

UNIT 3: IONIC AND COVALENT COMPOUNDS**Chapter 5: Ions and Ionic Compounds****5.1: Ions**

Chemical Reactivity: - the tendency for an element or a compound to react with other substances.

Octet Rule: - the tendency for electrons to fill the second and third energy levels (8 valence electrons – for main groups – IA to VIIIA columns) to achieve stability.

- in most cases, this means having the same electron arrangement of the **nearest** noble gas.
- exceptions to the rule include helium (only 2 electrons to fill the first energy level), and the transition metals.

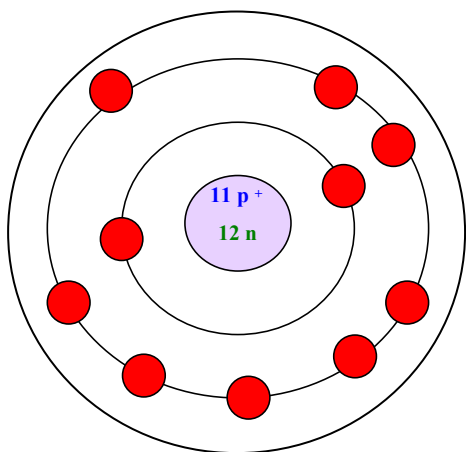
Groups or Families	Chemical Properties	Reasons
Alkali Metals (IA)	very reactive metals	Easily shed 1 valence e^- to achieve an octet
Alkaline Earth Metals (IIA)	less reactive than alkali metals	Can shed 2 valence e^- to achieve an octet
Halogens (VIIA)	very reactive non-metals	Easily gain 1 valence e^- to achieve an octet
Noble Gases (VIIIA)	very stable; all are gaseous state at room temperature	Already has an octet configuration as its valence energy level (always ends in np^6 – except for helium, which ends in $1s^2$)

Ions: - when atoms lose or gain electrons, they attain a positive or negative charge.

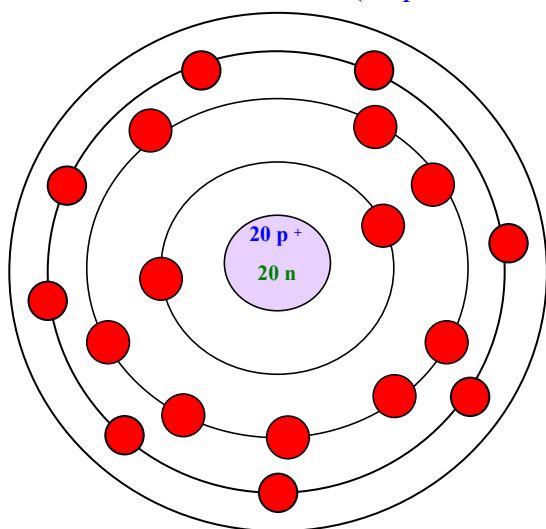
- Cations:** - positive charged ions (atoms that lose electrons).
- to name cations → element name follows by the word “ion”

Example 1: Draw the energy level diagram for the following cations.

- a. Sodium ion = Na^+ (11 p^+ and 10 e^-)

**Sodium Ion (Na^+)****Atomic Number: 11****Atomic Mass: 22.99****Nucleus: 11 p^+ and 12 n****2nd Energy Level: 8 e^-** **(8 valence e^- - Filled)****1st Energy Level: 2 e^-** **Total: 10 e^-** **Net Charge = 1+****Location on the Period Table of Elements:****Third Row; Column IA**

b. Calcium ion = Ca^{2+} (20 p^+ and 18 e^-)



Calcium Ion (Ca^{2+})

Atomic Number: 20

Atomic Mass: 40.08

Nucleus: 20 p^+ and 20 n

3rd Energy Level: 8 e^-

(8 valence e^- - Filled)

2nd Energy Level: 8 e^-

1st Energy Level: 2 e^-

Total: 18 e^-

Net Charge = 2+

Location on the Period Table of Elements:

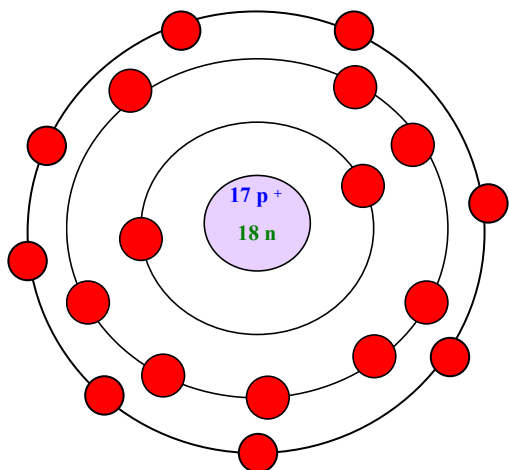
Fourth Row; Column IIA

2. **Anions:** - **negative** charged ions (atoms that gain electrons).

- to name anions → keep the first part of element name follow by suffix *-ide*

Example 2: Draw the energy level diagram for the following anions.

a. Chloride = Cl^- (17 p^+ and 18 e^-)



Chloride (Cl^-)

Atomic Number: 17

Atomic Mass: 35.45

Nucleus: 17 p^+ and 18 n

3rd Energy Level: 8 e^-

(8 valence e^- - Filled)

2nd Energy Level: 8 e^-

1st Energy Level: 2 e^-

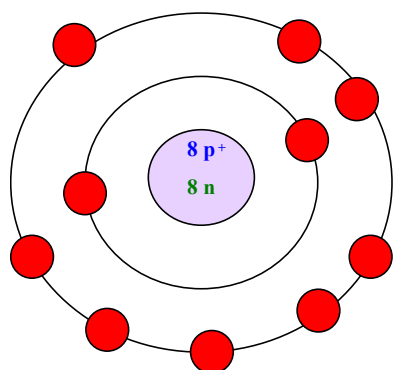
Total: 18 e^-

Net Charge = 1-

Location on the Period Table of Elements:

Third Row; Column VIIA

b. Oxide = O^{2-} (8 p^+ and 10 e^-)



Oxide (O^{2-})

Atomic Number: 8

Atomic Mass: 16.00

Nucleus: 8 p^+ and 8 n

2nd Energy Level: 8 e^-

(8 valence e^- - Filled)

1st Energy Level: 2 e^-

Total: 10 e^-

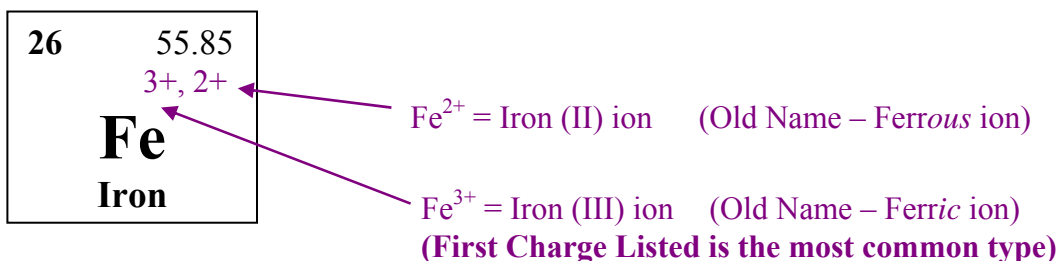
Net Charge = 2-

Location on the Period Table of Elements:

Second Row: Column VIA

Transition Metals (1B to 10B): - groups and periods of metals that can have **varying charges**.
 - use **Roman Numerals** as part of their ionic names.

