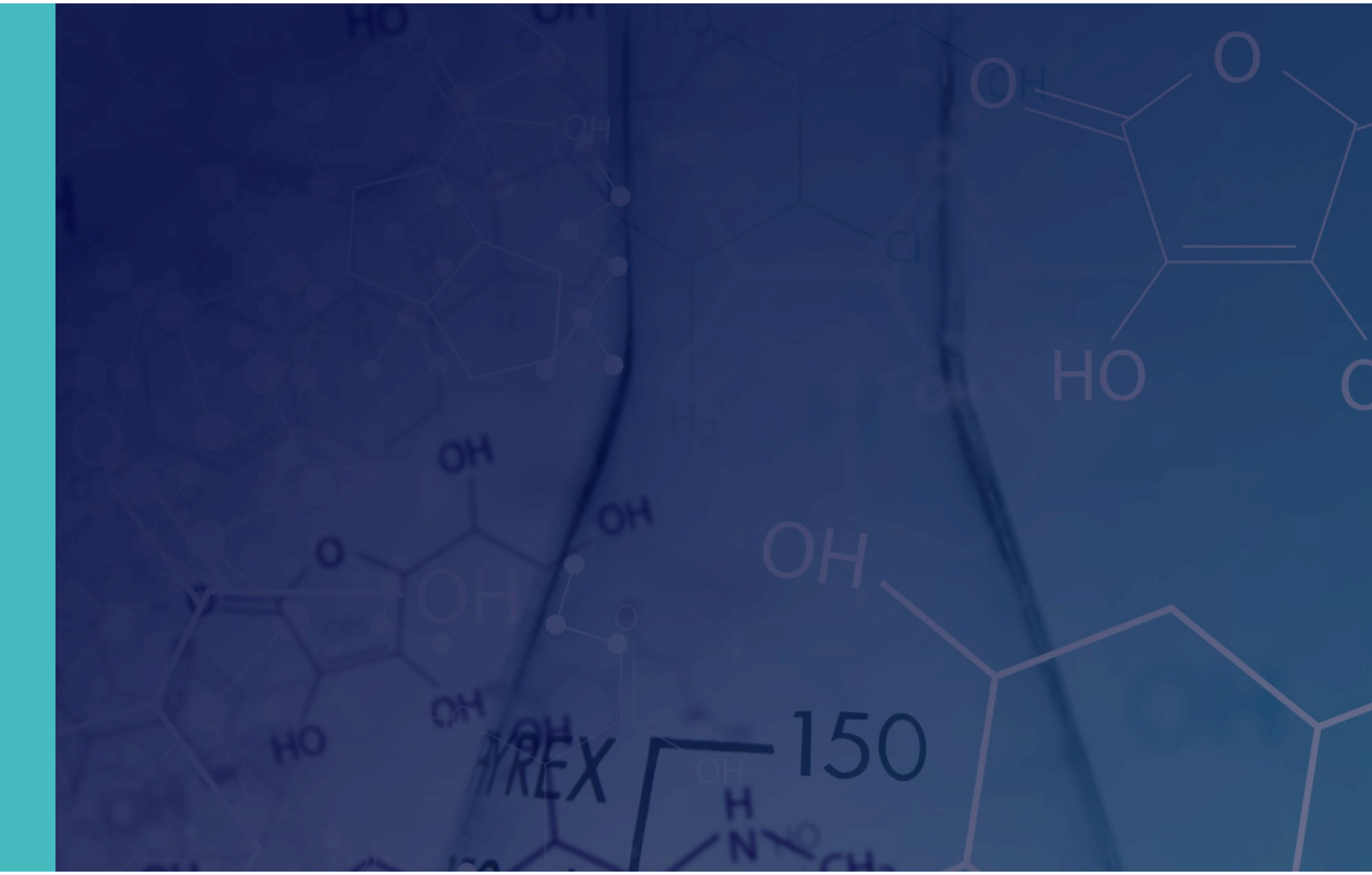


Ontario High School

Grade 11 Chemistry

Summer 2024, Chapter 3 Notes

Table of Contents

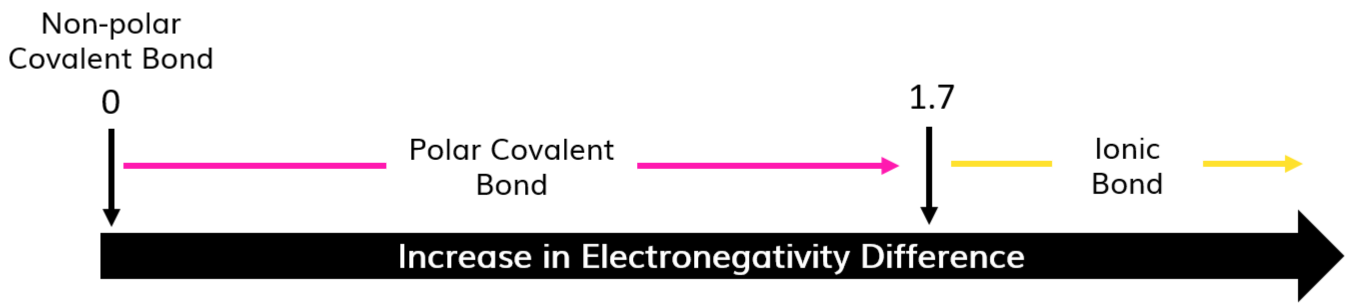
3. Chemical Bonding

3.1 Bonding Theory

3.1.1

Types of Chemical Bonds

- We classify chemical bonds based on the difference in electronegativity between the atoms that form said bond

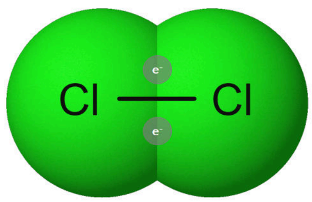


Covalent Bonds:

- Covalent bonds are bonds where electrons are **shared** between **two non-metals**

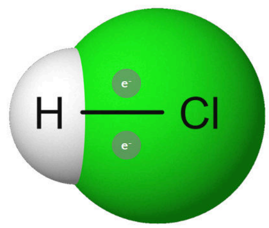
Non-polar Covalent Bonds:

- In non-polar covalent bonds, electrons are **shared equally** between **two of the same non-metals**
  Examples: H_2, O_2, N_2, Cl_2

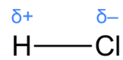


Polar Covalent Bonds:

- In polar covalent bonds, electrons are shared **unequally** between two **different non-metals**
- There is a difference in electronegativity, $0 < \Delta EN < 1.7$. Since electrons are shared unequally in this bond, we say that there is a **dipole moment**

