

# Chemistry Honors Study Guide

## Test 3 S1

Test date: November 15, 2024

## 1 LDS Continued

### Definitions

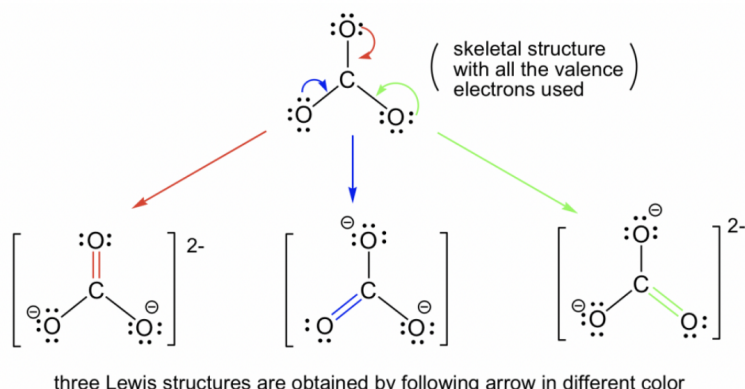
- **resonance:** The sharing of electrons between multiple bonds, esp. in polyatomic ions.

### Resonance LDS

#### Resonance Structures<sup>1</sup>

There are many possible resonance structures of one polyatomic anion depending on the number of double bonds needed to cancel out the formal charge on the central atom.

Example:  $\text{CO}_3^{2-}$  (carbonate)



This diagram shows the possibilities of where the double bond could be in carbonate.

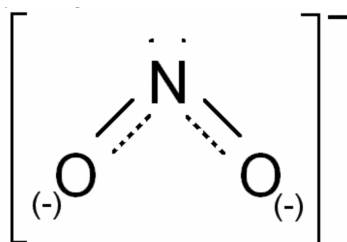
#### Combined Resonance Structures<sup>2</sup>

In a polyatomic ion, when the central atom has a formal charge of 0, the oxygen atoms around it share their remaining formal charges. A dotted line instead of a bond is drawn from each oxygen atom to the central atom to indicate the shared charge.

Example:  $\text{NO}_2^-$  (nitrite)

<sup>1</sup><https://kpu.pressbooks.pub/organicchemistry/chapter/1-3-resonance-structures/>

<sup>2</sup><https://chemfiesta.org/2015/09/18/resonance-structures/>



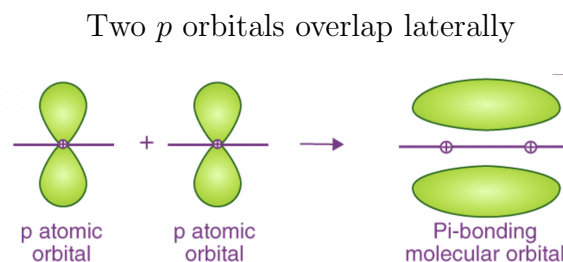
## 2 Bonding Theories and Geometry

### Definitions

- **molecular orbital:** An orbital that applies to the entire molecule.
- **electron domain:** A distinct region around an atom where lone pair electrons/bonds are found.
- **$\pi$  ( $\pi$ ) bond:** A type of covalent bond that is a combination of two atomic orbitals overlapping laterally.
- **$\sigma$  ( $\sigma$ ) bond:** A type of covalent bond that is a combination of two atomic orbitals from end to end. There is exactly one  $\sigma$  bond in each single, double, and triple bond (the rest are  $\pi$  bonds).
- **VSEPR (*ves-per*) theory:** A model predicting the geometric shape of molecules; stands for Valence Shell Electron Pair Repulsion.
- **hybridization:** The combination of two atomic orbitals to create a new type of hybrid orbital.