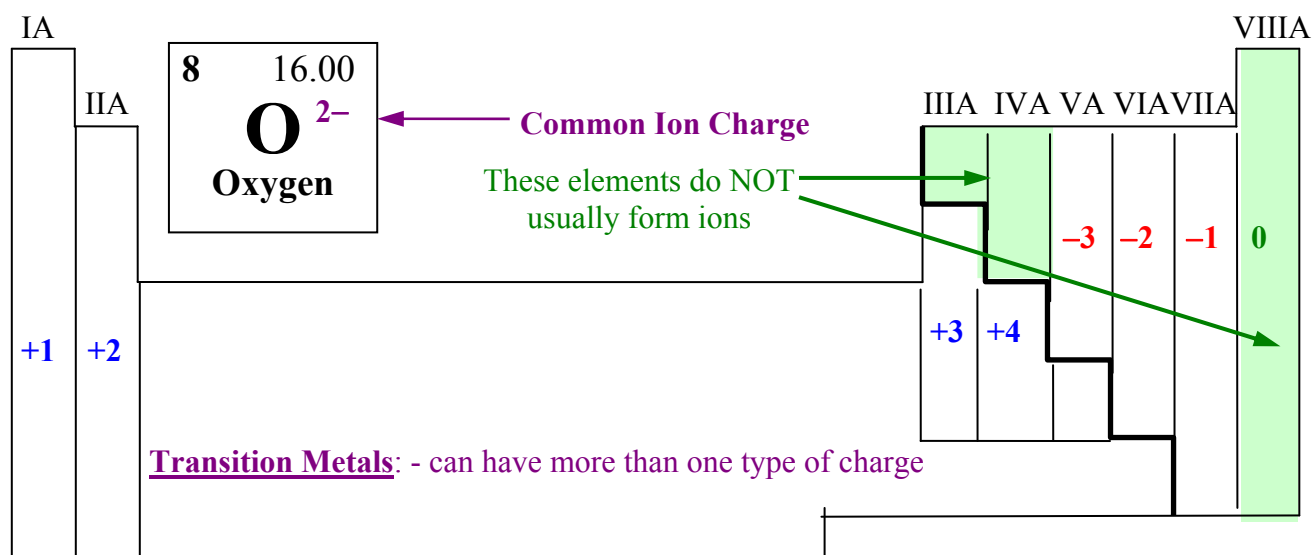
Example: Fe^{3+} and Fe^{2+} ions



Chemical Properties of Metals and Non-Metals:

- Metals lose electrons to become positive ions – cations.**
- Non-Metals gain electrons to become negative ions – anions.**
- Hydrogen usually loses an electron to become a H^+ ion. However, it can sometimes gain an electron to become H^- (Hydride).
- The last column of the Table of Elements does not usually form ions.** These elements are called the **Noble Gases** (Helium, Neon, Argon, Krypton, Xenon, and Radon).

The number of electrons an atom loses or gains depends on which column (vertical) the element is at the Table below.



The reason that noble gases (column VIII A) do not form ions is because their outermost shells are filled with the maximum number of electrons allowed. That is why we call this group of elements “noble gases”. They do not form ions because they are stable. Hence we use the word “noble” to describe them. All the other elements form ions because they want to achieve stability like the noble gases. If you observe carefully, oxide has the same number of electrons as the nearest noble gas, neon. On the other hand, calcium ion has the same number of electrons as the nearest noble gas, argon. In terms of stability, which is another word for lower energy state, these ions are more stable than their respective atoms.

Since the number of valence electrons of an atom is the same as its column number, all the elements of column IA have 1 valence electron. As we see with lithium, all they have to do is to lose that valence electron to achieve a noble gas “like” state. For elements in column IIA, they all have 2 valence electrons. Hence, they lose 2 electrons to acquire stability and become ions with a net charge of +2. The following table summarises these points.

Column	Number of Valence Electrons	Methods to achieve a Stable State	Net Charge of Ions
IA	1	lose 1 electron or gain 7 electrons	+1
IIA	2	lose 2 electrons or gain 6 electrons	+2
IIIA	3	lose 3 electrons or gain 5 electrons	+3
IVA	4	lose 4 electrons or gain 4 electrons	+4
VA	5	lose 5 electrons or gain 3 electrons	-3
VIA	6	lose 6 electrons or gain 2 electrons	-2
VIIA	7	lose 7 electrons or gain 1 electron	-1
VIIIA	8	already has the maximum number of electrons allowed in the outermost electron shell.	0

Ionic Equation: - a chemical equation that shows non-metal elements gaining electrons to become anions, or metal elements losing electrons to become anions,

Electron Configurations of Ions: - for **metals, which like to lose electrons to form cations**, they have the same electron configurations as the **noble gas of the previous row**.
 - for **non-metals, which like to gain electrons to form anions**, they have the same electron configurations as the **noble gas at the end of the same row**.

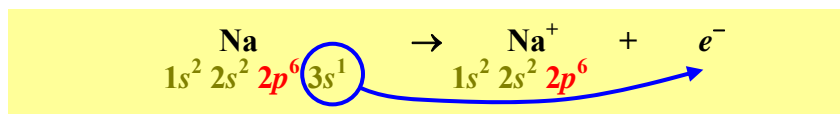
Alkali Cations: - cations that were the result as alkali metals (Group 1 or IA) losing one valence electron.
 - they are Li^+ , Na^+ , K^+ , Rb^+ , Cs^+ and Fr^+

Alkaline Cations: - cations that were the result as alkaline earths (Group 2 or IIA) losing two valence electrons.
 - they are Be^{2+} , Mg^{2+} , Ca^{2+} , Sr^{2+} , Ba^{2+} and Ra^{2+}

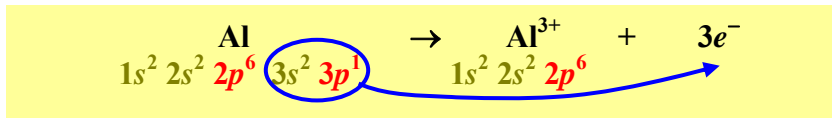
Halide Ions: - anions that were the result as halogen (Group 17 or VIIA) gaining one valence electron to satisfy the octet rule.
 - they are F^- , Cl^- , Br^- , I^- and At^-

Example 3: Write the electron configuration and an ionic equation showing the gaining or losing of electrons when each of the following atoms becomes their most popular ions.

