Chapter 9: Bonding & Molecular Structure: Fundamentals

Recall: Valence electrons are the outermost s & p electrons = Group No.
- determine chemical properties of atom
- loss & gain or rearrangement of these e\(^-\) caused by reactions

Lewis dot symbols for main group atoms (we valence electrons)
1A 2A 3A 4A 5A 6A 7A 8A
Li: • Be: • B: • O: • N: • O: • F: • Ne:

Chemical bonds: attractive forces that hold atoms together in compounds

1) Ionic bonds: electrostatic interaction between ions
- 
- Involves transfer of 1 or more electrons
- is proportional to distance between ions
- occurs between metals (low EN) and non-metals (high EN)

Mg and N combine together to form what simple binary compound?
\[ \text{Mg} + \text{N} \rightarrow \text{?} \] (ans. Mg\(_3\)N\(_2\))

First ask how do Mg and N both gain noble gas configuration?

\[ \text{Mg}^+ + \text{N}^3- \rightarrow \text{Mg}^{2+} \text{ and N}^{3-} \text{ in compound} \]

- Mg must lose 2e\(^-\)
- N must gain 3e\(^-\)
- iso-electric with Ne

But total number of e\(^-\) lost by Mg = total number of e\(^-\) gained by N
We need 3 Mg atoms to react with 2 N atoms

\[ \text{Mg} \overset{\text{[Ne]}}{\text{[Ne]}} \rightarrow \text{Mg}^{2+} \quad \text{[Ne]} \]
\[ \text{N} \overset{\text{[Ne]}}{\text{[Ne]}} \rightarrow \text{N}^{3-} \quad \text{[Ne]} \]

\[ \text{Mg}^+ \quad \text{N}^3- \quad \text{Mg}\_3\text{N}_2 \]

So \[ 3 \text{Mg} + 2 \text{N} \rightarrow \text{Mg}\_3\text{N}_2 \]
(2) Covalent bonds: involve sharing of one or more electron pairs. Occurs between nonmetals only (similar EN).

1 pair shared electrons (2e⁻): Single bond \( A-B \) or \( A\equiv B \)
2 pairs shared electrons (4e⁻): Double bond \( A= B \) or \( A\equiv B \)
3 pairs shared electrons (6e⁻): Triplet bond \( A≡B \) or \( A\equiv B \)

Examples:

- Lewis dot formula: Each electron expressed by dot.
- Dash formula: Each electron pair is expressed by dash.

\[
\begin{align*}
\text{H}_2 & \quad \text{H} : \text{H} \\
\text{H}_2 \text{F} & \quad \text{H} : \overset{\vdots}{\text{F}} \\
\text{H}_2 \text{O} & \quad \text{H} : \overset{\vdots}{\text{O}}
\end{align*}
\]

The electrons in bond may not be shared equally. If the electronegativities of the 2 elements are different, the more EN element will pull electrons closer to itself, creating polar bond. If they are the same, the bond is non polar.

Polar Bond:
- \(+8 \rightarrow -8\)
- \(+8\) small + charge
- \(-8\) small - charge
- \(\Delta \text{EN} = 1.9\)
- \(\text{H}_2 \text{F}:\) is a polar covalent bond.

Non-Polar Bonds:
- \(\text{Cl} - \overset{\vdots}{\text{Cl}}\)
- \(\text{P} - \overset{\vdots}{\text{As}}\)
- \(\text{C} - \overset{\vdots}{\text{S}}\)

EN 2.1 2.1 2.5 2.5

How do we tell if a molecule is polar? By its geometry!

**Step 1:** Lewis dot formula for molecules and polyatomic ions.

Representative ("A") elements usually achieve noble gas configuration in most compounds. Since all noble gases (but He) have 8e⁻ in outer shell, this is called "octet rule" - a guideline for determining structure.