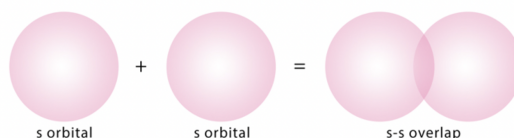
### Molecular Orbitals

#### $\pi$ Bonding<sup>3</sup>



#### $\sigma$ Bonding<sup>4</sup>

$s$ - $s$  overlap (2 half-filled  $s$  orbitals)

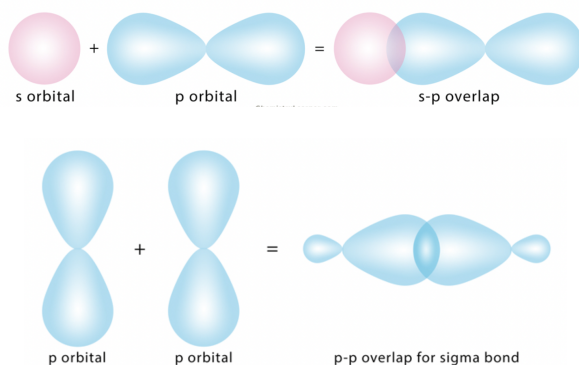


$s$ - $p$  overlap (1 half-filled  $s$  orbital and 1 half-filled  $p$  orbital)

$p$ - $p$  overlap (2 half-filled  $p$  orbitals)

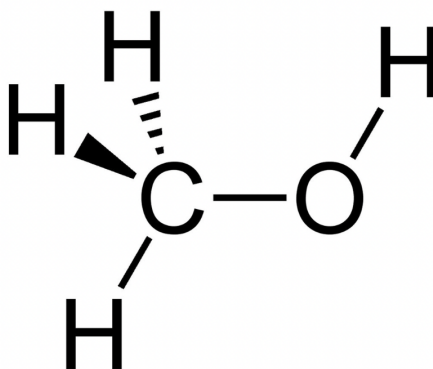
<sup>3</sup><https://byjus.com/chemistry/sigma-and-pi-bond/>

<sup>4</sup><https://www.chemistrylearner.com/sigma-and-pi-bonds.html>



## Drawing a 3-D Molecule<sup>5</sup>

- line = bond on the plane
- 3 lines = bond into the plane
- triangle = bond out of the plane



## VSEPR Theory

(ed = electron domains; lp = lone pair(s))

### Electron Geometry

# ed	Name
2	<i>linear</i>
3	<i>trigonal planar</i>
4	<i>tetrahedral</i>
5	<i>trigonal bipyramidal</i>
6	<i>octahedral</i>

<sup>5</sup><https://www.thoughtco.com/wedge-and-dash-projection-definition-602137>

## Molecular Geometry

# ed	0 lp	1 lp	2 lp	3 lp	4 lp
2	<i>linear</i>	n/a	n/a	n/a	n/a
3	<i>trigonal planar</i>	<i>bent</i>	n/a	n/a	n/a
4	<i>tetrahedral</i>	<i>trigonal pyramidal</i>	<i>bent</i>	n/a	n/a
5	<i>trigonal bipyramidal</i>	<i>seesaw</i>	<i>t-shaped</i>	<i>linear</i>	n/a
6	<i>octahedral</i>	<i>square pyramidal</i>	<i>square planar</i>	<i>t-shaped</i>	<i>linear</i>

## Hybridization

(The numbers before the orbital type are the number of that orbital type there are, not the energy level.)

$1(s) + 3(p) \rightarrow 4(sp^3)$  (**tetrahedral**  $e^-$  geometry)

$1(s) + 3(p) \rightarrow 3(sp^2) + 1(p)$  (**trigonal planar**  $e^-$  geometry)

$1(s) + 3(p) \rightarrow 2(sp) + 2(p)$  (**linear**  $e^-$  geometry)

## 3 Polarity

### Definitions

- **polarity:** Occurs when one atom attracts electron cloud more than another, therefore making one end more negative.
- **dipole:** A pair of separated equal and opposite electric charges.

### Bond Polarity by Electronegativity

Polarity is determined by the following equation:

$$\Delta EN = |EN_1 - EN_2|$$

- If  $\Delta EN \leq 0.4$ , the particle is **nonpolar covalent**.
- If  $0.4 < \Delta EN \leq 1.0$ , the particle is **moderately polar covalent**.
- If  $1.0 < \Delta EN \leq 2.0$ , the particle is **highly polar covalent**.
- If  $\Delta EN \geq 2.0$ , the particle is **ionic**.

### Molecular Polarity

- If a molecule has no polar bonds, it is not polar.
- If a molecule is symmetrical with the same polarity in each bond, it is not polar. (Molecules with linear, trigonal planar, tetrahedral, trigonal bipyramidal, and octahedral molecular geometry and the same element around the central atom are not polar.)
- If a molecule contains (a) more electronegative atom(s) on one end than another, the polarity arrow goes in the direction of the electronegative end.