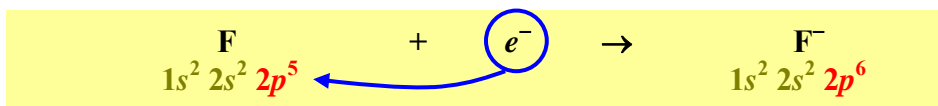
a. sodium atom to sodium ion



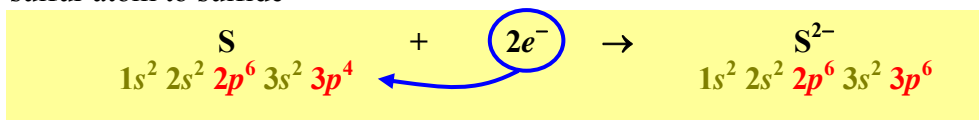
b. aluminium atom to aluminium ion



c. fluorine atom to fluoride



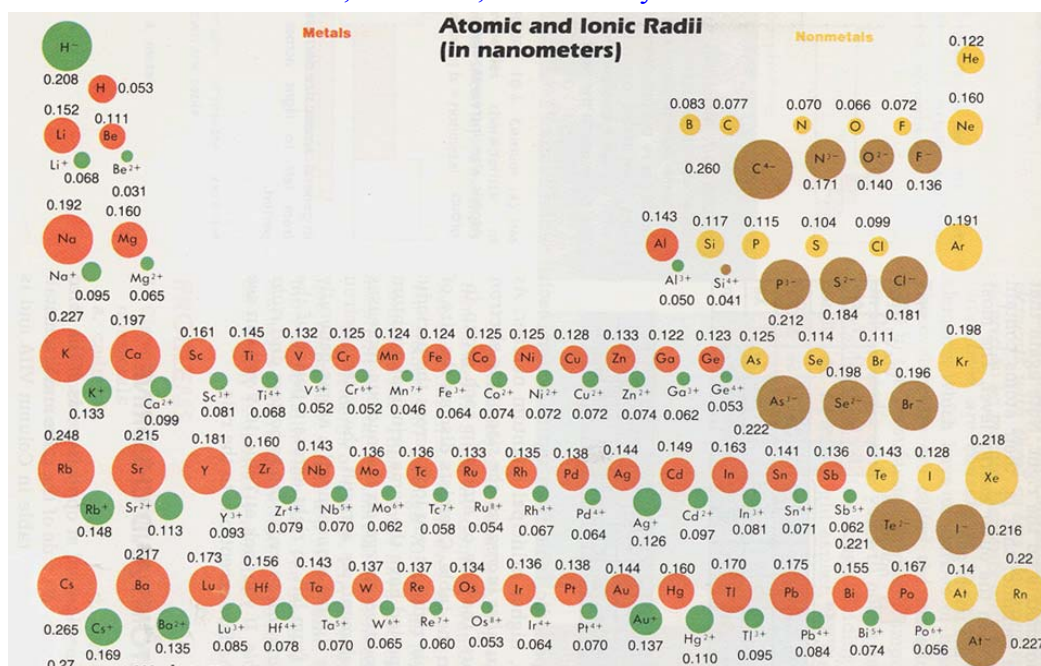
d. sulfur atom to sulfide



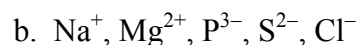
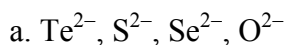
Ions Sizes: - the size of cations and anions.

Several Notes on Trends in Ion Size

- 1. Metal Cations are generally Smaller than Non-Metal Anions WITHIN the Same Period.** Metal cations tend to lose electrons to achieve the electron configuration of a previous noble gas. Thereby, losing an energy level in the process. **Metal Cations are always Smaller than the Parent Neutral Atoms. Non-Metal Anions are always Larger than the Parent Neutral Atoms.**
- 2. In general, Ion Sizes Decrease as one move from LEFT to RIGHT of a period WITHIN the METAL GROUPS and WITHIN the NON-METAL GROUPS.** This is because the increases in protons in the same row increase the effective nuclear charge (the protons in the nucleus have more pull on the outer electrons, decreasing shielding), thus drawing these outer electrons closer to the nucleus, decreasing in sizes as the result.
- 3. Ionic Radii INCREASES Down a Group.** This is due to the fact there are more orbitals as the number of row increases. The outer electrons are, of course, further away from the nucleus



Example 4: Order the following ions from the smallest to the largest.



These anions are within the same Group (column). As we move down the column, ion size increases. Therefore,



These ions are within the same Period (row). As we move to the right, ion size decreases within each of the metal and non-metal groups. Therefore,



Assignment

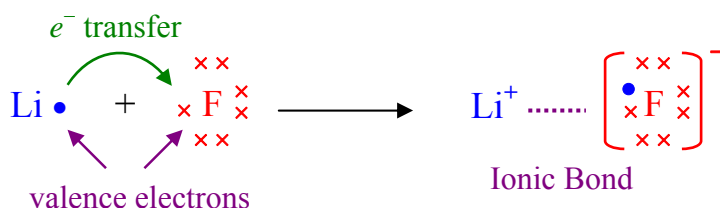
5.1 pg. 165 # 1 to 13; pg. 141 #8 and 13

5.2: Ionic Bonding and Salts

Ionic Compound:

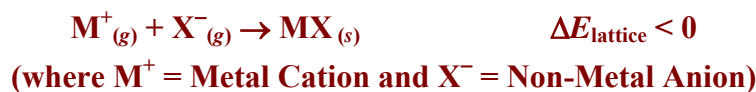
- when a **metal** element combines with a **non-metal** element.
- forms **ionic bonds** (electrons are “stolen” or “transferred” from one atom to another).
- dissociates into **electrolytes** (forms ions, cations and anions, when dissolve in water).
- the resulting ionic compounds are commonly referred to as **“salt”**.

Example: LiF



Lattice Energy:

- the energy involved when one mole of an ionic solid is formed from separated gaseous ions.
- lattice energy is usually exothermic (energy is released).

