**Vocabulary:**
1. Ionic bonds
2. Covalent bonds
3. Molecule
4. Metallic bond
5. Octet rule
6. VSEPR
7. Single bond
8. Double bond
9. Triple bond
10. Polar
11. Nonpolar
12. Hydrogen bonding
13. Dipole-dipole force
14. London dispersion force
15. Hybridization
16. Intermolecular force (IMF)

**Formulas/Constants:** Memorize and be able to use VSEPR chart.

**Objectives:**
- Be able to describe and identify the three types of bonds: ionic, molecular, and metallic.
- Be able to describe characteristics of ionic, molecular, and metallic compounds.
- Be able to describe the difference between polar and nonpolar molecular bonds.
- Be able to determine bond type using the periodic table.
- Be able to describe single, double and triple bonds; including length and strength.
- Be able to draw Lewis structures for various compounds and polyatomic ions.
- Be able to apply the VSEPR theory in determining molecular geometry and polarity.

**VSEPR & Molecular Geometry**

<table>
<thead>
<tr>
<th>Molecular Shape</th>
<th>Usually Polar or Non Polar??</th>
<th>Atoms Bonded to Central Atom</th>
<th>Lone Pairs of e- on Central Atom</th>
</tr>
</thead>
<tbody>
<tr>
<td>Linear</td>
<td>2 surrounding atoms; no lone e-s</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Bent</td>
<td>2 surrounding atoms – 1 pair of lone e-s</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Trigonal Planar</td>
<td>3 surrounding atoms; no lone e-s</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Tetrahedral</td>
<td>4 surrounding atoms; no lone e-s</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Trigonal Pyramidal</td>
<td>3 surrounding atoms – 1 pair of lone e-</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Bent</td>
<td>2 surrounding atoms – 2 pairs of lone e-</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
TYPES OF BONDING review sheet fill in blanks!

3 types of BONDS

Covalent
(_______e-)
Between a _______ and __________
Properties
Strength :_____
MPt, BPt:
Appearance:
Conductors?

Ionic
(_______of e-)
Between a _______ and __________
OR ________________
Properties
Strength:_____
MPt, BPt:
Appearance:
Conductors?
3 examples

Metallic
(_______ of electrons)
Between a _______ and __________
Properties
Malleable=_____________
Ductile = _______________
Appearance:
Conductors?
3 examples

Non-Polar =
3 examples:

Polar =
3 examples

All three bonds types follow the octet rule:
MOLECULAR STRUCTURES:
Review: Molecular = _________________________ bond.
Octet rule = ______________________________________________________________________
H and He are exceptions to octet rule – WHY? _____________________________________________________________________________

LEWIS STRUCTURE:
Symbol dots dashes

Single bond =

Steps for Lewis Structure Problem (and an example)
Example: draw the Lewis Structure of iodomethane - CH₃I

1. Determine the number and type of atoms
2. Determine the total number of valence electrons in the atoms to be combined. (use periodic table)
3. Arrange the atoms to form a skeletal structure. (First atom is central unless otherwise specified or H is first. Hydrogen can never be the central atom)
4. Add unshared pairs of electrons so that each atom is surrounded by 8 electrons (octet rule) (except H – it can only hold 2 e-s)
5. Count the number of e-s in the structure to be sure the number of valence electrons used equals the number available.

Practice: together: ammonia, NH₃ and silicon tetrafluoride

You try: BrI, CH₃Br

You try 2 water
DOUBLE BOND =

TRIPLE BOND =

Comparison of bond strength and length: shortest & strongest = ________________________________
Longest & weakest = ________________________________

Use multiple bonds when there are not enough valence electrons to complete octets by adding unshared electron pairs (step 5 = numbers do not match)

5b. If too many electrons have been used, subtract two unshared pairs, changing it into a shared pair until the total # of electrons is correct.