Example: \(\text{N}_2\)
- \(\text{N}\) is a SA - has 5e⁻ in outer shell; each N wants 8e⁻.
- We have 10e⁻ to play with.
- \(\text{N} = \overset{\vdots}{\text{N}}\) - triplet bond.
Note: Octet Rule: \( S = N - A \) where 
\[ \begin{align*} 
S &= \# \text{shared } e^- \text{ in bonds} \\
N &= \# \text{ needed } e^- \\
A &= \# \text{ available } e^- \text{ (valence)} 
\end{align*} \]

Example 2. \( \text{CH}_4 \) (C is IVA: has 4 e\(^-\) in outer shell 
- each C wants 8 e\(^-\) 
- each H wants 2 e\(^-\) 
- they will share electrons in such a way so that they will get them.

\[
\begin{align*}
\text{H} & \quad \text{H} \\
\text{C} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\end{align*}
\]

\[
S = N - A = (4 \times 8 + 4 \times 2) - (4 + (4 \times 1)) = 16 - 8
\]

Arrangement of elements:
- rules:
  1. \(\text{the element needing the most electrons to fill its octet is usually the central atom.}\)
  2. \(\text{the most symmetrical skeleton is usually correct.}\)
  3. \(\text{oxygen atoms are usually bonded to another nonmetal - not to each other}\)
  4. \(\text{in ternary acids, e.g., H}_2\text{SO}_4, \text{H are bonded to O. (A ternary acid contains H, O and another nonmetal) Always obeys octet rule.}\)

Example 3. \( \text{NH}_3 \) (N is IVA: has 5 e\(^-\) in outer shell 
- each N wants 8 e 
- each H wants 2 e

\[
\begin{align*}
\text{H} & \quad \text{H} \\
\text{N} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\end{align*}
\]

\[
S = N - A = \left[4 \times 8 + (3 \times 2)\right] - \left[1 \times 5 + 3 \times 1\right] = 14 - 8
\]

\( \text{H-N-H} \) 

6 electrons shared.

Example 4. \( \text{NH}_4^+ \) (N is IVA BUT lost 1 electron 
- has 4 e\(^-\) only in outer shell.

\[
\begin{align*}
\text{H} & \quad + \\
\text{H} & \quad \text{H} \\
\text{N} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\end{align*}
\]

\[
S = N - A = \left[(1 \times 8) + (4 \times 2)\right] - \left[1 \times 5 + 4 \times 1 \right] = 16 - 8
\]

\[
\begin{align*}
\left[\text{H-N-H}\right]^+ \\
\text{H} & \quad \text{H} \\
\text{H} & \quad \text{H} \\
\end{align*}
\]

8 electrons shared.
More Rules of Thumb:

(1) Halogens and H always share one electron to complete outer shell.

(2) Oxygen can do several things depending on the molecule
   (a) single bond by sharing an electron
   \[ \text{H}_2\text{O} \quad \text{H} : \overset{\text{O}}{\text{O}} : \quad S = N - A \]
   \[ = (8 + 4) - (6 + 2) \]
   \[ = 12 - 8 = 4 \]
   
   (b) single bond by accepting 2 electrons from another atom and not sharing at all.
   \[ \text{HClO}_3 \quad \text{H} : \overset{\text{O}}{\text{O}} : \overset{\text{Cl}}{\text{Cl}} : \]
   \[ S = N - A \]
   \[ = (4 \times 8 + 2) - (7 + 3 \times 6 + 1) \]
   \[ = 34 - 26 \]
   \[ = 8 \text{ electrons shared.} \]
   
   Recall: in a ternary acid, H must be bonded to an oxygen atom.

   (c) double bond by sharing 2 of its electrons.
   \[ \text{CO}_2 \quad \overset{\text{O}}{\text{O}} : \overset{\text{C}}{\text{C}} : \overset{\text{O}}{\text{O}} : \]
   \[ S = N - A \]
   \[ = (3 \times 8) - (4 + 2 \times 6) \]
   \[ = 24 - 16 \]
   \[ = 8 \text{ electrons shared.} \]
   
   Rarely, exceptions C\text{N}^- + similar compounds. CO always obeys octet rule.

(3) There are never any lone pairs of electrons on a carbon atom.

(4) When forming multiple bonds between atoms (i.e., double + triple bonds), both atoms donate the same number of electrons.