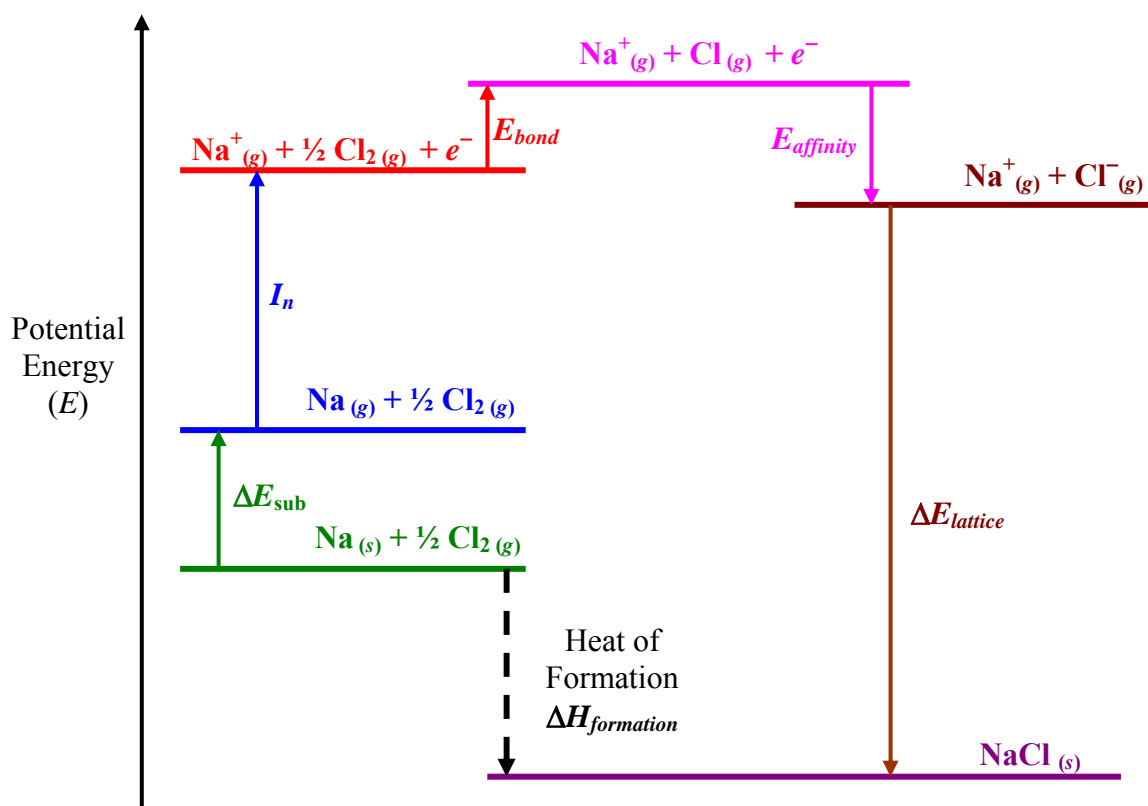


Determining the Heat to Form Ionic Compounds

1. **Enthalpy (Heat) Sublimation of Solid Metal into Gaseous Metal (ΔE_{sub}).** $\text{M}_{(s)} \rightarrow \text{M}_{(g)}$
2. **Ionization of Gaseous Metal into Gaseous Cation (I_n).** $\text{M}_{(g)} \rightarrow \text{M}^+_{(g)} + e^-$
3. **Dissociation of Non-Metal Molecules into Atoms (E_{bond} = Bond Energy).** $\frac{1}{2} \text{X}_2_{(g)} \rightarrow \text{X}_{(g)}$
4. **Formation of Anion from Non-Metal Atom (E_{affinity} = Electron Affinity).** $\text{X}_{(g)} + e^- \rightarrow \text{X}^-_{(g)}$
5. **Formation of Solid Ionic Compound from Gaseous Ions ($\Delta E_{\text{lattice}}$).** $\text{M}^+_{(g)} + \text{X}^-_{(g)} \rightarrow \text{MX}_{(s)}$
6. **Summation of all Energies involved for $\Delta H_{\text{formation}}$** $\text{M}_{(s)} + \frac{1}{2} \text{X}_{(g)} \rightarrow \text{MX}_{(s)}$

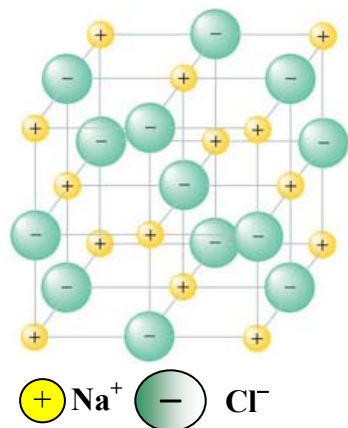
Potential Energy Diagram: - diagram that shows the energy changes of various stages of a reaction.

Example: Potential Energy Diagram for the formation of sodium chloride (NaCl).

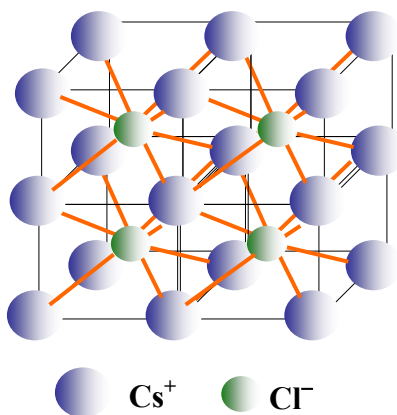


Properties of Ionic Crystalline Solids

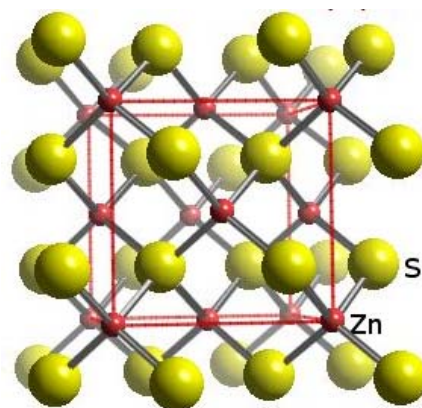
- 1. Ionic Compounds have a definite Crystalline Structure and are Poor Conductors of Electricity and Heat in their Solid Form.** Conduction of electricity and heat requires ions to move freely within the solid. The lattice structures of the solid ionic compounds do not allow ions to move freely.



Simple Cubic Unit Crystalline Structure of NaCl



Body Centred Cubic Unit Crystalline Structure of CsCl



Face Centred Cubic Unit Crystalline Structure of ZnS

- The smallest unit of an Ionic Compound is called a “Unit Cell”**. Because of the regular crystal structure of an ionic compound, they are not referred to as molecules.
- Ionic solids are generally High Melting Points (typically 300°C to 1000°C)**. Since a strong force can only shatter the crystal but not bend it as in metals, **the energy needed to completely break up the lattice structure (lattice energy)** is very large and it is the same energy needed to melt the ionic compounds.
- Ionic solids are Hard and Brittle**. The lattice structure of all ionic compounds holds the ions in definite positions. When the compound encountered a strong force, the close proximity of the ions stay close together. This causes the crystal to **shatter**, not bent like metal solid would.
- Ionic solids can be Melted to form Liquids that are Electrical Conductors**. Ionic solids melt when the ions gain enough energy to break the lattice structure. They are move freely and can carry electrical charge through the liquid. This explains why a molten ionic substance conducts electricity, but a solid ionic material doesn't. The ions move through the liquid can carry charge from one place to another.
- Soluble ionic solids dissolve to form solutions that are Electrical Conductors. (Not all ionic substances are soluble in water.)** Soluble ionic compounds form **electrolytes (ions in aqueous from)** that allow the conduction of electricity.

Assignment**5.2 pg. 175 #1 to 6, 8 to 11****5.3: Names and Formulas of Ionic Compounds****Monoatomic Ions**: - ions that came from a single atom (include metal cations and non-metal anions).

- monoatomic anion ends with suffix *-ide*.
- some transition metal ions require Roman Numeral when there can be more than one type of charge.

Examples: Na^+ = sodium ion, Cl^- = chloride, Pb^{4+} = lead (IV) ion, Zn^{2+} = zinc ion**Polyatomic Ions**: - ions that contain many atoms.

- mostly anions (except NH_4^+ = ammonium ion).
- most ends with suffixes *-ate* or *-ite* (some ends with suffix *-ide*).
- some ends with suffix *-ide* still (cyanide - CN^- and hydroxide - OH^-).

Examples: CO_3^{2-} = carbonate, $\text{Cr}_2\text{O}_7^{2-}$ = dichromate, OH^- = hydroxide, SO_3^{2-} = sulfite**Nomenclature**: - a naming system.**IUPAC**: - International Union of Pure and Applied Chemistry.

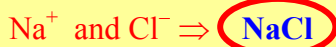
- an organization that oversees the standard regarding chemistry including chemical nomenclature.

Nomenclature of Ionic Compounds

1. **Balance the Cation and Anion Charges.**
2. Use **brackets** for **multiple Polyatomic Ions**.
3. When naming, use ***-ide*** for the **non-metal anions**.
4. **Metals** that can have **two or more different charges** must use **Roman Numerals** in the names.

