formula

Shows the type of atoms and numbers present in a molecule of a compound

- **Example:** Molecular formula of water- H_2O
  - Notice there is no need for a subscript next to oxygen

Tells us the composition of a molecule

- Does not tell us about the structure of the molecule (does not show the arrangement of the atoms)

CHAPTER 6.5
MOLECULAR COMPOUNDS

Properties of Molecular Substances

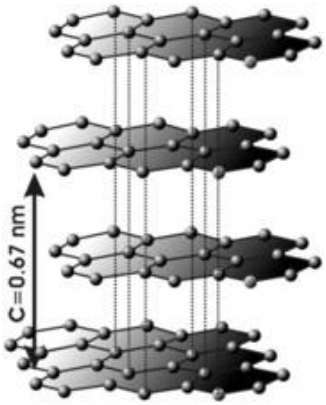
- Exist in all states of matter
- Melting points and boiling points are low compared to ionic compounds
- Some exceptions:
  - Network Solids – stable substances in which all of the atoms are covalently bonded to each other
  - All atoms are interconnected
  - Ex) Diamond & Silicon carbide

Allotropes

Definition – different forms of carbon with different types of bonding

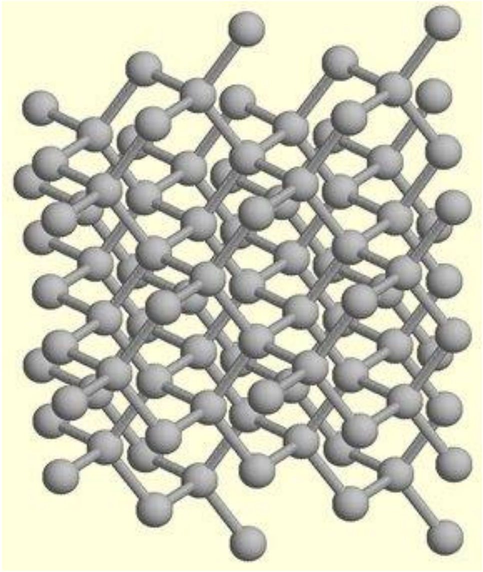
- Carbon – 3 allotropes

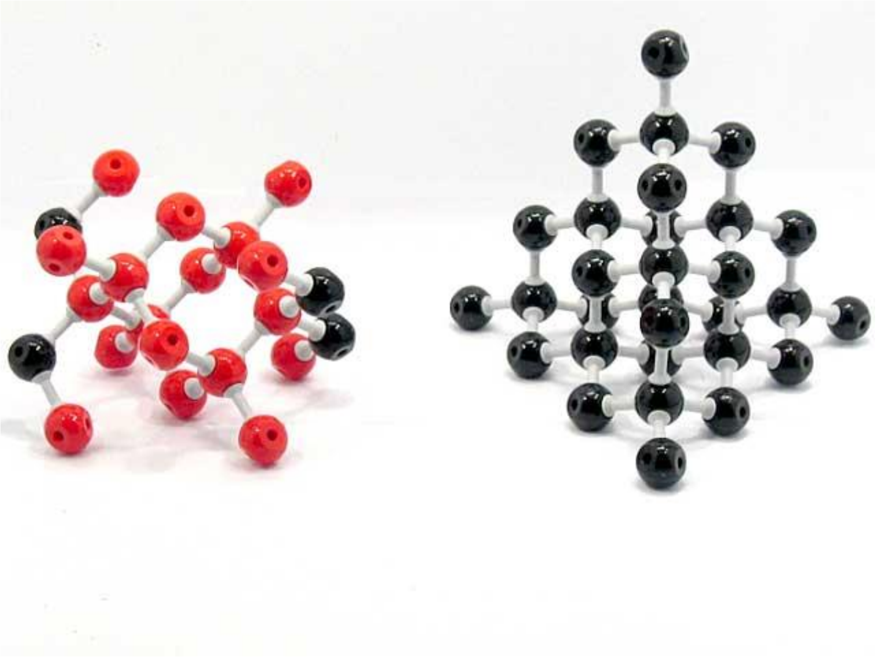
  1. Graphite



Allotropes

2. Diamond

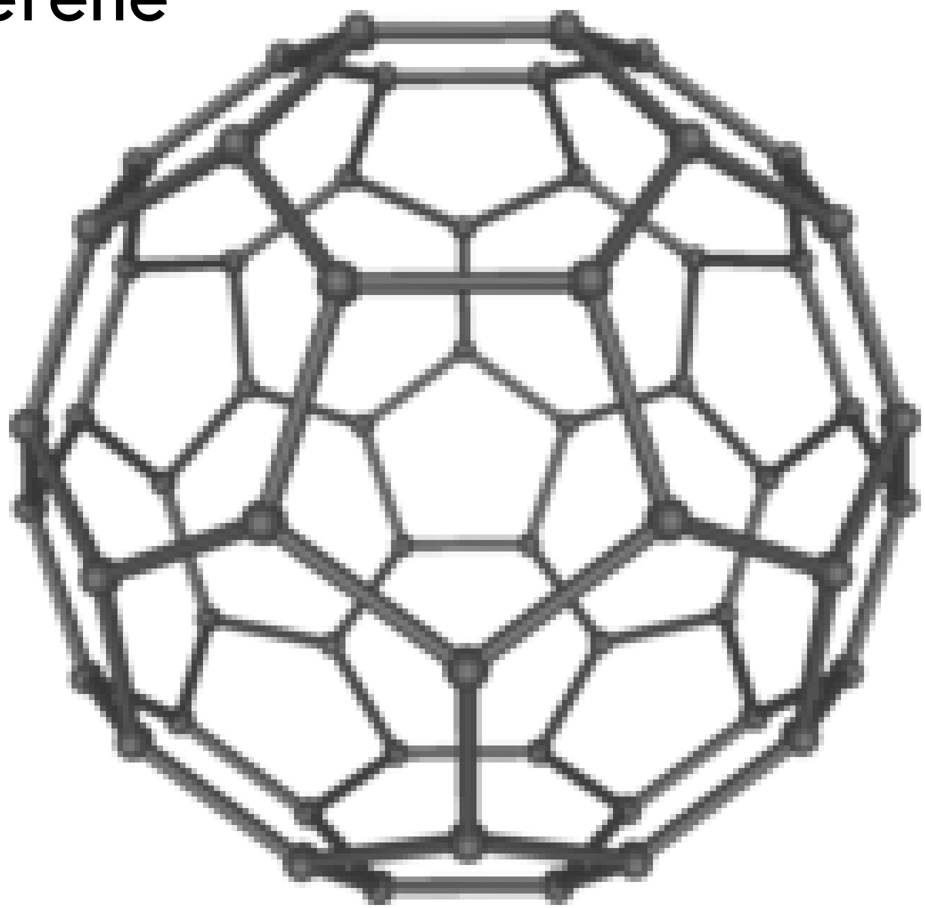




Allotropes

3. Buckminsterfullerene

   C_{60}



Binary molecular compounds:

- Composed of two non- metallic elements
- Ionic charges are not used to assign **formulas** or **names**
- When two non- metallic elements combine, they can often combine in more than one way
- Example: Carbon and Oxygen

Binary molecular compounds:

Two molecular compounds composed of only carbon and oxygen:

- Carbon dioxide CO_2
- Carbon monoxide CO

Prefixes- tell how many atoms of each element are present in each molecule

Binary molecular compounds:

PREFIX	NUMBER
MONO	1
DI	2
TRI	3
TETRA	4
PENTA	5
HEXA	6
HEPTA	7
OCTA	8
NONA	9
DECA	10

Rules for writing molecular names:

1. Use the prefix to tell you the subscript of each element in the formula
2. Write the correct symbols for the two elements, with the appropriate subscripts

Example:

Tetraiodine nonoxide

- tetra = four, so I_4
- non or nona = nine, so O_9

Formula: I_4O_9

Examples:

Write the following molecular formulas:

- sulfur trioxide
- phosphorus pentafluoride

Rules for writing molecular formulas:

1. Use the subscript to tell you the prefix of each element in the formula.
2. Write the name of the first element, using a prefix if the number is greater than one. (<u>No</u> <u>Mono</u>)
3. Write the name of the second element using a prefix, end the name with –ide.