Ex: (together) CH₂O

Practice: carbon dioxide and HCN

POLYATOMIC IONS:

Additional Step: when counting valence electrons must include the charge:

- If positive subtract electrons
- If negative add electrons
- Put brackets with a charge around the final structure

Ex: nitrate Practice: sulfate
LEWIS STRUCTURES: Draw the following Lewis Structures

1. TeCl₂
2. NO₃⁻
3. phosphate
4. I₂
5. ICl
6. H₂S
7. nitrite
8. carbonate
9. water
10. P₂
11. PCl₃
12. SF₂
13. bromine monochloride
14. ammonium
HYBRIDIZATION: Mixing of two or more atomic orbitals of similar energy to make new orbitals of equal energies. This allows the sharing of electrons in some covalent bonds to occur.

Ex. For Carbon, draw the hybridization and name the new hybrid!

You try: predict the hybridization of the molecules on the Lewis Structure worksheet on pg 5

VSEPR Theory – valence-shell, electron pair repulsion
- Way to predict molecular geometry (shape)
- There is a repulsion between valence e- pairs

Steps to work problems:
(Be sure to refer to the VSEPR table on the same handout as your vocabulary – first six shapes)

1) Draw Lewis Structure of Molecule
2) Put molecule in AB₁E_z form, where:
   - A represents the central atom
   - B represents the atoms bonded to A (y is the # of B atoms)
   - E represents the lone pair e- on A (z is the # of lone pairs)
3) Predict the shape based on the AB₁E_z form

Example 1: Use VSEPR Theory to predict the shape of sulfur difluoride (SF₂)

Example 2: CBr₄.

You try: sulfite

You try: predict the VSEPR shapes of the molecules on the Lewis Structure worksheet on pg 5
The polarity of molecules

Many of the properties of a molecule are directly related to the shape and from there the polarity. For example, water is polar and is a liquid at room temperature. Larger molecules like carbon dioxide, CO₂ – which is nonpolar, are gasses. To determine polarity, you must now bring several things together. You must be able to:

1. Find the Lewis structure and VSEPR shape of the molecule.
2. Determine if the molecule is symmetrical or not

If molecules are symmetrical: they are usually non-polar (if the non-central atoms are all the same, so they can “cancel” each other out).

If a molecule is not symmetrical – (the polarities do not “cancel” and) the overall molecule is polar.

Symmetrical VSEPR shapes –typically nonpolar (If all surrounding atoms are the same): linear, trigonal planar, tetrahedral. These 3 shapes are nonpolar as long as the surrounding atoms match – if they don’t match then the molecule is polar.

Asymmetrical VSEPR shapes- always polar: bent (both), trigonal pyramidal.
Examples together: draw the VSEPR shape first then determine polarity

1) water:

2) carbon dioxide

3) HCN

You try: determine the molecule polarity for all of the molecules on your VSEPR worksheet, pg5

INTERMOLECULAR FORCES and Strengths

NOTE: Ionic compounds are held together by opposite charges – they are the strongest force, IMFs only apply to molecules (2 or more nonmetals in a compound!)

Intermolecular forces are the forces between molecules that hold the molecules together. For example keeping all the water molecules in a glass of water “stuck together”. These forces are also called Van der Waals forces. They are generally weaker forces than bonds.

There are three types of IMFs.

Hydrogen Bonding: Hydrogen attracted to a neighboring N, O, or F. Not bonded to these elements! This is the strongest intermolecular force.

Dipole-dipole force: Mainly present in polar molecules that don’t have hydrogen bonding IMFs present.

London Dispersion: Only force in non-polar molecules. They are present in polar but basically irrelevant due to other stronger IMFs. This is the weakest intermolecular force.

ALL IMFs are weaker than all bonds. See scale below.

Relative strength of bonds/IMF
Strongest force ------------------------------------------weakest force

Bonding intermolecular forces
Ionic cmpds > *molecular cmpds > > > > > H bonding > dipole-dipole > dispersion forces

Examples: Let’s determine IMFs for the three compounds on the top of this page.