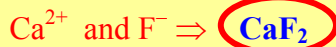
Example 1: Write the chemical formula of the followings.

a. sodium chloride



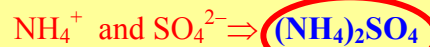
Need 1 Na^+ & 1 Cl^- to balance charges

b. calcium fluoride



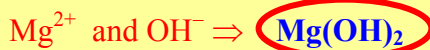
Need 1 Ca^{2+} & 2 F^- to balance charges

c. ammonium sulfate



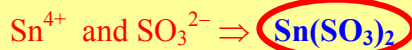
2 (NH_4^+) & 1 SO_4^{2-} to balance charges

d. magnesium hydroxide



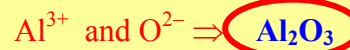
Need 1 Mg^{2+} & 2 (OH^-) to balance charges

e. tin (IV) sulfite



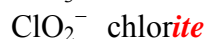
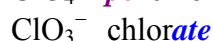
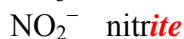
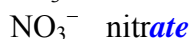
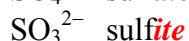
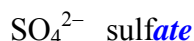
Need 1 Sn^{4+} (IV means 4+ charge) & 2 (SO_3^{2-}) to balance charges

f. aluminium oxide



Need 2 Al^{3+} & 3 O^{2-} to balance charges

Oxoanions: - a series of polyatomic ions that contains different number of oxygen atoms.



Example 2: Name the following ionic compounds.

a. Na_2S

Na^+ = sodium
 S^{2-} = sulfide

sodium sulfide

b. KMnO_4

K^+ = potassium
 MnO_4^- = permanganate

potassium permanganate

c. $\text{Li}_2\text{S}_2\text{O}_3$

Li^+ = lithium
 $\text{S}_2\text{O}_3^{2-}$ = thiosulfate

lithium thiosulfate

d. CuCr_2O_7

Cu^{2+} = copper (II)
 $\text{Cr}_2\text{O}_7^{2-}$ = dichromate

copper (II) dichromate

e. $\text{Fe}_3(\text{BO}_3)_2$

Fe^{2+} = iron (II)
 BO_3^{3-} = borate

iron (II) borate

f. $\text{Ti}(\text{ClO}_3)_3$

Ti^{3+} = titanium (III)
 ClO_3^- = chlorate

titanium (III) chlorate

Hydrate: - ionic compounds sometimes come with water molecule locked in their crystal form.

- to name hydrates → use the ionic compound name, then write the prefix follow by the word “hydrate”.

Prefixes for Hydrates

1 – mono 2 – di 3 – tri 4 – tetra 5 – penta
6 – hexa 7 – hepta 8 – octa 9 – nona 10 – deca

Example: $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ copper (II) sulfate pentahydrate

Assignment

**5.3 pg. 180 (Practice) #1; pg. 180 #1 to 8 and
Worksheet: Nomenclatures of Ionic Compounds
Chapter 5 Review pg. 183–184 #1 to 29**

Worksheet: Nomenclatures of Ionic Compounds

Write chemical formula for the compounds below. Show subscript numbers where needed.

- | | | |
|-------------------------------|-----------------------------------|--------------------------------------|
| 1. sodium chloride | 2. strontium nitrate | 3. calcium fluoride |
| 4. lithium acetate | 5. hydrogen phosphate | 6. potassium sulfate |
| 7. magnesium chromate | 8. calcium hydroxide | 9. aluminum sulfite |
| 10. ammonium sulfate | 11. ammonium iodide | 12. iron (III) nitrite |
| 13. lead (II) sulfite | 14. copper (I) sulfide | 15. aluminum oxide |
| 16. sodium chlorate | 17. sodium hydrogen carbonate | 18. nickel (III) hypobromite |
| 19. silver dichromate | 20. barium cyanide | 21. titanium (III) periodate |
| 22. aluminum oxalate | 23. cobalt (II) borate | 24. lithium hydrogen phosphate |
| 25. barium chloride dihydrate | 26. iron (II) nitrate hexahydrate | 27. nickel (II) sulfate heptahydrate |
| 28. vanadium (V) sulfide | 29. cadmium phosphide | 30. gold (III) thiosulfate |

Name the ionic compounds for the chemical formula below. Use Roman numerals where needed.

- | | | | | |
|--|--|--|----------------------------------|------------------------------|
| 31. Na_2S | 32. Na_2SO_3 | 33. Na_2SO_4 | 34. $(\text{NH}_4)_2\text{CO}_3$ | 35. $\text{Fe}(\text{OH})_3$ |
| 36. $\text{Zn}(\text{CH}_3\text{COO})_2$ | 37. SnCl_4 | 38. $\text{Ti}_2(\text{CrO}_4)_3$ | 39. NH_4MnO_4 | 40. AgNO_2 |
| 41. Cu_2O | 42. $\text{Mn}(\text{OOC}\text{COO})_2$ | 43. $\text{Ni}(\text{BrO}_2)_2$ | 44. $\text{Sc}(\text{IO})_3$ | 45. HgCO_3 |
| 46. $\text{Cu}(\text{NO}_3)_2 \cdot 3\text{H}_2\text{O}$ | 47. $\text{Pb}(\text{SO}_4)_2 \cdot 2\text{H}_2\text{O}$ | 48. $\text{CoF}_2 \cdot 4\text{H}_2\text{O}$ | 49. Zn_3P_2 | 50. GeO_2 |