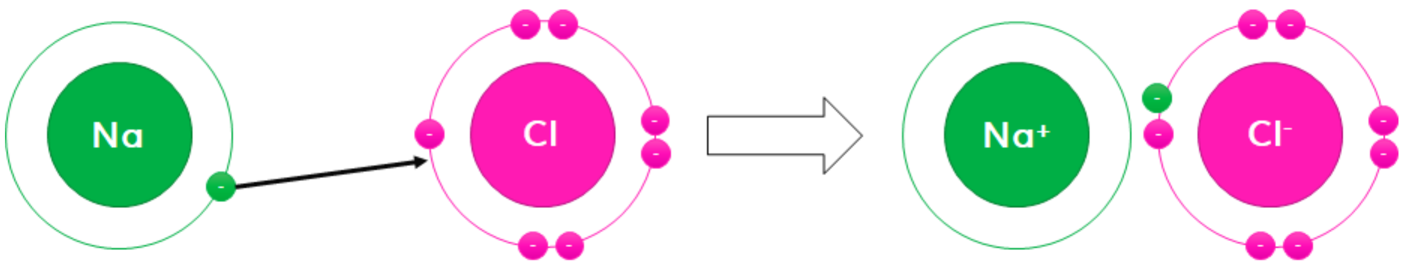


- The dipole moment is a vector with both magnitude and direction
  - Partial negative charge (**δ-**) is assigned to the atom with the **higher electronegativity**
  - Partial positive charge (**δ+**) is assigned to the atom with the **lower electronegativity**
- **The greater the difference in electronegativity (EN), the greater the dipole moment!**



Ionic Bonds

- Are between a **metal and a non-metal.** You might see ionic compounds called salts like NaCl
- There is a **large difference in EN** ($1.7 < \Delta EN$) in these bonds
- The **metal gives electrons to the non-metal.** There is a complete transfer of electrons
  Example: NaCl



3.1.2 **Electronegativity**

Electronegativity

- **Valence electrons** are involved in chemical bonding.
- The type of bond depends on the **difference in electronegativity (ΔEN)** between bonding species.
- **Electronegativity** is the tendency for an atom to draw bonding electrons to itself.