Double bond: \[ \text{H}_2\text{O}, \quad \text{H} : \overset{\text{O}}{\text{O}} : \]

Triple bond: \[ \text{HCC}, \quad \text{H} : \overset{\text{C}}{\text{C}} : \overset{\text{C}}{\text{C}} : \overset{\text{H}}{\text{H}} \]
Note: two notable exceptions to these rules are the heteronuclear diatomic species

\[ \text{CO} : \text{C} \equiv \text{O} \leftrightarrow : \text{C} \equiv \text{O}^2^- \]

\[ \text{CN}^- : \text{C} \equiv \text{N} \]

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

\[ \begin{align*}
\text{[O} \equiv \text{C} \equiv \text{O}]^{-2} & \leftrightarrow \begin{array}{c}
\text{[O} \equiv \text{C} \equiv \text{O}]^{-2}
\end{array} \\
\text{[C} \equiv \text{C} \equiv \text{O}]^{-2}
\end{align*} \]

\[ S = N - A \\
= (4 \times 8) - (3 \times 6 + 1 \times 4 + 2) \\
= 32 - 24 \\
= 8 \]

The "\( \leftrightarrow \)" means that the true structure is somewhere in between the three structures. These 3 structures are called RESONANCE structures.

A molecule or ion for which 2 or more equally acceptable dot formulas are available to describe the bonding is said to exhibit resonance. We can represent \( \text{CO}_3^{2-} \) ion as

\[ \begin{align*}
\text{delocalized structure}
\end{align*} \]

the C=O bonds are neither double nor single bonds but somewhere in between.

- C-O single bond length is \( \sim 1.43 \) Å
- C=O double bond length is \( \sim 1.22 \) Å (shorter)
- C=O in carbonate ion is \( 1.29 \) Å

Exceptions to OCTET Rule: (where \( S = N - A \) does NOT apply).

1. most covalent compounds of Be (IIA)

\[ \text{BeCl}_2 \]

Cl atoms have octet

Be only has 4 e⁻

2. most covalent compounds of Group IIIA

\[ \text{BF}_3 \]

F atoms have octet

B only has 6 e⁻
3. Compounds in which the central element must have a share in more than 8 valence electrons to accommodate all attached atoms.

\[
\begin{align*}
PF_5 & \quad \text{P has 10\text{e}^- around it since each F atom had to have 1 extra e\text{-} to complete its octet.}
\end{align*}
\]

Note: In compounds \& ions, outside atoms ALWAYS obey octet rule. Central atoms may or may not obey octet rule.

4. Compounds containing d\text{-} or f\text{-} transition elements.

5. Species containing odd number of electrons. - free radicals. - very reactive.

\[
\begin{align*}
\text{Nitrogen oxide} & \quad \text{need resonance to explain}
\end{align*}
\]

\[
\begin{align*}
\begin{array}{c}
\text{N} \\
\text{O}
\end{array}
\quad \begin{array}{c}
\text{5 valence e}^- \\
\text{6 valence e}^-
\end{array}
\quad \begin{array}{c}
\text{11 valence e}^-
\end{array}
\end{align*}
\]

\[
\begin{align*}
\text{N} \cdots \text{O} \leftrightarrow \text{N} \cdots \text{O} \leftrightarrow \text{N} \cdots \text{O}
\end{align*}
\]

Now that we can figure out Lewis dot structures, next: Molecular Shapes.

Valence Shell Electron Pair Repulsion Theory (VSEPR) is based on the idea that bonding and lone pairs of electrons in the valence shell of an atom want to be as far apart as possible to gain maximum stability. This results in electron pair geometry. If we are concerned with the actual shape of the molecule or ion, the lone pairs of \text{e}^-\text{s are invisible, but their presence is still felt. In fact, lone pairs of e}^-\text{s take up more space than electrons in bonds. In this class, we are only concerned with "ideal" bond angles.