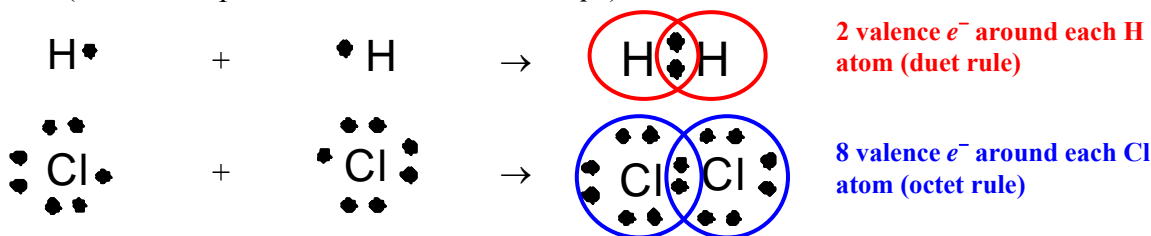
Answers

- | | | | | |
|--|---|---------------------------------------|--|---|
| 1. NaCl | 2. $\text{Sr}(\text{NO}_3)_2$ | 3. CaF_2 | 4. LiCH_3COO | 5. H_3PO_4 |
| 6. K_2SO_4 | 7. MgCrO_4 | 8. $\text{Ca}(\text{OH})_2$ | 9. $\text{Al}_2(\text{SO}_3)_3$ | 10. $(\text{NH}_4)_2\text{SO}_4$ |
| 11. NH_4I | 12. $\text{Fe}(\text{NO}_2)_3$ | 13. PbSO_3 | 14. Cu_2S | 15. Al_2O_3 |
| 16. NaClO_3 | 17. NaHCO_3 | 18. $\text{Ni}(\text{BrO})_3$ | 19. $\text{Ag}_2\text{Cr}_2\text{O}_7$ | 20. $\text{Ba}(\text{CN})_2$ |
| 21. $\text{Ti}(\text{IO}_4)_3$ | 22. $\text{Al}_2(\text{OOC}\text{COO})_3$ | 23. $\text{Co}_3(\text{BO}_3)_2$ | 24. Li_2HPO_4 | 25. $\text{BaCl}_2 \cdot 2\text{H}_2\text{O}$ |
| 26. $\text{Fe}(\text{NO}_3)_2 \cdot 6\text{H}_2\text{O}$ | 27. $\text{NiSO}_4 \cdot 7\text{H}_2\text{O}$ | 28. V_2S_5 | 29. Cd_3P_2 | 30. $\text{Au}_2(\text{S}_2\text{O}_3)_3$ |
| 31. sodium sulfide | 32. sodium sulfite | 33. sodium sulfate | | |
| 34. ammonium carbonate | 35. iron (III) hydroxide | 36. zinc acetate | | |
| 37. tin (IV) chloride | 38. titanium (III) chromate | 39. ammonium permanganate | | |
| 40. silver nitrite | 41. copper (I) oxide | 42. manganese (IV) oxalate | | |
| 43. nickel (II) bromite | 44. scandium hypoiodite | 45. mercury (II) carbonate | | |
| 46. copper (II) nitrate trihydrate | 47. lead (IV) sulfate dihydrate | 48. cobalt (II) fluoride tetrahydrate | | |
| 49. zinc phosphide | 50. germanium oxide | | | |

Chapter 6: Covalent Compounds

6.1: Covalent Bonds

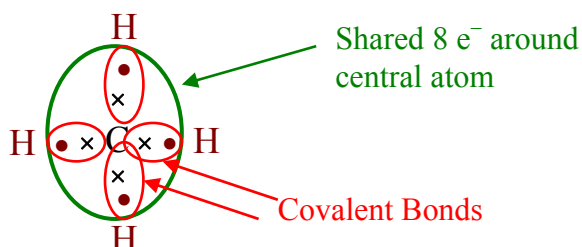
Covalent Bond: - an attraction force between two non-metal elements where the sharing of valence electrons is taken place so each atom can achieve the stable duet (for hydrogen) or octet (for other representative elements – ns^2np^6) conditions.



Covalent Compound: - forms **covalent bonds** (electrons are “shared” between atoms). The space these electrons share is referred to as **molecular orbitals**.

- sometimes refers to as **molecular compound**.
- usually forms when a **non-metal** element combines with a **non-metal** element.
- forms **non-electrolytes** (do not dissociate into ions when dissolve in water).

Example: CH_4

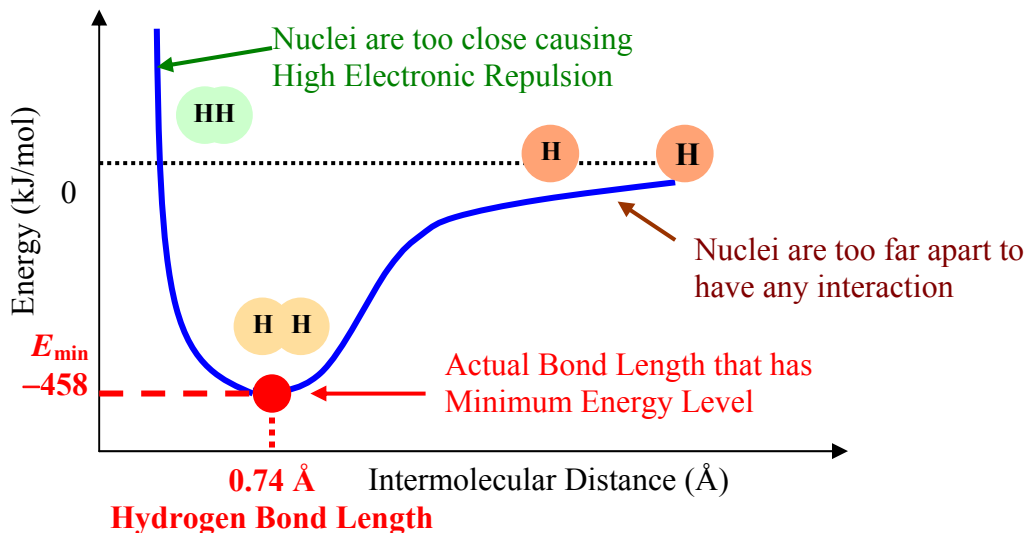


Chemical Bond Model: - chemical bonds are the result when atoms are arranged to achieve a lower energy state.

Molar Bond Energy: - the energy to required breaking or released from forming one mole of a particular chemical bond.

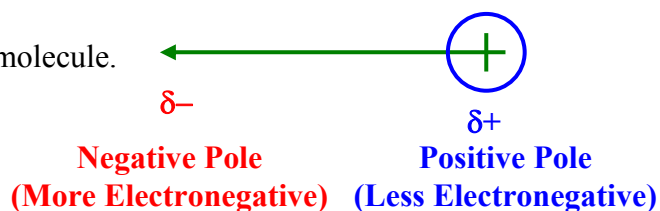
- in general, **Bond Energy > 0 (endothermic – requires energy input) when breaking a bond;**
Bond Energy < 0 (exothermic – releases energy to the surroundings) when forming a bond.
- the **more energy the reaction is released when a bond is formed, the more stable is the molecule.**

Bond Length: - the length of a chemical bond as measured by the **lowest energy level**.

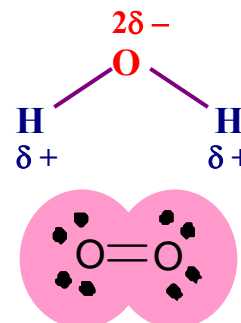


Dipole: - the direction of the charge distribution of a polar molecule.

- the length of the vector indicates the strength of the bond polarity, whereas the arrow head indicates the direction of the higher electronegative atom (negatively shifted).



Polar Covalent Bond: - a bond where electrons are not evenly distributed, resulting in localized charges at either end.



Non-Polar Covalent Bond: - a bond where electrons are evenly distributed, resulting in a non-polar molecule.

Bond Polarity: - the difference in electronegativity in a molecule.

- the larger the bond polarity, the higher the bond energy, and the stronger is the bond. This is why ionic bonds are very strong and ionic compounds are hard.