> **WATCH OUT!**
> **This is similar to electron affinity but not the same! Electron affinity involves a single atom/ion, whereas electronegativity involves two bonded atoms.**

Increase in Electronegativity →

H 2.2																	
Li 1.0	Be 1.6											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.3											Al 1.6	Si 1.9	P 2.2	S 2.6	Cl 3.2
K 0.8	Ca 1.0	Sc 1.4	Ti 1.5	V 1.6	Cr 1.7	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.8	Ge 2.0	As 2.2	Se 2.6	Br 3.0
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.3	Nb 1.6	Mo 2.2	Tc 1.9	Ru 2.2	Rh 2.3	Pd 2.2	Ag 1.9	Cd 1.7	In 1.8	Sn 2.0	Sb 1.9	Te 2.1	I 2.7
Cs 0.8	Ba 0.9	Lu 1.1	Hf 1.3	Ta 1.5	W 2.4	Re 1.9	Os 2.2	Ir 2.2	Pt 2.3	Au 2.5	Hg 2.0	Tl 1.6	Pb 2.3	Bi 2.0	Po 2.0	At 2.2

Increase in Electronegativity ↑

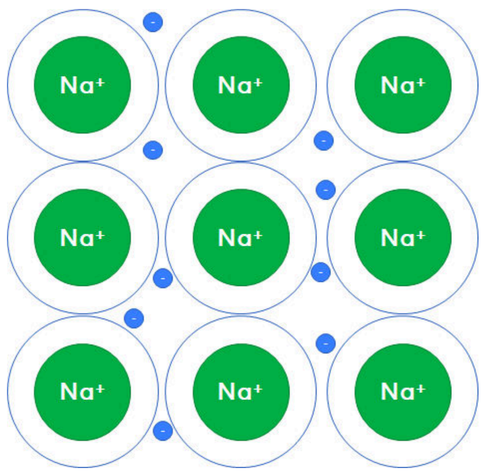
3.1.3

Metallic Bonding

- The diagram below shows many atoms of a metal element surrounded by a **sea of electrons** that are free to move around; we say that the electrons are **delocalized**
  Examples: Aluminum, Iron, Zinc

- **Conduction electrons** are valence electrons that are free to move

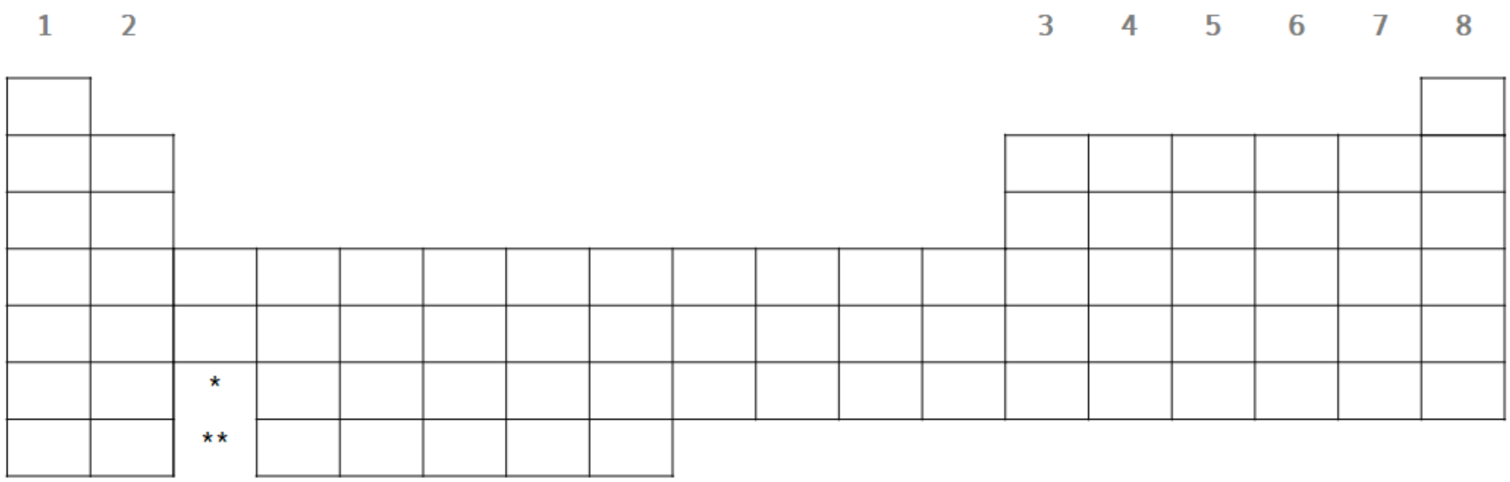
- These conduction electrons are what give the metal their properties!
  - Ductile, malleable, conduct thermal energy, conduct electricity, have luster and shine