You try: determine the Intermolecular forces (IMFs) present for all of the molecules on worksheet pg5
Intermolecular forces (IMF) practice sheet
Use the IMF notes that are in your packet to fill in the following:

1. Define intermolecular forces: ____________________________________________

2. Fill in the following table describing the 3 types of IMFs

<table>
<thead>
<tr>
<th>Name</th>
<th>Relative Strength/boiling &amp; melting pt</th>
<th>Description/types of atoms</th>
<th>examples</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen bonding</td>
<td>Strongest IMF Higher bpt &amp; mpt</td>
<td>Attraction between H(in a polar molecule) and F, O, or N/H and F, O or N</td>
<td></td>
</tr>
<tr>
<td>Dipole-dipole</td>
<td></td>
<td>Attraction between 2 polar molecules/usually 2 different nonmetals (usually results in gases)</td>
<td></td>
</tr>
<tr>
<td>Dispersion force</td>
<td></td>
<td>Attraction between 2 nonpolar molecules/usually 2 of the same nonmetal or a nonpolar VSEPR shape (results in gases)</td>
<td>Chlorine gas, Cl₂</td>
</tr>
</tbody>
</table>

3. What is the intermolecular force present for each of the following compounds?
   water ___________________________ phosphorus trichloride PCl₃ ___________________________
   nitrogen N₂ ___________________________ Hydrobromic acid (HBr) ___________________________
   ammonia NH₃ ___________________________

4. Place the above: H₂O, N₂, NH₃, PCl₃ and HBr into the following categories:

<table>
<thead>
<tr>
<th>Strongest force/highest bpt/mpt</th>
<th>Medium force/medium bpt/mpt</th>
<th>Weakest force/lowest bpt/mpt</th>
</tr>
</thead>
</table>

5. Complete the table for substance A and B:

<table>
<thead>
<tr>
<th>A</th>
<th>B</th>
</tr>
</thead>
<tbody>
<tr>
<td>Freezing Point = -78°C</td>
<td>Freezing Point = 125°C</td>
</tr>
<tr>
<td>Boiling Point = -6°C</td>
<td>Boiling Point = 338°C</td>
</tr>
<tr>
<td>State of matter at -25°C =</td>
<td>State of matter at 0°C =</td>
</tr>
<tr>
<td>State of matter at 0°C =</td>
<td>State of matter at 100°C =</td>
</tr>
<tr>
<td>State of matter at 100°C =</td>
<td>State of matter at 200°C =</td>
</tr>
</tbody>
</table>

Which substance has the greater intermolecular attractions? (A or B)
1. Fill in the following table:

<table>
<thead>
<tr>
<th>Bond Type</th>
<th>Type of atoms</th>
<th>One Property</th>
<th>One example</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
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<tr>
<td></td>
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<td></td>
</tr>
</tbody>
</table>

2. Explain why the oxidation number of Magnesium is 2+ in ionic compounds (It is recommended you use an electron configuration diagram in your explanation)

3. What type of bond is found in lead (II) phosphate?

4. Draw the Lewis structure for the following then, identify their VSEPR shape and molecular polarity

   PCl₃ and Carbonate.

Answers:

1. Refer to front page with exact same table!
2. Mg electron configuration = 1s² 2s² 2p⁶ 3s² – has 2 valence electrons – can gain 6 e-s or lose 2 e-s to have 8 (octet rule = 8 e-s full outer level = very stable) Mg will lose its 2 val electrons resulting in a positive 2 charge.
3. The phosphate (between the Os and P) contains covalent bonds – between the lead and the phosphate is an ionic bond.
4. P in the middle, 3 single bonded Cls, 2 dots on P, 6 dots on each Cl = AB₃E₁ = trigonal pyramidal
   Polar

C in the middle, 2 single bonded Os with 6 dots and a double bonded O with 4 dots brackets with a negative 2 at the top right = AB₃ = trigonal planar and nonpolar