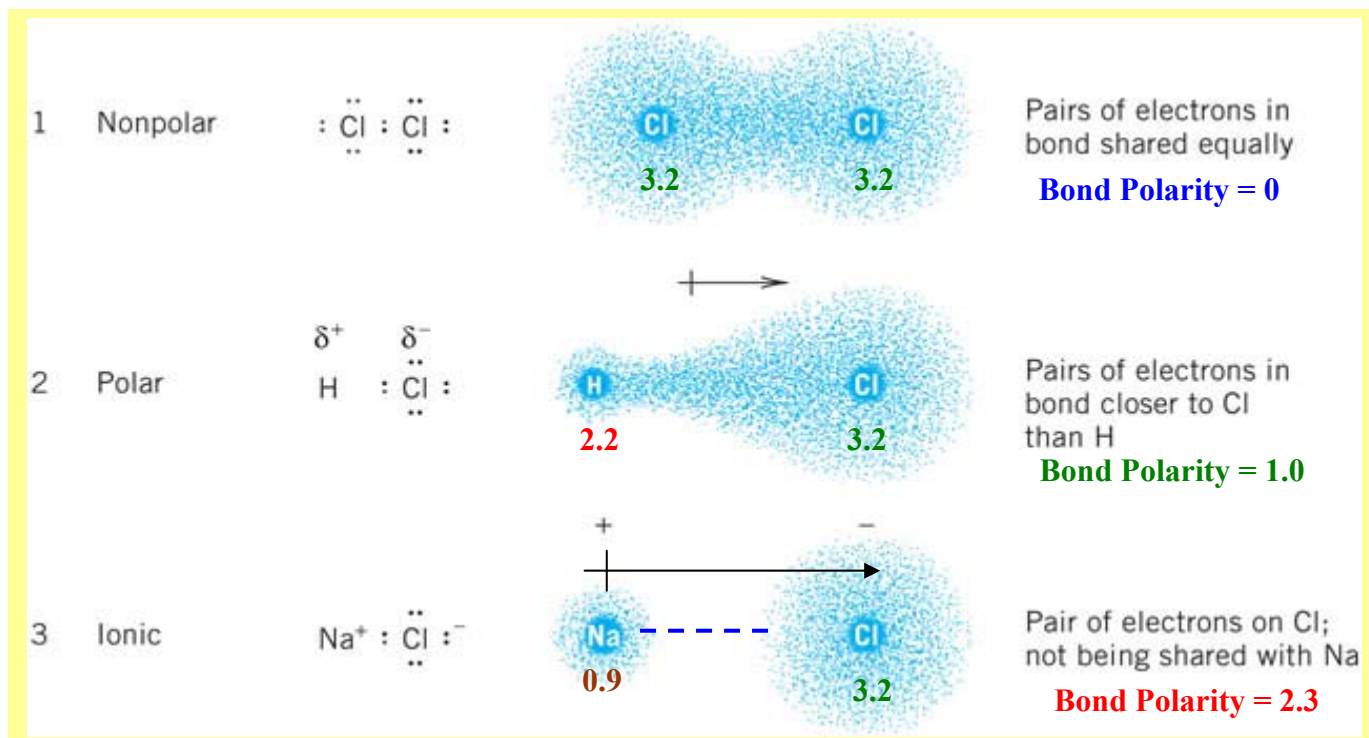
1. **Non-Polar Covalent Bonds have No Bond Polarity.**

2. **Polar Covalent Bonds have Some Bond Polarity.**

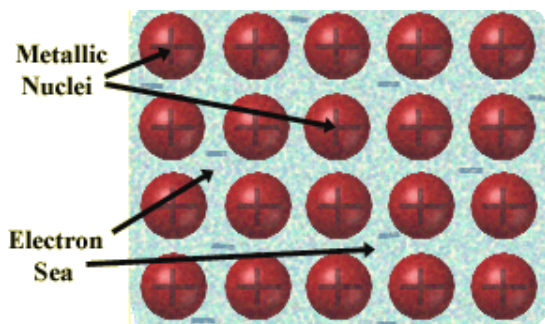
3. **Ionic Bonds have High Bond Polarity.**

Bond Polarity \uparrow Bond Energy \uparrow
Bond Strength \uparrow

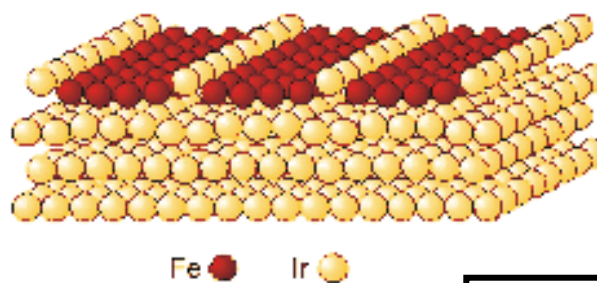
Example 1: Determine the bond polarity of Cl_2 , HCl , and NaCl . Order them from least to greatest and classify them.



Metallic Bonds: – the result of the attraction between the electrons in the outermost energy level of each metal atom and all the other atoms in the solid metal.



Metallic Bonds in metals and Alloys with metallic nuclei in a sea of valence electrons.



Iron-Iridium Alloy where atoms are organized in layers

Assignment
6.1 pg. 198 #1 to 14

6.2: Drawing and Naming Molecules

Nomenclature of Molecular Compounds

- Do NOT use charges to balance subscripts.** Use prefixes to name or write the formula's subscripts.
- If the first element has one atom in the molecule, do NOT use *mono~* as a prefix.
- The last element uses the suffix *-ide*.

Prefixes for Binary Molecular Compounds

1 – mono	3 – tri	5 – penta	7 – hepta	9 – nona
2 – di	4 – tetra	6 – hexa	8 – octa	10 – deca

Example 1: Name the following molecular compounds.

a. CO

1 Carbon and 1 Oxygen

Carbon monoxide

b. CO₂

1 Carbon and 2 Oxygen

Carbon dioxide

c. N₂O₄

2 Nitrogen and 4 Oxygen

dinitrogen tetraoxide

Example 2: Provide the chemical formula for the following compounds.

a. sulfur trioxide

1 S and 3 O ⇒ **SO₃**

b. diphosphorus pentoxide

2 P and 5 O ⇒ **P₂O₅**

c. silicon dioxide

1 Si and 2 O ⇒ **SiO₂**

Common Names for Some Molecular Compounds (Memorize!)

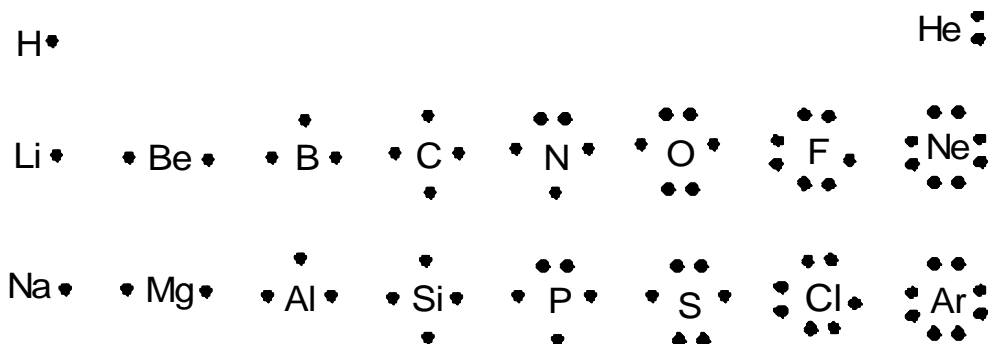
H ₂ O	Water	H ₂ O ₂	Hydrogen Peroxide	O ₃	Ozone	CH ₄	Methane
C ₃ H ₈	Propane	NH ₃	Ammonia	CH ₃ OH	Methanol	C ₂ H ₅ OH	Ethanol
C ₆ H ₁₂ O ₆	Glucose	C ₁₂ H ₂₂ O ₁₁	Sucrose				

Note: Do NOT use prefixes for the above common molecular compounds!

Lewis Structure: - sometimes refer to as Lewis Dot Diagram.