CHAPTER 16

Covalent Bonding

Covalent Bonds

Covalent bond: A bond between two **nonmetals** in which they **share electrons** to form a stable octet.

Atoms can share 2 (single bond), 4 (double bond), or 6 (triple bond) electrons.

Chemical Bonding

An electrostatic force of attraction between two atoms, ions, or molecules

The Octet Rule – all atoms want 8 valence e- (full s and p)

- Exceptions: Hydrogen and Helium:
  - Only want 2 e^-
- Valence e^-: Valence electrons: are the electrons in the highest occupied energy level of an element's atoms

Lewis Dot Structures

- Use the name of the representative element group to determine the # of **valence e^-'s**

Valence electrons

Grps	1A	2A	3A	4A	5A	6A	7A	8A
Per. 2	Li	Be	B	C	N	O	F	Ne
e^- config	s^1	s^2	s^2p^1	s^2p^2	s^2p^3	s^2p^4	s^2p^5	s^2p^6
Lewis Dot	1	2	3	4	5	6	7	8

Lewis Dot Structures

Electron dot structures are used to show the valence electrons.

- They are diagrams that show valence electrons as dots.
- Since valence electrons refer to the s and p sublevels, there can be a total of 8 electrons.
- Each "dot" represents an electron.

PERIODIC TABLE ELEMENTS 1–20

HYDROGEN 1 H							HELIUM 2 He
LITHIUM 3 Li	BERRYLLIUM 4 Be	BORON 5 B	CARBON 6 C	NITROGEN 7 N	OXYGEN 8 O	FLOURINE 9 F	NEON 10 Ne
SODIUM 11 Na	MAGNESIUM 12 Mg	ALUMINUM 13 Al	SILICON 14 Si	PHOSPHORUS 15 P	SULFUR 16 S	CHLORINE 17 Cl	ARGON 18 Ar
POTASSIUM 19 K	CALCIUM 20 Ca						

Rules for Drawing Dot Structures

1. Determine the number of shared electrons. (How many electrons do they need to obtain an octet?)
2. Place 1 pair of electrons in each bond.
3. Decide where any leftover bonding electrons should go.
4. Fill in the molecule with the rest of the electrons to give all atoms an octet.

Structural formulas

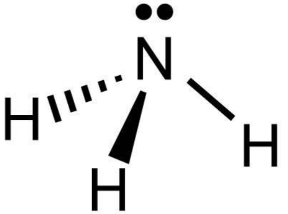
Electron pairs in dot structures can be replaced by lines to make a structural formula.

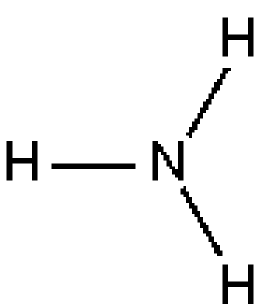
Single bond = 1 line	Cl-Cl
Double bond = 2 lines	O=O
Triple bond = 3 lines	N≡N

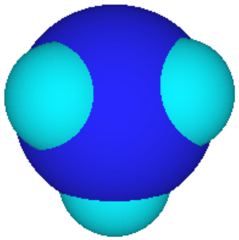
Molecular models:

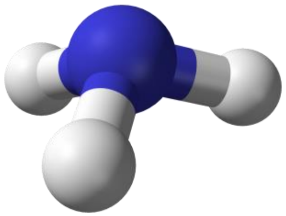
There are a variety of models that describe the arrangement of molecules:

NH_3







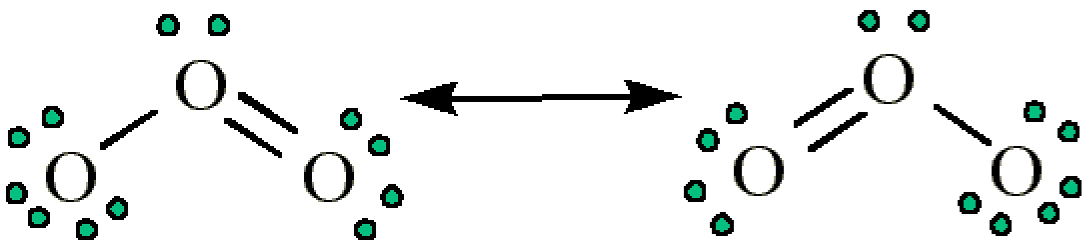


Resonance Structures

Resonance: when 2 or more equally valid electron dot structures can be written for a molecule