## 4 Intermolecular Forces (IMFs)

### IMF Types (weakest to strongest)

#### London Dispersion Forces (LDFs)

Occur between all molecules. Only IMF between nonpolar molecules

- weakest IMF
- attractions between instantaneous dipoles due to moving electrons
- also induced dipoles
- $\uparrow$  molecule size, polarizability =  $\uparrow$  LDF

#### Dipole-Dipole Forces (Dip-Dip)

Formed between polar molecules due to attraction between opposite permanent poles

**Dipole notation:**<sup>6</sup>  $\delta^+$  on positive end(s);  $\delta^-$  on negative end(s)

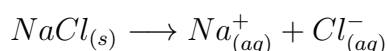


#### Hydrogen “Bonding” (H-Bonding)

Occurs between molecules with N—H, F—H, and O—H bonds. A type of dipole-dipole interaction.

#### Ion-Dipole Interactions

Formed between ionic compounds and polar compounds.



Adding salt to water creates a cage effect, increasing density, because molecules are more tightly packed together due to stronger bonds. In order for ionic compounds to dissolve, ion-dipole forces must be stronger than ionic and H bonds.

### Macro Effects of IMFs

$\uparrow$  IMFs =  $\uparrow$  boiling point, melting point, surface tension, viscosity

Name	Formula	Strongest IMF	Boiling Point
Water	H <sub>2</sub> O	H-bonds	100°C
Methane	CH <sub>4</sub>	LDFs	-162°C
Formaldehyde	CH <sub>2</sub> O	Dip-dip	-19°C
Propane	C <sub>3</sub> H <sub>8</sub>	LDFs	-19°C
Hexane	C <sub>6</sub> H <sub>14</sub>	LDFs	68°C

---

<sup>6</sup><https://byjus.com/chemistry/dipole-moment/>

## Trends

- The relative boiling points of molecules of similar sizes depend on the IMFs present. ( $\uparrow$  force =  $\uparrow$  b.p.)
- The relative boiling points of molecules with the same IMFs depend on the sizes of the molecules. ( $\uparrow$  size =  $\uparrow$  polarizability =  $\uparrow$  b.p.)

## Miscibility

“**Like dissolves like**” — polar + polar and nonpolar + nonpolar are miscible and create a homogeneous mixture; nonpolar + polar are not miscible and create a heterogeneous mixture

# 5 Atoms, Molecules, Moles, Grams

## Definitions

- **mole:** A unit of quantification;  $6.02 \times 10^{23}$  (Avogadro’s number).
- **molar mass:** How many grams a mole of an atom/compound weighs. Units are g/mol.

## Conversions

Example: How many atoms are in 17 grams of  $\text{MgCl}_2$ ?

**Step 1: Draw a conversion map.** In order to get from grams to atoms, the steps are as follows:

$$g \longrightarrow mol \longrightarrow \text{molecules} \longrightarrow \text{atoms}$$

**Step 2: Set up the conversion.** Multiply by the conversion factor(s), (a) unit fraction(s) (equal to 1). Set it up so that the same units cancel out in order to get the desired unit in the end.

$$(17\text{ g}) \times \left( \frac{1\text{ mol}}{95.211\text{ g}} \right) \times \left( \frac{6.02 \times 10^{23}\text{ molecules}}{1\text{ mol}} \right) \times \left( \frac{3\text{ atoms}}{1\text{ molecules}} \right)$$

**Step 4: Calculate.**

$$= 3.225 \times 10^{23}\text{ atoms}$$

**Step 5: Check for reasonableness.** 17 grams of  $\text{MgCl}_2$  should contain a large amount of atoms, which is reflected in the answer.

## Percent Composition

$$\frac{\text{mass of element}}{\text{mass of compound}} \times 100 = \% \text{ composition}$$

Example: What is the percent composition of iron (Fe) in  $\text{Fe}_2\text{O}_3$ ?

$$\frac{2(55.85 \text{ g/mol})}{2(55.85 \text{ g/mol}) + 3(15.999 \text{ g/mol})} \times 100 = 69.94\%$$