Unit 1 – Chemical Bonding

In this unit students will closely examine the nature of ionic and molecular bonding. We will study the 3D shapes of molecules and answer questions such as:

Why are ionic compounds (such as table salt – sodium chloride) always solids at room temperature?

Why is water a liquid at room temperature, but carbon dioxide a gas?

What forces holds a water droplet together?

There are 2 main outcomes that make up this unit. Student achievement in each outcome will be formatively (not included in mark) assessed by in class work and hands on activities. Student achievement in each outcome will be summatively (included in mark) assessed by 1 assignment and 1 quiz. Both outcomes will also assessed by a unit test at the end of the unit.

Outcome 1: Bonding Basics and Ionic Compounds (30 %)

Topic 1: Bonding Vocabulary and Lewis Diagrams (structures)
Topic 2: Types of Bonding and Ionic Compounds
Topic 3: Reduction and Oxidation

Outcome 2: Molecular Compounds and Intermolecular Forces (70%)

Topic 1: Lewis and Structural Diagrams for Molecular Compounds
Topic 2: Shape Diagrams for Molecular Compounds
Topic 3: Polarity
Topic 4: Intermolecular Forces and Boiling Points of Molecules

Unit Timeline:

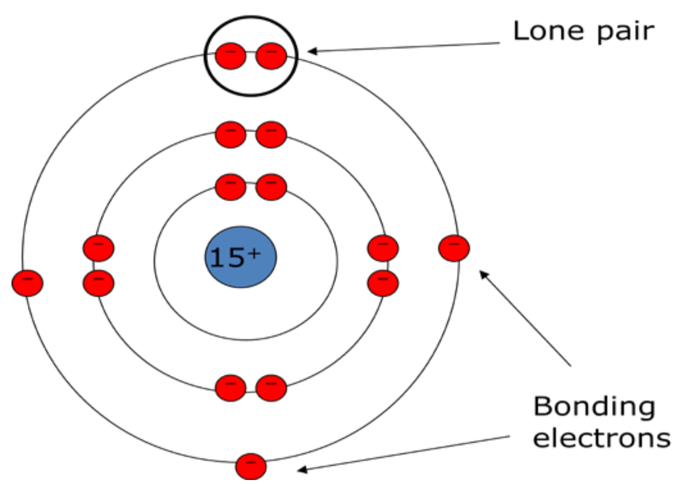
DAY	TOPIC	INCLASS/HOMEWORK
	Outcome 1: Topic 1	Practice Sheet 1
	Topic 2	Practice Sheet 2
	Topic 3	Outcome 1 - Assignment
	Outcome 2: Topic 1	Practice Sheet 3 - front
	Quiz – Outcome 1	
	Topic 1	Practice Sheet 3 - back
	Topic 2	Practice Sheet 4
	Topic 2	
	Topic 2 – Hands on	
	Topic 3	Practice Sheet 5
	Topic 3	
	Topic 4	Practice Sheet 6
	Quiz – Outcome 2	
	Review	
	Test	

Outcome 1 – Bonding Basics and Ionic Compounds

Topic 1: Bonding Vocabulary and Lewis Diagrams (structures)

- **Energy level** – level of energy around an atom where electrons are found
    - The period # gives the # of energy levels of an atom
    - 1st level holds 2 e^-, 2nd and 3rd holds $8e^-$
- **Orbital** – region of space around an atom's nucleus where an electron may exist
- **Valence Orbital** – orbitals found in the outer most energy level of an atom
    - Electrons in the valence orbital are called ____.

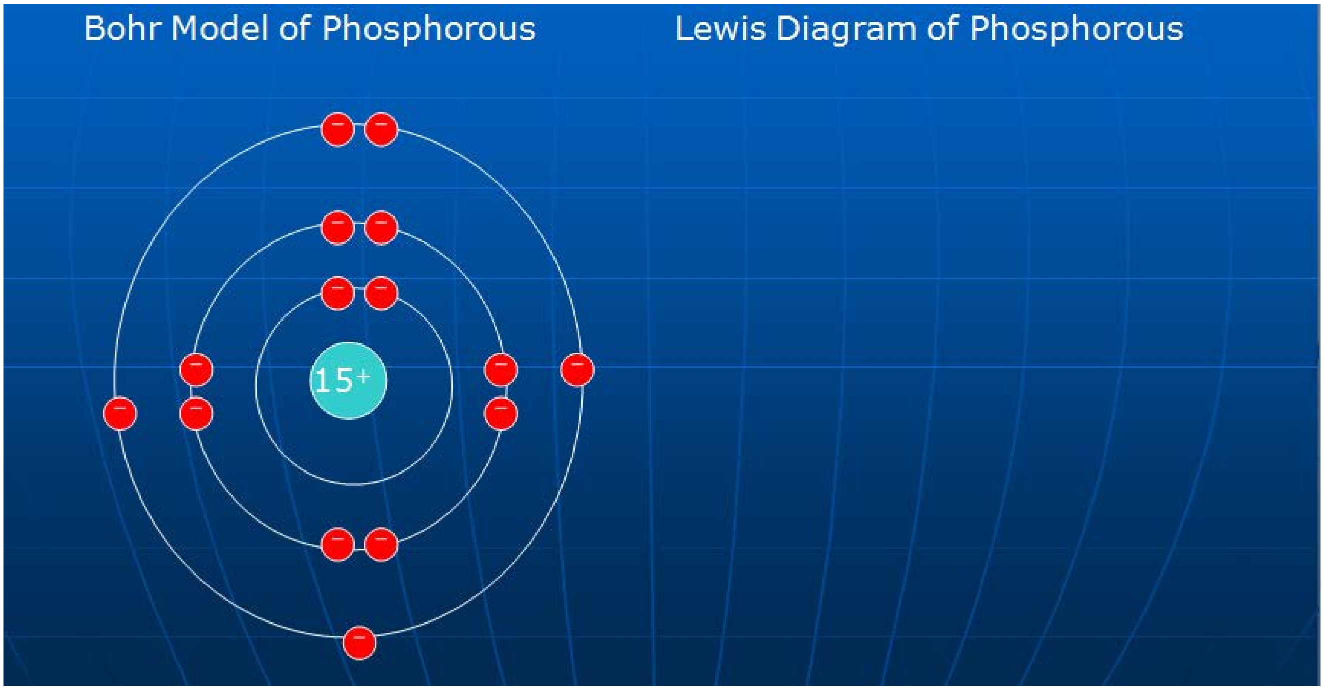
- **Bonding electron** –
    - Bonding electrons are involved in bonding.
- **Lone pair** –