3.1.4

Valance Electrons and Ions

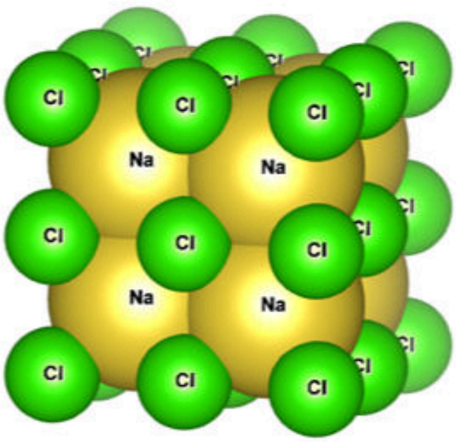
- **Valence electrons** are the electrons found in the outermost shell of an atom. These are the electrons that participate in **bonding**
- The simplest way of determining the number of valence electrons an atom has is by looking at which group an atom is in



- Atoms will form ions by losing or gaining electrons, such that they obtain a full valence shell (full octet).
    - Metals will **lose electrons** to form **cations**
    - Non-metals will **gain electrons** to form **anions**
- **Multivalent** atoms are atoms that can form more than one stable ion. Most transition metals are multivalent
- **Polyatomic ions** are ions containing more than one atom

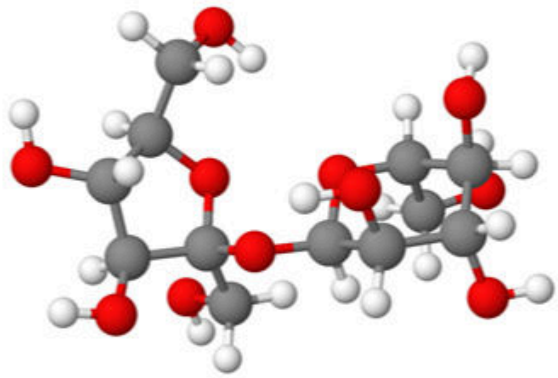
3.1.5

Properties of Ionic and Molecular Compounds



Ionic Compounds

- •Formed by electrostatic attraction between cations and anions
- •Formula has a combination of a metal and a non-metal
- •Most ionic compounds are solid
- •Conduct electricity in aqueous solutions



Molecular Compounds

- •Formed by the sharing of electron pairs between atoms
- •Formula has a combination of two or more non-metals
- •Most are liquids and gases
- •Poor electrical conductivity in aqueous solutions (with some exceptions)

3.1.6 Example: Identifying Chemical Bonds

Example: Identifying Chemical Bonds

Identify the type of bonding between the atoms of the following molecules

H 2.2																
Li 1.0	Be 1.6											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.3											Al 1.6	Si 1.9	P 2.2	S 2.6	Cl 3.2
K 0.8	Ca 1.0	Sc 1.4	Ti 1.5	V 1.6	Cr 1.7	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.8	Ge 2.0	As 2.2	Se 2.6	Br 3.0
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.3	Nb 1.6	Mo 2.2	Tc 1.9	Ru 2.2	Rh 2.3	Pd 2.2	Ag 1.9	Cd 1.7	In 1.8	Sn 2.0	Sb 1.9	Te 2.1	I 2.7
Cs 0.8	Ba 0.9	Lu 1.1	Hf 1.3	Ta 1.5	W 2.4	Re 1.9	Os 2.2	Ir 2.2	Pt 2.3	Au 2.5	Hg 2.0	Tl 1.6	Pb 2.3	Bi 2.0	Po 2.0	At 2.2

a) CO

b) F_2

c) BeO

3.1.7

Out of the following choice, which bond is the most polar?

- H-Br
- H-F
- H-Cl
- H-I

3.1.8

Which of the following compounds displays ionic bonding?