Let us go through examples, in order.
I. Two (2) regions of high electron density around central atom.
   sp hybridization (180°)
   electronic geometry: linear

   example: BeCl₂

   \[ \begin{array}{c}
     \text{Cl} \\
     \text{Be}
   \end{array} \]

   \[ \begin{array}{c}
     \text{Cl} \\
     \text{Be} \quad \text{Cl}
   \end{array} \]

   Be-Cl bond is polar but the bond polarities cancel.

   molecular geometry: linear

II. Three (3) regions of high electron density around central atom.
   sp² hybridization (120° ideal)
   electronic geometry: trigonal planar

   example: BF₃

   \[ \begin{array}{c}
     \text{F} \\
     \text{F} \\
     \text{F}
   \end{array} \]

   B-F bond is polar but the bond polarities cancel.

   molecular geometry: trigonal planar.

\[ \text{NO}_2^- \]

\[ \begin{array}{c}
   \text{O} \\
   \text{N} \\
   \text{O}
\end{array} \]

Bond polarities do NOT cancel.

molecular geometry: angular

\[ \begin{array}{c}
   \text{N} \\
   \text{O} \\
   \text{O}
\end{array} \]

\[ \begin{array}{c}
   \text{O} \\
   \text{N} \\
   \text{O}
\end{array} \]

\[ \begin{array}{c}
   \text{O} \\
   \text{N} \\
   \text{O}
\end{array} \]

\[ \begin{array}{c}
   \text{O} \\
   \text{N} \\
   \text{O}
\end{array} \]

\[ \begin{array}{c}
   \text{O} \\
   \text{N} \\
   \text{O}
\end{array} \]

EN 3.5  3.0  3.5
\[ \text{P}_\text{π}  \Delta 0 \]

\[ \begin{array}{c}
   \text{O} \\
   \text{N} \\
   \text{O}
\end{array} \]

Bond polarities do NOT cancel.
III. Four (4) regions of high electron density

$sp^3$ hybridization

electronic geometry: tetrahedral (109° 28' ideal)

Examples: CH₄ and CF₄

\[
\begin{align*}
\text{H} & \quad \text{H} & \quad \text{F} & \quad \text{F} \\
\text{H} & \quad \text{C} & \quad \text{H} & \quad \text{F} \\
\text{H} & \quad \text{H} & \quad \text{F} & \quad \text{F} \\
\text{H} & \quad \text{2.1} & \quad \text{F} & \quad \text{4.0}
\end{align*}
\]

C-H bonds are polar
but CH₄ is nonpolar
since polarities cancel

C-F bond is polar
but CF₄ is nonpolar

molecular geometry = electronic geometry = tetrahedral for CH₄, CF₄

\[
\begin{align*}
\text{H} & \quad \text{H} & \quad \text{H} & \quad \text{Cl} \\
\text{H} & \quad \text{C} & \quad \text{Cl} & \quad \text{H} \\
\text{H} & \quad \text{Cl} & \quad \text{H} & \quad \text{H}
\end{align*}
\]

Examples: NH₃

\[
\begin{align*}
\text{H} & \quad \text{N} & \quad \text{H} \\
\text{H} & \quad \text{N} & \quad \text{H} \\
\text{H} & \quad \text{2.1}
\end{align*}
\]

electronic geometry: tetrahedral
molecular geometry: pyramidal

Example: SO₃²⁻

\[
\begin{align*}
\text{O} & \quad \text{S} & \quad \text{O} \\
\text{O} & \quad \text{O} & \quad \text{O}
\end{align*}
\]

electronic geometry: tetrahedral
molecular geometry: pyramidal

POLAR
Examples:

$H_2O$

$$\begin{align*}
\text{H} & \quad \text{O} \\
\text{H} & \quad \text{H}
\end{align*}$$

Electronic geometry: tetrahedral
Molecular geometry: angular
Polar

IV. Five (5) regions of high electron density (does not obey octet rule)

$\text{sp}^3d$ hybridization

Electronic geometry: trigonal bipyramidal

Example:

$\text{PF}_5$

(does not obey octet rule)

Electronic geometry = molecular geometry = trigonal bipyramidal

Example: $\text{SF}_4$

(I lone pair of electrons on central atom)