- **Electronegativity** –
  - Atoms have different abilities to attract valence e^-
  - Farther away that e^- are from the nucleus, the weaker they are held by the atom
  - Values are assigned from 0 to 4
  - Highest electronegativity is Fluorine at 4.0
  - Lowest electronegativity is Francium at 0.7

- **Octet Rule** –
  - All atoms want to achieve a full valence energy level
  - Atoms bond to achieve a full valence energy level
  - When metals react they lose electrons to make this happen
  - Non-metals react and gain electrons to make this happen

- **Lewis Diagrams** –
  - Write element symbol to represent nucleus
  - Add a dot to represent each valence electron
  - Start by placing valence electrons into each of four valence orbitals
  - Once each orbital is half-filled start putting second electrons in each orbital



Practice Sheet 1

1. Read pgs 78-79 and answer the following question

Place the following chemistry concepts in the order in which they were created:

(a) Lewis symbols (b) Kekulé structures (c) empirical formulas
(d) quantum mechanics (e) Dalton atom

2. Write the Lewis formula, the electron energy-level diagram, and the electronegativitiy value for each of the following atoms:

(a) aluminium (b) calcium (c) chlorine (d) argon

3. (a) Draw the Lewis symbol for a calcium atom, but omit the two dots for valence electrons. Show that the remaining structure has a double positive charge by enclosing the Lewis structure in large square brackets and writing the overall charge to the upper right side, outside the brackets.

(b) What structure is represented by your symbol? What is it called?

4. In a Lewis symbol of a potassium atom, describe what entities are assumed to be included in its element symbol.

5. Draw a Lewis Diagram for oxygen and label the following: lone pair, bonding electron.

6. List 3 ways that Lewis diagrams for metal atoms differ from Lewis diagrams for non-metal atoms.

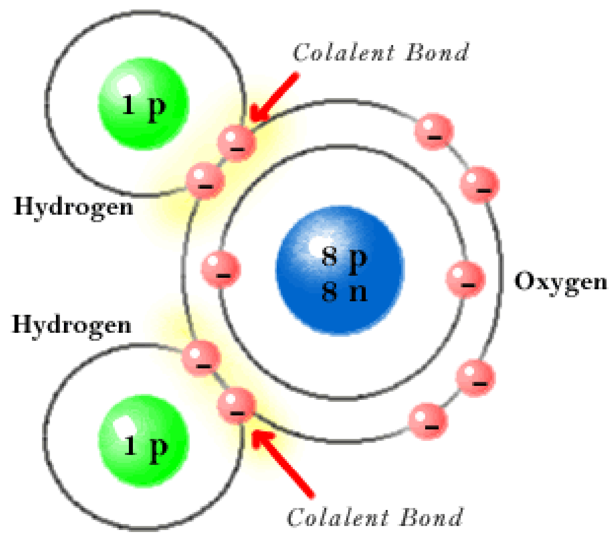
Topic 2: Types of Bonding and Ionic Compounds

- Bonding is the result of two or more atoms in a tug of war for electrons
    - Who wins??

- By looking at the electronegativity of atoms we can predict the kind of bonding that will take place.