- NH_3
- NaBr
- CO_2
- Na_2SO_4

3.1.9

Practice: Identifying Chemical Bonds

Determine if the elements in the following compounds are metals or non-metals. Describe the type of bonding that occurs in the compound. Use the following abbreviations:

M = metal
NM = non-metal
I = ionic bond
C = covalent bond

Compound	Element 1 (metal or non-metal?)	Element 2 (metal or non-metal?)	Bond Type
NaCl			
CaO			
HF			
AgBr			
CO			

3.2 Naming Conventions

3.2.1

Naming Ionic Compounds

Naming Binary Ionic Compounds

- Binary ionic compounds are composed of a monoatomic cation and a monoatomic anion
  Example: NaCl is made up of Na^+ cations and Cl^- anions
- To name a binary ionic compounds. **first name the metal cation**, followed by the **base name of the non-metal anion with the ending *-ide.*** Subscripts in the formula do not affect the name
  Example: MgO ______________

Naming Binary Ionic Compounds with Multivalent Cations

- Transition metals could form different charged ions. These atoms are known as **multivalent** atoms. You have to indicate the charge when naming the compound

Atom	Ions		
Cr	$3+$	$6+$	
Mn	$2+$	$4+$	$7+$
Fe	$2+$	$3+$	
Co	$2+$	$3+$	

- To name this type of ionic compounds, first **name the metal cation**, followed by the **ion charge in roman numerals in parenthesis** then the **base name of the non-metal anion with the ending *-ide***

  Example: Fe_2O_3 is ________________

Naming Ionic Compounds with Polyatomic Ions

- Naming ionic compounds that contain polyatomic ions are very similar to the naming of binary ionic compounds. **First name the cation**, followed by the **base name of the anion.** Subscripts in the formula do not affect the name

  Example: $NaNO_3$ is ______________

Ion	Name	Ion	Name	Ion	Name
NH_4^+	ammonium	OH^-	hydroxide	CN^-	cyanide
SO_4^{2-}	sulfate	O_2^{2-}	peroxide	CNO^-	cyanate
HSO_4^-	bisulfate	CH_3COO^-	acetate	SCN^-	thiocyanate
SO_3^{2-}	sulfite	$C\ell O_4^-$	perchlorate	CO_3^{2-}	carbonate
NO_3^-	nitrate	$C\ell O_3^-$	chlorate	HCO_3^-	bicarbonate
NO_2^-	nitrite	$C\ell O_2^-$	chlorite	$OOCOO^{2-}$	oxalate
PO_4^{3-}	phosphate ion	$C\ell O^-$	hypochlorite ion	$S_2O_3^{2-}$	thiosulfate
HPO_4^{2-}	hydrogen phosphate	CrO_4^{2-}	chromate	Hg_2^{2+}	mercury
$H_2PO_4^-$	dihydrogen phosphate	$Cr_2O_7^{2-}$	dichromate ion	H_3O^+	hydronium
PO_3^{3-}	phosphite ion	MnO_4^-	permanganate ion	SiO_3^{2-}	silicate

Naming Molecular Compounds

- Covalent bonds are formed between two non-metals. The resulting compound is known as a **molecular element or compound.**
- **Molecular elements** are named the same as the element itself

  Example: O_2 is __________, H_2 is ____________
- **Molecular compounds** are usually written from least to most electronegative.
- To name a molecular compound:
  - Name the first element using the proper prefix (di, tri, etc.). If only one atom, avoid the prefix mono
  - Name the second element with the proper prefix (including mono) using the "ide" ending

Number of Atoms	Prefix	Number of Atoms	Prefix
1	mono-	6	hexa-
2	di-	7	hepta-
3	tri-	8	octa-
4	tetra-	9	nona-
5	penta-	10	deca-

Example:

CO = ____________________

CO_2 = __________________

N_2O_4 = _______________________

3.2.3

Example: Providing Chemical Formulae for Ionic Compounds

Give a chemical formula for the following compounds:

a) Potassium fluoride

b) Manganese(II) oxide

c) Ammonium sulfate

Example: Naming Molecular Compounds

Name the following molecular compounds:

a) NF_3

b) CCl_4

c) N_2O

3.2.5

Practice: Naming Conventions

In naming a binary molecular compound, the number of atoms of each element present in the molecule is indicated by:

- roman numerals
- prefixes
- superscripts
- suffixes

3.2.6

MARK YOURSELF QUESTION

1. Grab a piece of paper and try this problem yourself.
2. When you're done, check the "I have answered this question" box below.
3. View the solution and report whether you got it right or wrong.