Electronic geometry = trigonal bipyramidal
Molecular geometry: see saw
Molecule is polar
Example: 2 lone pairs of electrons on central atom

\[
\text{BrF}_3 \\
\overset{VIIA}{\text{Br}} \overset{VIIA}{\text{F}}
\]

\[
: \text{F} - \text{Br} - : \text{F} \\
\overset{\text{Br}}{\text{F}} \overset{\text{F}}{\text{Br}} \overset{\text{F}}{\text{F}}
\]

Electronic geometry: trigonal bipyramidal  
Molecular geometry: T-shaped and molecule is polar

Example: 3 lone pairs of electrons on central atom

\[
\text{ICl}_2^- \\
\overset{\text{VIIA}}{\text{I}} \overset{\text{VIIA}}{\text{Cl}}
\]

With 1 extra electron

\[
\begin{bmatrix}
\overset{\text{Cl}}{\text{Cl}} \\
\overset{\text{Cl}}{\text{I}} \\
\overset{\text{I}}{\text{Cl}}
\end{bmatrix}
\]

Electronic geometry: trigonal bipyramidal  
Ionic geometry (since this is an ion): linear + NONPOLAR

V. Six (6) Regions of High Electron Density (does not obey octet rule)  
\(\text{sp}^3\text{d}^2\) hybridization  
Electronic geometry: octahedral (90°, 180° bond angle)

Example: No (0) lone pairs of electrons on central atom

\[
\text{SF}_6 \\
\overset{\text{F}}{\text{S}} \overset{\text{F}}{\text{F}} \overset{\text{F}}{\text{F}}
\]

Electronic geometry = Molecular geometry = octahedral  
Molecule is nonpolar
Example: 1 lone pair of electrons on central atom

\[ \text{IF}_5^- \]

VIIA VIIA

- Electronic geometry: octahedral
- Molecular geometry: square pyramidal; molecule is polar.

Example: 2 lone pairs of electrons on central atom

\[ \text{XeF}_4 \]

- Electronic geometry: octahedral
- Molecular geometry: square planar; molecule is nonpolar.

Note: In discussing bond angles, we have only considered IDEAL bond angles. For instance, in a molecule with tetrahedral electronic geometry, the bond angle is 109°28' ideally.

In reality, lone pairs of electrons occupy more space than bonding pairs with the following result: a decrease in the angles between bonding pairs.
<table>
<thead>
<tr>
<th>Regions of High Electron Density Around Central Atom</th>
<th>Hybridization (Ch10)</th>
<th>Electron Pair (Electronic) Geometry</th>
<th>No. of Lone Pairs of Electrons Around Central Atom</th>
<th>Molecular (Ionic) Geometry</th>
<th>Examples</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>sp</td>
<td>linear</td>
<td>0</td>
<td>linear</td>
<td>BeCl₂, CO₂</td>
</tr>
<tr>
<td>3</td>
<td>sp²</td>
<td>trigonal planar</td>
<td>0</td>
<td>bent or angular</td>
<td>BF₃, CO₃²⁻</td>
</tr>
<tr>
<td>4</td>
<td>sp³</td>
<td>tetrahedral</td>
<td>0</td>
<td>tetrahedral</td>
<td>NO₂⁻</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>1</td>
<td>pyramidal</td>
<td>HClO₄</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>2</td>
<td>bent or angular</td>
<td>CH₄, CH₃Cl, NH₄⁺</td>
</tr>
<tr>
<td>5</td>
<td>sp³d² or dsp³</td>
<td>trigonal bipyramidal</td>
<td>0</td>
<td>trig. bipy. see saw T-shaped</td>
<td>NH₃, SO₃²⁻, H₃O⁺</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>1</td>
<td>linear</td>
<td>H₂O</td>
</tr>
<tr>
<td>6</td>
<td>sp³d² or dsp³</td>
<td>octahedral</td>
<td>0</td>
<td>octahedral</td>
<td>PF₅, SF₄</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>1</td>
<td>square planar</td>
<td>BrF₅, ClF₃</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>2</td>
<td></td>
<td>I⁻, XeF₂</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>SF₆</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>IF₅⁻, BrF₅</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>IF₄⁻, XeF₄</td>
</tr>
</tbody>
</table>
Formal Charges

The concept of formal charges helps us choose the correct Lewis structure for a molecule.