Covalent Bonding –

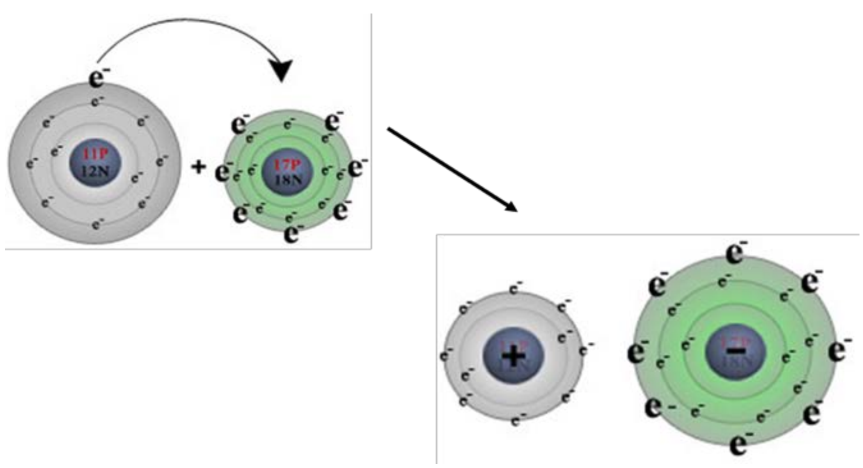
- If electronegativities of both atoms are **high**, neither atoms wins tug of war and they share the electron



Bohr Model of H_2O

Ionic Bonding –

- If electronegativities are **different (>2)** then the atom with the higher will steal the electron from the other one



Animation

http://bcs.whfreeman.com/thelifewire/content/chp02/02020.html

Ionic compounds are:

- **__________________**, which means they have full valence energy levels
- **_____________________** which means they are made up of charged particles
- Formed from the collision of **metal and non-metal** atoms.

Lewis structures for Ionic Compounds

- We can show formation of ionic comp using Lewis Structures

Ex. lithium chloride

Ex. calcium chloride

You try:

sodium oxide

aluminum sulphide

Practice Sheet 2

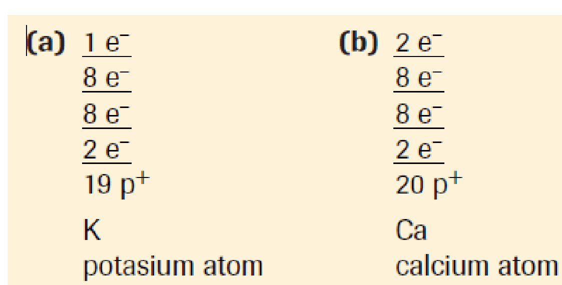
1. Use bonding theory to describe the following in terms of electrons and orbitals: bonding electron, lone pair.

2. Fill in the chart below for all the elements in period 3.

Symbol	Electronegativity	Group #	# of valence e-	Lewis Symbol	# of bonding e-	# of lone pairs

3. Using the electronegativity data in the periodic table, describe the variation in electronegativities within a group and a period.

4. Potassium and calcium both have valence electrons in the fourth energy level, presumably about the same distance from their nuclei, yet calcium has higher electronegativity. Why?



5. Potassium chloride is a substitute for table salt for people who need to reduce their intake of sodium ions. Use Lewis formulas to represent the formation of potassium chloride from its elements.

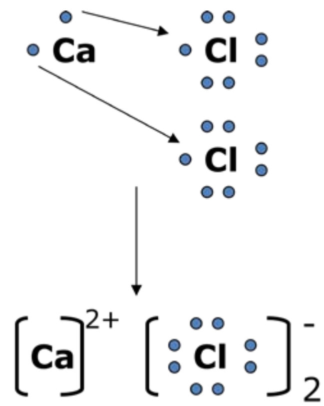
6. Use Lewis formulas to represent the reaction of calcium and oxygen atoms. Name the ionic product.

7. The empirically determined chemical formula for magnesium chloride is MgCl2. Create Lewis formulas to explain the empirical formula of magnesium chloride.

8. Create Lewis formulas to predict the chemical formula of the product of the reaction of aluminium and oxygen.

9. Based only on differences in electronegativity, what compound would you expect to be the most strongly ionic of all binary compounds?

Topic 3: Reduction and Oxidation



- When **<u>simple ionic compounds</u>** form from their elements, there is always a loss and a gain of electrons by the atoms involved.

- We use **<u>half reactions</u>** to show the loss and the gain of electrons separately.

- 2 types of half reactionOxidation
    -
    -

- **<u>Oxidation</u>** is the____.

Examples:

$Na_{(s)} \rightarrow$

$Ca_{(s)} \rightarrow$

$Li_{(s)} \rightarrow$

$Al_{(s)} \rightarrow$

$Zn_{(s)} \rightarrow$

- **<u>Reduction</u>** is the ____.