Practice: Naming Compounds

Name the following compounds.

a. $KClO$

b. $BaOH$

c. PCl_3

d. FeO

3.2.7

MARK YOURSELF QUESTION

1. Grab a piece of paper and try this problem yourself.
2. When you're done, check the "I have answered this question" box below.
3. View the solution and report whether you got it right or wrong.

Practice: Determining Chemical Formulas based on Names

Write out the formula for the following. Make sure to use subscripts and brackets where appropriate.

a. magnesium hydroxide:

b. sodium chromate:

c. ammonium nitrate:

d. iron (III) phosphate:

3.3 Lewis Diagrams

3.3.1

Lewis Dot Diagrams

- **Lewis dot diagrams** are diagrams in which the **valence electrons** of an atom are shown as dots distributed around the element’s symbol.

Bonds and Lewis Diagrams

- We can use Lewis diagrams to show the formation of **ionic compounds**

- We also use Lewis diagrams to show the formation of **molecular compounds.** A line is used to indicate a shared pair of electrons. These diagrams are also known as **Lewis structures** or **structural formulas**

- Molecular compounds can have double (they share two pairs of electrons) and triple bonds (they three pairs of electrons). We show these using double or triple lines

Drawing Lewis Structures

Steps for Drawing Lewis Structures

Example: CCl_4

1. Calculate the total number of **valence electrons** for the molecule.
   If drawing the Lewis structure for a ion, subtract one electron per positive charge and add one electron per negative charge

2. Write out all atoms, with the **least electronegative atom in the middle**

3. Connect surrounding atoms to the central atom with **single bonds**.

4. Place remaining electrons as **lone pairs** on surrounding atoms (except hydrogen), completing an octet around each atom

5. Place remaining electrons as lone pairs on the central atom.

6. Shift lone pairs to make double or triple bonds to satisfy the **octet rule** for all atoms

3.3.3

Example: Lewis Structures

Draw the Lewis structure of PCl_3

3.3.4

Example: Drawing Lewis Structure with a Double Bond

Draw the Lewis structure for CO_2.

3.3.5

Example: Drawing Lewis Structures for Ions

Draw the Lewis structure for the BF_4^- ion.

3.3.6

Practice: Periodic Table and Valence Electrons

Which group of the periodic table can most easily donate an electron in a chemical reaction?

- Halogens
- Noble Gases
- Alkali metals
- Transition metals
- Alkaline earth metals

3.3.7

Practice: Valence Electrons

Which molecule has the highest number of valence electrons?

- NH_3
- CCl_4
- CO
- Br_2

3.3.8

Practice: Drawing Lewis Structures

Draw the Lewis structure for the following molecules, then use the Lewis structures you drew to answer the following questions.

a. CH_2O
b. H_3O^+
c. NH_4^+

Part 1

Whose Lewis structure has double bonds?

- CH_2O
- H_3O^+
- NH_4^+
- none of them have double bonds

Practice: Drawing Lewis Structures

Draw the Lewis structure for the following molecules, then use the Lewis structures you drew to answer the following questions.

a. CH_2O
b. H_3O^+
c. NH_4^+

Part 2

Whose Lewis structure has a lone pair on the central atom?

- CH_2O
- H_3O^+
- NH_4^+
- none have lone pairs on the central atom

3.4 Valence Shell Electron Pair Repulsion Theory (VSEPR)

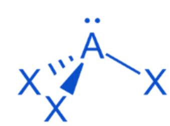
3.4.1

Valence Shell Electron Pair Repulsion Theory (VSEPR)

- **The VSEPR theory** states that repulsion of electron groups (lone pairs and bonds) in the valence shell will determine the 3D geometry of a molecule.
- In VSEPR theory, we focus on the number of **electron groups** around a central atom.
  - Bonding groups
  - Non-bonding groups
- To draw a VSEPR structure, you need the Lewis structure and the number of electron groups.
  - **Wedges** mean the bond is coming out of the page
  - **Dashes** mean the bond is going into the page
- We use all the electron groups to get the **electron-pair geometry**.
  - To get the **molecular geometry**, we look at the atoms present

WIZE CONCEPT

1. Draw Lewis structure
2. Determine the electron-group geometry based on the total electron groups
   a. Two electron groups: linear
   b. Three electron groups: trigonal planar
   c. Four electron groups: tetrahedral
3. Determine the molecular geometry that minimizes repulsions