Rules for assigning formal charges to Group A elements:
1. For a molecule, the sum of the formal charges for all elements is zero.
   For an ion, the sum of the formal charges for all the elements equals charge on ion.
2. Formal charge, $\text{FC} = \text{Group No.} - (\text{no. of bonds} + \text{no. unshared e}^-)$
   Note: Group Number of noble gases = VII; double bond counts as 2 bonds.
3. A good Lewis dot structure has
   - low or zero values of FC for all elements
   - adjacent atoms are not given FC values of the same sign
   - the more electronegative atom in a bond is given the more negative FC.

Consider two Lewis dot structures for sulfuric acid, $\text{H}_2\text{SO}_4$.

(a) $\text{H} : \text{O} : \text{S} : \text{O} : \text{H}$

   $O_{(2)}$  

   FC  
   $\text{H} = 1 - (1 + 0) = 0$  
   $O_{(1)} = 6 - (2+4) = 0$  
   $O_{(2)} = 6 - (1+6) = -1$  
   $S = 6 - (4+0) = +2$

   This is an OK dot structure, but there might be a better one where FC of $S$ is closer to zero. Consider

(b) $\text{H} : \text{O} : \text{S} : \text{O} : \text{H}$

   $O_{(2)}$  

   Note: sulfur does not obey the octet rule

   FC  
   $\text{H} = 1 - (1 + 0) = 0$  
   $O_{(1)} = 6 - (2+4) = 0$  
   $O_{(2)} = 6 - (2+4) = 0$  
   $S = 6 - (6-0) = 0$

   Structure is a little strange, but all elements have zero FC. Good!

There is data to support this structure in $\text{H}_2\text{SO}_4$. 
**Bond properties**

**Bond order** - number of bonding electron pairs shared by 2 atoms

<table>
<thead>
<tr>
<th>Bond Order</th>
<th>bond</th>
<th>Note: can be fractional eq 1.5</th>
<th>(between single &amp; double bond)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>single</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2</td>
<td>double</td>
<td></td>
<td></td>
</tr>
<tr>
<td>3</td>
<td>triple</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Bond length** - distance between nuclei of 2 bonded atoms

- **units**: Å $1 \times 10^{-10}$ m
- hm $1 \times 10^{-9}$ m
- pm $1 \times 10^{-12}$ m

**In general**

- single bonds $>$ double bonds $>$ triple bonds
- longest weakest bond
- shortest strongest bond
- (smallest bond energy) (largest bond energy)

**Bond energy** -

- Bond dissociation energy, $D$, is enthalpy change ($^{\Delta}H$) for breaking a bond in a gaseous molecule; process is exothermic.
- You can use information to estimate $^{\Delta}H_{\text{rxn}}$.

**Example**: Calculate $^{\Delta}H_{\text{rxn}}$ for $\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow 2\text{HCl}(\text{g})$

**Using bond energies**:

- H–H 436 kJ/mol
- H–Cl 432 kJ/mol
- Cl–Cl 242 kJ/mol

**Note**: it takes energy to break bonds; energy is released when making bonds.

**Reaction**

\[ \text{H–H (g)} + \text{Cl–Cl (g)} \rightarrow \text{H–Cl (g)} + \text{H–Cl (g)} \]

\[ ^{\Delta}H_{\text{rxn}} = D_{\text{H}_2} + D_{\text{Cl}_2} - 2 \times D_{\text{HCl}} \]

\[ = +436 \text{ kJ/mol} + 242 \text{ kJ/mol} - 2 \times 432 \text{ kJ/mol} \]

\[ = -186 \text{ kJ} \]

**Note**: $^{\Delta}H_{\text{rxn}} = 2 \times ^{\Delta}H_f [\text{HCl (g)}] = -184.6 \text{ kJ} \quad \text{very close!!} \]