Examples:

$I_2 \quad + \quad 2e^- \quad \rightarrow \quad 2\ I^-$

$P_4 \quad + \quad 12\ e^- \quad \rightarrow \quad 4\ P^{3-}$

$S_{8(s)} \quad \rightarrow$

$Cl_{2(g)} \quad \rightarrow$

$N_{2(g)} \quad \rightarrow$

$O_{2(g)} \quad \rightarrow$

- Remember LEO says GER

- Loss of electrons is oxidation, gain of electrons is reduction

Net Reactions:

- A net reaction combines the reduction and oxidation half reactions together.
- e- lost in the oxidation half reaction must equal the e- gained in the reduction half reaction

Example: Determine the net reaction of sodium and chlorine

Ox:

Red:

Net:

Example: Determine the net reaction of magnesium and nitrogen

Ox:

Red:

Net:

Example: Determine the net reaction of aluminum and sulfur

Ox:

Red:

Net:

Outcome 2: Molecular Compounds and Intermolecular Forces

Topic 1: Lewis and Structural Diagrams for Molecular Compounds

Molecular (covalent) Bonding

- Occurs between non-metal atoms

- If the atoms are the same the result is a molecular element.
    - Eg. $N_{2(g)}$, $O_{2(g)}$, $P_{4(s)}$

- If the atoms are different, the result is a molecular compound.
    - E.g. $CO_{2(g)}$, $H_2O_{(l)}$, $N_2O_{(g)}$

Animation
http://bcs.whfreeman.com/thelifewire/content/chp02/02020.html

Types of formulas:

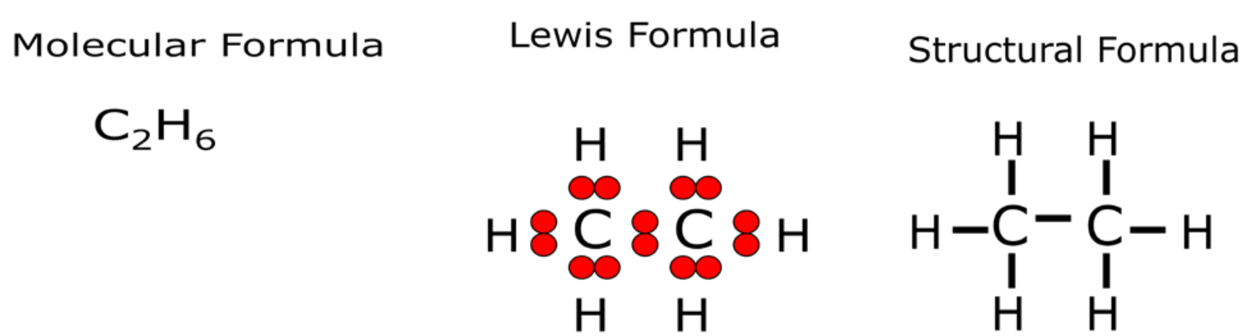
Molecular formula – shows all atoms in the compound, often in the order that they are bonded

Lewis formula – uses Lewis symbols to show bonding

Structural formula – shows which atoms are bonded using lines between atom symbols

Sterechemical formula – represents the 3D shape of the compound

Example - ethane



Stereochemical formula

Bonding Capacity

- (Def) –
- Equals the number of bonding electrons that an atom has
- What is the bonding capacity of:

Cl N O P C

Creating the Lewis Diagram for Molecular Compounds

- Count total valence e^- in the compound by adding the valence e^- for each atom.
  - Eg. SO_2 –
    - S has 6, O has (6 x 2 atoms) 12
    - 6 + 12 = 18 valence electrons
  - You try – NH_3 –
    - N has 5, H has (1 x 3 atoms)
    - 5 + 3 = 8 valence electrons
- Arrange peripheral atoms around central atom
  - Central atom is one with highest bonding capacity
- Place one pair of electrons between each peripheral atom and the central atom
- Place more pairs of electrons on all the peripheral atoms to complete their octets.
- Place any remaining valence electrons on the central atom as lone pairs

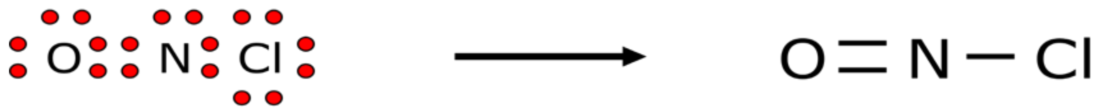
Example: NH_3

- If the central atoms octet is not complete move a lone pair from a peripheral atom to a position between it and the central atom (keep bonding capacity in mind)

Example: NOCl

Structural Diagrams

- To create a structural diagram
  - First start with the Lewis Diagram
  - Replace every pair of <u>shared</u> electrons with a single line
  - Don't show lone pairs



Examples: Draw the Lewis and Structural Diagrams for the following compounds

CCl4 OF_2 C_2H_4

Practice Sheet 3

1. Use bonding theory to draw Lewis formulas for the elements in the halogen family. How are these diagrams consistent with the concept of a chemical family?

2. Use a Lewis formula to explain the molecular formula for nitrogen, $N_2(g)$. Recall that N atoms always obey the octet rule.

3. Use both Lewis formulas and structural formulas to represent molecules of the following compounds:

(a) $CS_2(l)$, carbon disulfide

(b) $H_2Se(g)$ hydrogen selenide

(c) $PH_3(g)$, phosphine

(d) $CH_3SH(g)$, methanethiol

(e) $H_2S(g)$, hydrogen sulfide

4. Why is it incorrect to write the structural formula of the H2S molecule as H—H—S?

5. Why is the molecular formula for the methanol molecule usually written as CH3OH instead of CH4O? (Do the Lewis diagram to help you answer this)

6. For each of the following molecular compounds, name the compound, and explain the empirically determined formula by drawing a Lewis formula and a structural formula:

(a) HCl

(b) CO2

7. Use the bonding capacities listed in Table 1 (page 87) to draw a structural formula of each of the following entities. In each case (for these particular molecules), every C must connect to 4 lines, every N to 3 lines, every O to 2 lines, and every H to one line because C, N, and O atoms always obey the octet rule. Hint: Use the sequence of atoms in some of these molecular formulas to guide you.