Total Pairs	**Bonding Pairs**	**Lone Pairs**	**Electron Pair Arrangement**	**Molecular Geometry**	**Stereochemical Drawing**
2	2	0	Linear	Linear	
3	3	0	Trigonal Planar	Trigonal Planar	
4	4	0	Tetrahedral	Tetrahedral	
4	3	1	Tetrahedral	Trigonal Pyramidal	
4	2	2	Tetrahedral	Bent	

3.4.2 Example: VSEPR Shapes

Example: VSEPR Shapes

Determine the electron pair arrangement and molecular geometry for NH_3

3.4.3

Practice: VSEPR Terms and Definitions

The VSEPR model includes several concepts related to atomic theory. Match the following terms and definitions.

A. the outermost energy level or orbit of an atom or ion

B. a shared pair of electrons

C. all pairs of valence electrons repel each other

D. a pair of electrons that is not involved in bonding

☐ valence shell

☐ bonding pair

☐ lone pair

☐ electron pair repulsion

Practice VSEPR Shapes

For the following molecular geometries, fill in the table below with:

- the total number of electron groups
- the number of lone pairs on the central atom
- the number of bonding groups

Molecular Geometry	Total number of electron groups	Number of lone pairs around the central atom	Number of bonding groups
Linear			
Trigonal Planar			
Bent			
Trigonal Pyramidal			
Tetrahedral			

3.4.5

Practice: Determining Electron Pair Arrangement and Molecular Geometries

Use the VSEPR theory with the following molecules to answer the following questions.

a. PF_3
b. NH_4^+
c. H_2S

Part 1

Which molecule has a tetrahedral electron pair arrangement?

- PF_3
- NH_4^+
- H_2S
- all of them have a tetrahedral electron pair geometry

Practice: Determining Electron Pair Arrangement and Molecular Geometries

Use the VSEPR theory with the following molecules to answer the following questions.

a. PF_3
b. NH_4^+
c. H_2S

Part 2

Which molecule has a trigonal pyramidal molecular geometry?

- PF_3
- NH_4^+
- H_2S
- none of them have a trigonal pyramidal molecular geometry

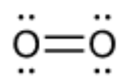
3.5 Intermolecular Forces

3.5.1

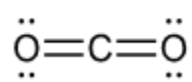
Molecular Polarity

- In **polar covalent bonds**, electrons are shared unequally between two different non-metals
- To determine if a molecule is polar, we need to look at the whole molecule, and not just the individual bonds

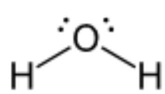
   i. If there are **no polar bonds** in the molecule, the molecule is **non-polar!**
   Example: O_2



   ii. If there are **polar bonds that are symmetrical**, the dipole moments cancel each other out so the molecule is **non-polar!**
   Example: CO_2



   iii. If there are **polar bonds that are not symmetrical**, the dipole moments don't cancel each other out and we are left with a **net dipole moment,** hence the molecule is **polar**!
   Example: H_2O



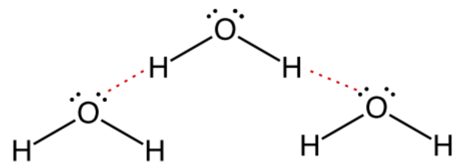
3.5.2

Types of Intermolecular Forces

- **Intermolecular forces** are forces of attraction between molecules.
- Intermolecular forces define physical properties of compounds (boiling points, melting points, surface tension etc.)
- The stronger the intermolecular forces are, the higher the boiling and melting points will be

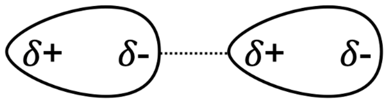
Hydrogen Bonding

- The strongest of the intermolecular forces
- **Hydrogen bonded to N, O, or F** is attracted to lone pairs of **electrons on other N, O or F**



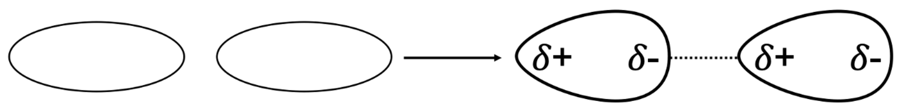
Dipole - Dipole

- Second strongest intermolecular force
- The more polar a molecule is, the stronger the forces
- Interaction between **two polar molecules;** opposite dipoles attract (δ– and δ+)