- Ozone: O_3

- Proof: bond lengths are the same, there is no clear side for the single and the double bond



VSEPR Theory

- Unpaired electrons around a central atom play a large role in determining a molecule's 3-D shape
- Negatively charged electrons <u>repel</u> one another
  - electron pairs in different orbital stay as far apart as possible

VSEPR THEORY

Valence Shell Electron Pair Repulsion

- Predicts the shapes of molecules

Bonds are made from **electron pairs**

- "bonding pairs"
- "lone pairs"

The bonding pairs and lone pairs around an atom are negatively charged and will get *as far apart from each other as possible.*

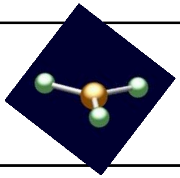
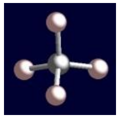
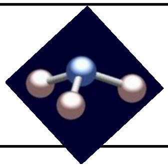
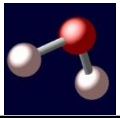
VSEPR Theory

The tendency of electron pairs to adjust the orientation of their orbitals to maximize the distance between them

- Depends on the number of electrons or atoms bonded to a central atom
- Bond angle: shape characterized between the central atom and the atoms bonded to it

VSEPR THEORY

The number of electron pairs will determine the shape of the molecule

Electron Pairs	Orbital Angles	Shape
2	180°	Linear 
3	120 °	Trigonal planar 
4	109.5 °	Tetrahedral 
4 (1 lone pair)	107 °	Trigonal pyramidal 
4	105 °	Bent 

VSEPR Shapes

Linear

Bent

Trigonal Planar or Bent

Trigonal Pyramidal

Tetrahedral

Polarity of Covalent Bonds

Polar bonds: Bonds with uneven sharing of electrons

Non Polar Bonds: Bonds with even Sharing of electrons

Polarity of Molecules

Polarity of molecules: Depends on bonds and shape of molecule.

Nonpolar bonds only = Nonpolar molecule

Polar bonds = Polar molecule or nonpolar molecule (equal and opposite pull)

Intermolecular Forces

Intermolecular Forces

- An attractive force that operates between molecules
- * DO NOT confuse with bonds! *
  - Bonds: attractive forces that hold atoms together in molecules
- IMF are much weaker than bonding forces

Intermolecular Forces

van der Waals forces: collection of the weak interactions

Types:

1. London dispersion force
2. Dipole-dipole force (already covered)
3. Hydrogen-bonding force

London Dispersion Forces

Electrons are in constant motion and aren't always equally distributed

- Therefore they develop a temporary dipole, known as an <u>induced dipole</u>
- The effect passes onto other atoms, like a domino effect… and so on, and so on…

<u>London Dispersion Forces</u>

- Attraction between temporary dipoles of molecules

London Dispersion Forces (L.D.F.)

What do we know?

1. Occur between <u>all</u> atoms and molecules
2. The <u>only</u> intermolecular force at work in <u>nonpolar</u> substances
3. Relatively weak

Dipole-Dipole Forces

Dipole-Dipole Force:

- Attractions among ***polar molecules***
- Electronegativity of atoms determines which part is the:
  - Partial positive (δ +)
  - Partial negative (δ -)
    - Positive and negative parts attract!

Hydrogen Bonding

Hydrogen Bonding:

- An especially strong dipole-dipole force between polar molecules that contain hydrogen attached to a highly electronegative element
  - Although, there is no bond between molecules in the "usual" sense
  - H-bond is a special type of dipole-dipole force

Intermolecular forces

Dispersion forces: attraction present in all molecules (increases with size)

Dipole forces: attraction between polar molecules

Hydrogen bonds: a very strong dipole in molecules with H-Cl, H-F, H-N bond