(a) H_2O_2 (b) C_2H_4 (c) HCN

(d) C_2H_5OH (e) CH_3OCH_3 (f) CH_3NH_2

8. Using Lewis symbols, predict the simplest binary molecular compound and write the chemical name for a product of each of the following reactions. Include a structural formula for a molecule of the product.

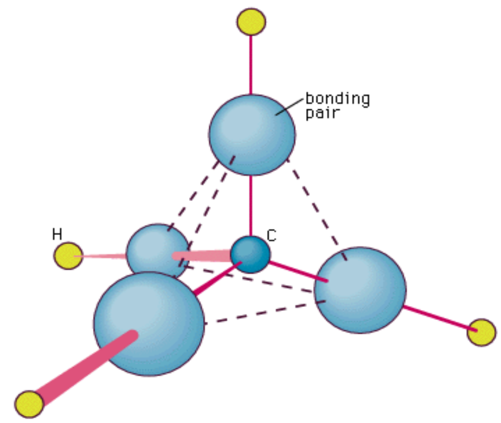
(a) $I_2(s) + Br_2(l)$ -->

(b) $P_4(s) + Cl_2(g)$ -->

(c) $O_2(g) + Cl_2(g)$ -->

(d) $C(s) + S_8(s)$ -->

Topic 2: Shape Diagrams for Molecular Compounds



Stereochemistry –

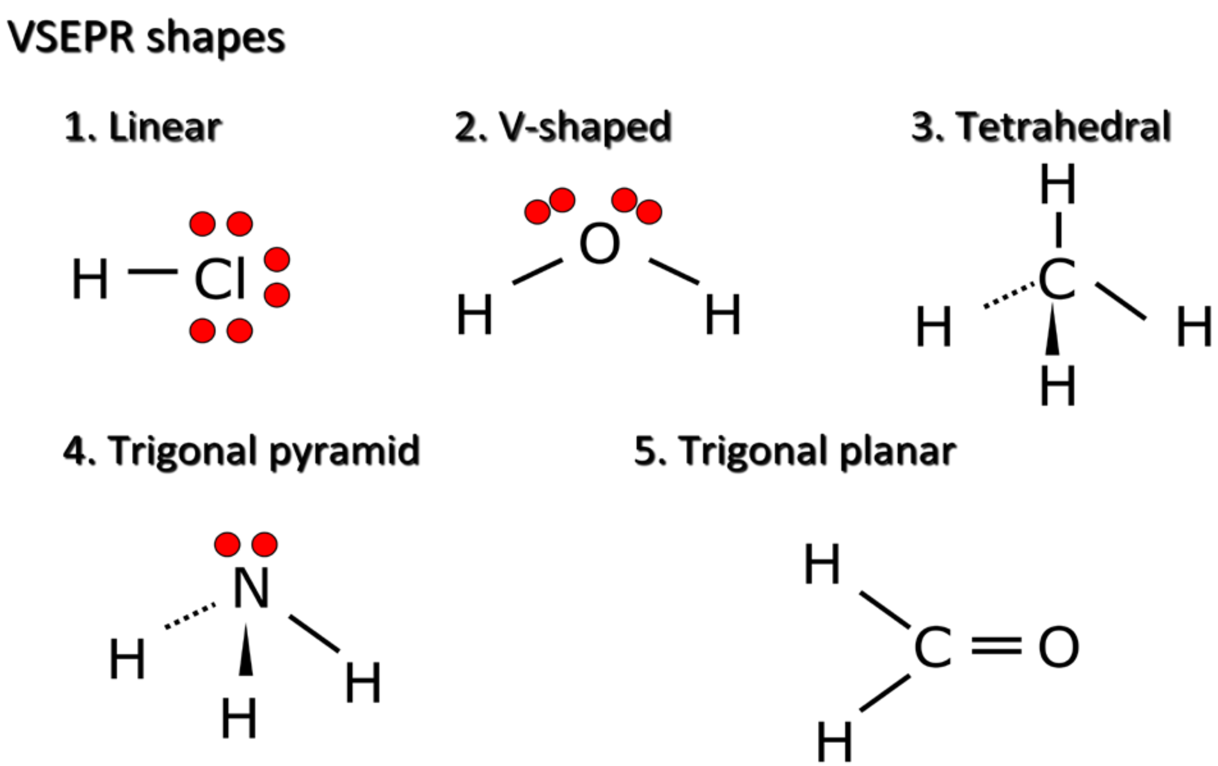
- study of 3D configuration of molecules

VSEPR theory

-
- Pairs of electron stay as far away from each other as possible

According to VSPER Theory

- Only valence electrons on the central atom are important for molecular shape
- Valence electrons are paired.
- Bonded electron pairs and lone pairs are equal
- Electron pairs repel each other
- Molecular Shape is determined when e^- pairs are a max distance apart



Patterns in VSEPR Theory

- Linear if:
  - Only two atoms involved or;
  - Central atom w/ no lone pairs bonded to 2 atoms.