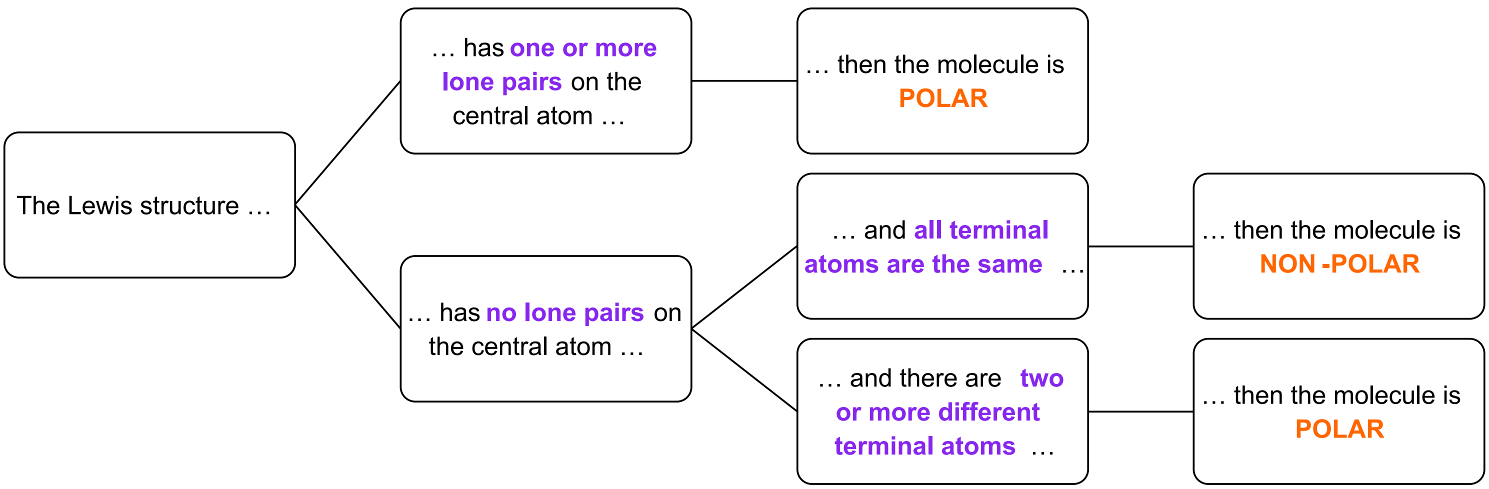
London Dispersion Forces

- The weakest of intermolecular forces.
- All molecules have London dispersion forces.
- The bigger the molecule, the stronger the forces
- Molecules interact randomly and distort each other's electron clouds, causing temporary dipoles
- **Electrons randomly move** around and at some point, **more electrons may be on one side than the other** in a non-polar molecule, creating a **temporary dipole.** This can happen with another non-polar molecule as well and the two can interact



3.5.3

How to Determine Molecular Polarity



3.5.4

Example: Molecular Polarity

Determine if the following molecules are polar or not: CCl_4 and $CHCl_3$

Example: Strength of Intermolecular Forces

Determine which compound in the following pairs of molecules will have the higher melting point.

a) I_2 or Br_2

b) HF or HBr

3.5.6

Practice: Intermolecular Forces and Boiling Points

Which of the following molecules would exhibit the highest boiling point based on intermolecular forces?

- H_2O
- CH_4
- NCl_3

Practice: Molecular Polarity

Which of the following molecules is non-polar?

- SCl_2
- OCl_2
- $BeCl_2$
- all of the above molecules are polar

3.5.8

Describe the strongest intermolecular forces in the following compounds.

A. Dipole-Dipole

B. Hydrogen Bonding

C. London Dispersion

	CCl_4
	NH_3
	H_2S

3.5.9

Practice: Comparing Boiling Points

Hexane, C_6H_{14} (M = 86g/mol) has a boiling point of 68°C. Ethanol, CH_3CH_2OH (M = 46g/mol) has a boiling point of 78°C. Mark each of the following statements as TRUE or FALSE.

a. Ethanol must have stronger intermolecular attractions, based on its higher boiling point.
b. Ethanol has a higher boiling point because of greater London dispersion force
c. Both hexane and ethanol have hydrogen bonding.
d. Ethanol has a higher boiling point due to hydrogen bonding.

a.

b.

c.

d.