- V-shaped if:
  - Central atom has 2 atoms bonded to it and 1 (or 2) lone pairs

- Tetrahedral if:
  - Central atom has 4 atoms bonded to it

- Trigonal planar if:
  - Central atom has 3 atoms bonded to it and no lone pairs

- Trigonal Pyramid if:
  - Central atom has 3 atoms bonded to it and 1 lone pair

Shape	# of Bonded Atoms on Central Atom	# of Lone Pairs on Central Atom
Linear	i. No central atom ii. 2 atoms	n/a 0
V – Shaped (Bent or Angular)	i. 2 atoms	1 or 2
Trigonal Planar	i. 3 atoms	0
Trigonal Pyramid	i. 3 atoms	1
Tetrahedral	i. 4 atoms	0

Steps to creating a VSEPR shape

- Draw Lewis and Structural diagrams

- Look at central atom and see how many atoms are bonded and how many lone pairs

- Examine previous chart to determine shape

- Draw VSEPR shape diagram

Examples: Draw the Lewis and Shape Diagrams for the following molecules

CCl4

OF_2

NF_3

SiI_2O

F_2

Practice Sheet 4

1. Use VSEPR theory to predict the geometry of a molecule of each of the following substances. Draw a stereochemical formula (shape diagram) for each molecule.

(a) $Cl_2O(g)$ (b) $H_2S(g)$ (c) $SiBr_4(l)$

(d) $PF_3(g)$ (e) HCl(g)

2. Use Lewis formulas and VSEPR theory to predict the shapes around each central atom of the following molecules:

(a) CO2(g), carbon dioxide (in "carbonated" beverages) (b) HCN(g), hydrogen cyanide (odour of bitter almonds, extremely toxic)

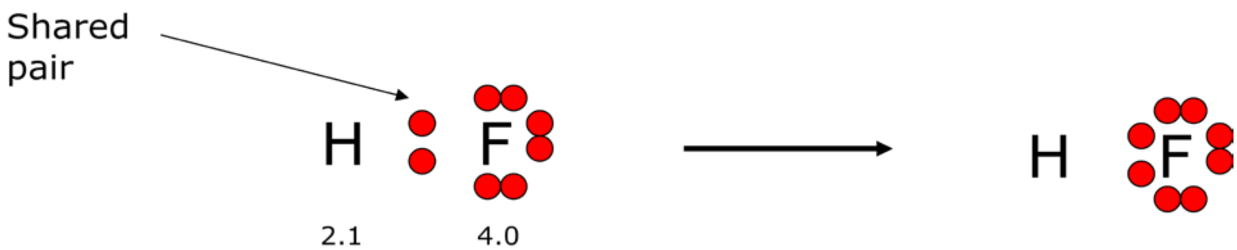
(c) CH2CHCH3(g), propene (monomer for polypropylene) (d) CHCCH3, propyne (in specialty fuels, for welding)

(e) H2CO(g), methanal (formaldehyde)

Topic 3: Polarity

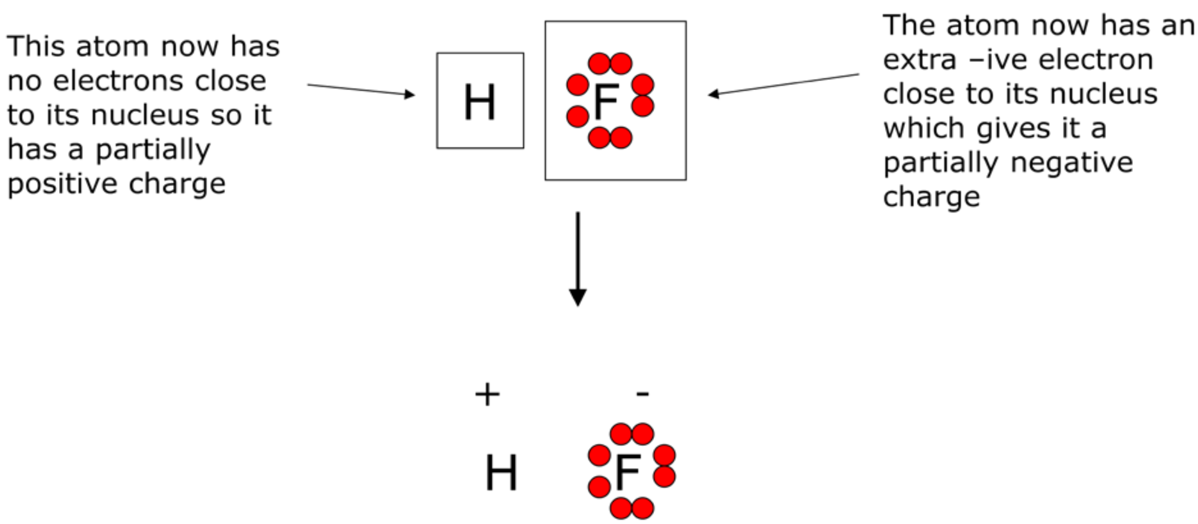
Electronegativity and Bond Polarity

- If two bonded atoms have different electronegativities, they will have unequal sharing of the shared pair of electrons
  - E.g $HF_{(g)}$



- Because the fluorine atom has a stronger attraction for electrons, it pulls the shared pair closer to its nucleus

- This unequal sharing gives the atoms partial charges

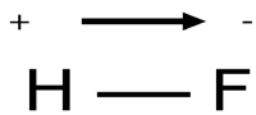


- Polar covalent bond
  - Covalent bond in which the two bonded atoms have a <u>different</u> electronegativity
  - Atoms ___________________________ share electrons

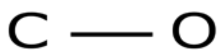
- Non-polar covalent bond
  - Covalent bond in which the two bonded atoms have the <u>same</u> electronegativity
  - Atoms _________________________ share electrons

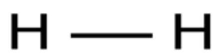
Bond Dipoles

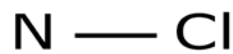
- We represent a polar bond with an arrow in the direction of the electron pull and the partial charge symbols

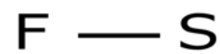


- Draw bond dipoles for the following bonds







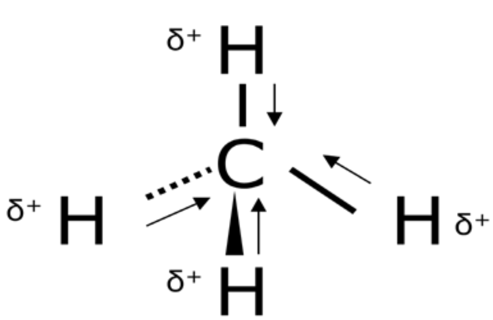


Polarity in Molecules

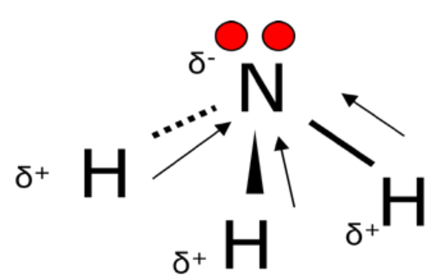
- Polar Molecule
  - A molecule in which the –ive charge is not distributed <u>symmetrically</u> among the molecule
- Non-Polar Molecule
  - A molecule in which the –ive charge is distributed symmetrically among the molecule
- To determine if a molecule is polar
  - Start by creating the VSEPR shape diagram
  - Draw in bond dipoles

Examples:

CH_4



NH_3



Practice Sheet 5

1. Predict the shape of the following molecules. Provide Lewis formulas and stereochemical formulas.
(a) silicon tetrabromide, $SiBr_4(l)$ (b) nitrogen trichloride, $NCl_3(l)$ (c) sulfur dichloride, $SCl_2(l)$

2. Predict the bond polarity for the following bonds. Use a diagram that includes the partial negative and positive charges and direction of the bond dipole:

(a) C-N in hydrogen cyanide

(c) P-S in P(SCN)3(s)

(b) N-O in nitrogen dioxide

(d) C-C in C8H18(l)

3. Predict the polarity of the following molecules. Include a stereochemical formula, bond dipoles, and the final resultant dipole (if nonzero) of the molecule.
(a) carbon disulfide, $CS_2(l)$ (b) oxygen difluoride, $OF_2(g)$ (c) phosphorus trichloride, $PCl_3(l)$

4. Use the empirical rules from Table 8, page 99, to predict the polarity of an octane, C8H18(l), molecule. Explain your answer without drawing the molecule

Topic 4: Intermolecular Forces and Boiling Points of Molecules

Intermolecular forces

- "Inter" means ___________________
- Molecular refers to <u>molecules</u>
- Forces between molecules
- These forces attract molecules to each other

Intramolecular forces

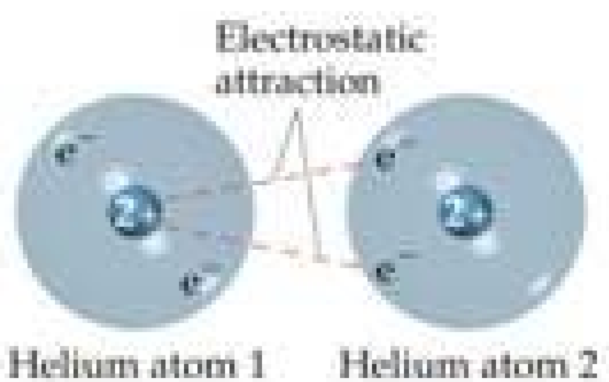
- "Intra" means ______________________
- Within a molecule
- Ex. Covalent bonds

Three types of intermolecular forces

-
-
-

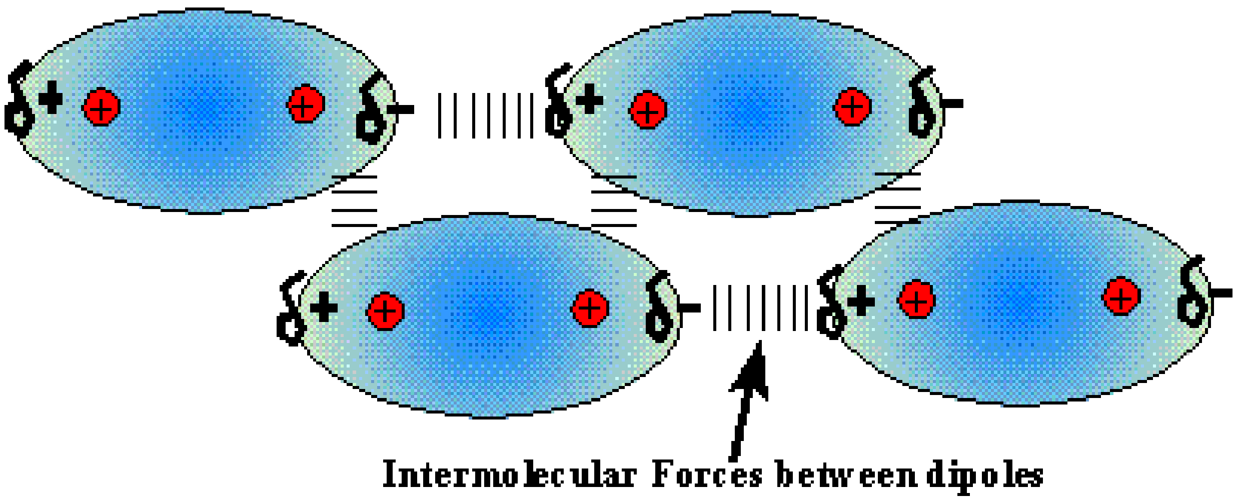
London Dispersion

- Positive and negative charges ***attract each other***.
- Every atom has ***positive protons*** in the ***nucleus*** and ***negative electrons*** on ***orbitals surrounding the nucleus***.
- ***positive protons of one molecule attract the negative electrons of the other***.



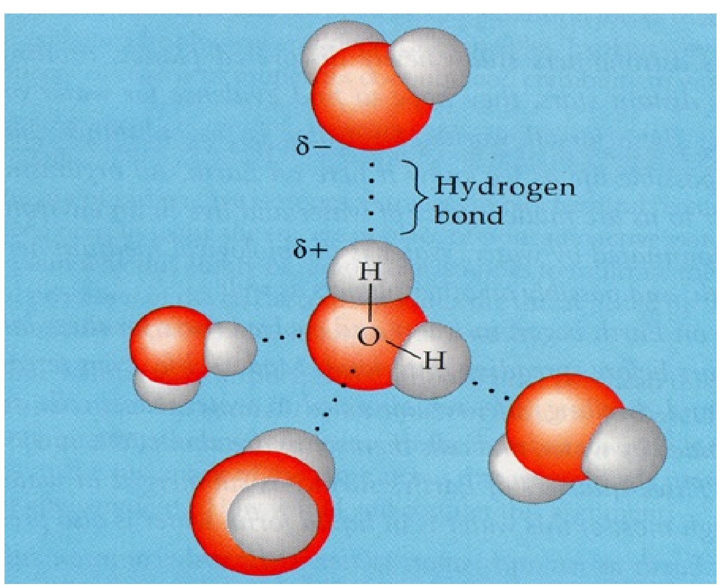
Dipole-Dipole

- Force of attraction between ____________________________.
- Slightly **negative poles** of one molecule attract slightly **positive poles** of another molecule
- Determines solubility of molecules
  - **Like dissolves like**
  - **Polar molecules dissolve in polar molecules**
  - **Non-polar molecules dissolve in non-polar molecules**



Hydrogen Bonding

- Force of attraction between a hydrogen bonded to __________________________ of one molecule and the **lone pairs** of O, F, or N on another molecule.
- It is the _______________________ intermolecular force



Effect of intermolecular forces

- Boiling point/Melting point
  - The higher the I.M. the higher the mp or bp
  - More forces to hold the molecules together
- Surface tension
  - Higher I.M. higher surface tension b/c forces hold surface of liquid intact

Example: Which of the following compounds would have a higher m.p/b.p.?

Cl_2 or Br_2 CH_4 or CCl_4 CCl_4 or $CHCl_3$

Practice Sheet 6

1. For each of the following pairs of chemicals, which one is predicted to have the stronger intermolecular attraction? Include a 1 sentence explanation with your answer.

a. chlorine or fluorine

b. fluorine or hydrogen chloride

c. methane or ammonia

d. water or hydrogen sulfide

e. silicon tetrahydride or methane

f. chloromethane (CH_3Cl) or ethanol

2. Hydrogen peroxide has hydrogen bonding forces holding its molecules together. Draw the shape diagrams for two molecules of hydrogen peroxide and include a dashed line showing where the hydrogen bonding force would be present.

3. A glass of water can be filled slightly above the rim without spilling over. Explain why this is using your knowledge of intermolecular forces.