- shows electrons of valence electron (s and p orbitals) as dots for all atoms of representative elements (Groups IA to VIIIA).
- max of eight electrons around each atomic symbol, and hence Lewis structure cannot be used on transition metals.

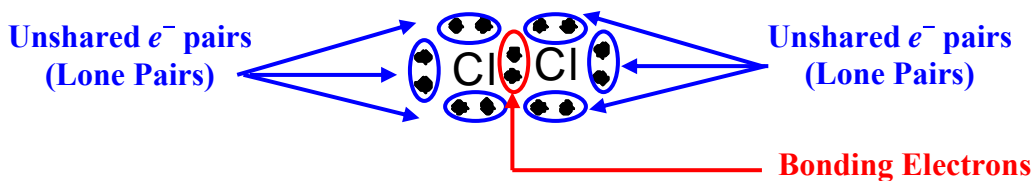
Lewis Structure for Atoms in the First Three Rows of the Periodic Table of Elements



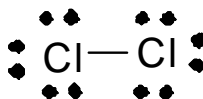
Types of Covalent Chemical Bonds:

- Single Covalent Bond:** - covalent bond where there are only **two valence electrons** being shared between two atoms (also refer to as **bonding electrons**).

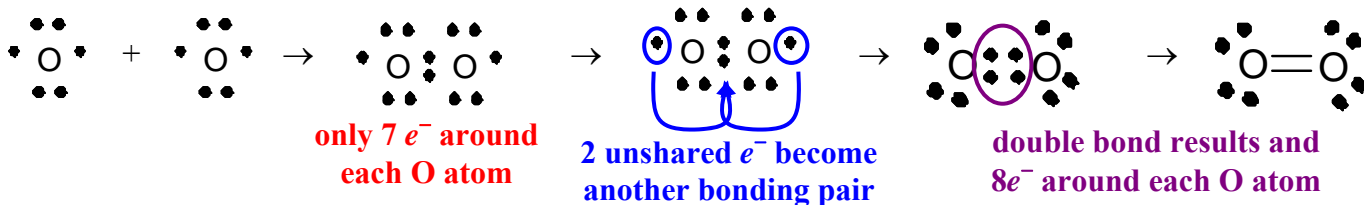
Unshared Electron Pairs (Lone Pairs): - pairs of electrons **not** involving in the covalent bond.



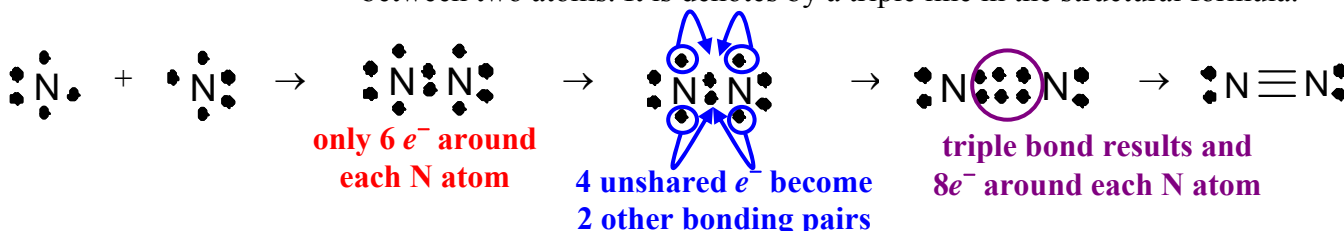
Structural Formula: - a Lewis dot diagram of a molecule where the bonding electrons are replaced by a short line.



- Double Covalent Bond:** - covalent bond where there are only **four valence electrons** being shared between two atoms. It is denoted by a double line in the structural formula.



- Triple Covalent Bond:** - covalent bond where there are only **six valence electrons** being shared between two atoms. It is denoted by a triple line in the structural formula.



Note: In general, bond length decreases with multiple bonds.

Bond and Bond Type	Bond Length (pm)
C – C (single bond)	154
C = C (double bond)	134
C ≡ C (triple Bond)	120

Note: In general, bond strength increases with multiple bonds.

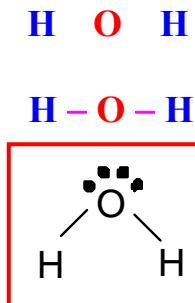
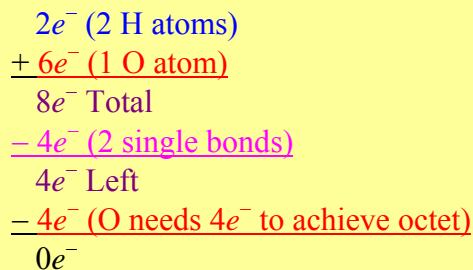
Selected Bond Energies (kJ/mol)							
Single Bonds			Multiple Bonds				
H – H	436	N – H	391	O – H	464	C = C	614
H – F	569	N – N	159	O – O	138	C ≡ C	837
H – Cl	431	N – F	272	O – F	190		
H – Br	363	N – Cl	200	O – Cl	203	O = O	498
H – I	295	N – Br	243	O – I	234		
		N – O	201			C = O	803
C – H	414			S – H	339	C ≡ O	1072
C – C	347	F – F	154	S – F	327		
C – N	389	F – Cl	253	S – Cl	253	N = N	418
C – O	358	F – Br	237	S – Br	218	N ≡ N	946
C – F	485	Cl – Cl	243	S – S	266		
C – Cl	339	Cl – Br	218			N = O	631
C – Br	276	Br – Br	193	Si – Si	340		
C – I	240	I – I	149	Si – H	393	C = N	615
C – S	259	I – Cl	208	Si – C	360	C ≡ N	891
		I – Br	175	Si – O	452		

Procedure to draw Lewis Dot Diagram for Covalent Molecules

- Add up all the valence electrons available from all atoms and extra charges from complex ions (ex: charge of 1⁻ means one more electron; charge of 1⁺ means one less electron).**
- Decide on the central atom. Usually, it is the atom with the most available bonding sites (carbon and silicon has the most bonding sites – 4 valence e^-) or it is element with the least number of atoms. Place a pair of electrons to form a single bond between each atom.**
- Arrange any remaining electrons around all atoms such that the octet rule is followed for the second and third periods atoms and duet rule is satisfied for the first period atoms. This may sometimes mean setting up multiple bonds (moving lone pairs to the central atom so that it satisfies the octet rule).**
- Final Lewis structure of Polyatomic Ion must include a big square bracket and the indication of net charge.**

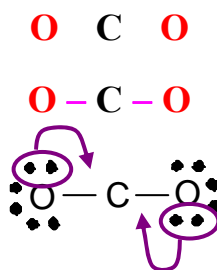
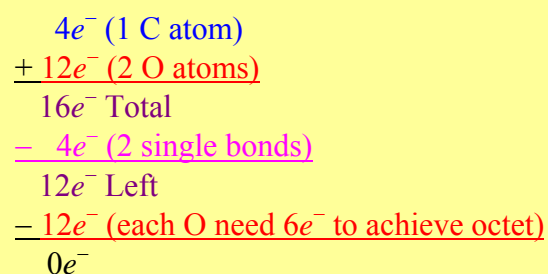
Example 3: Draw the Lewis structure for the following molecules and polyatomic ions.

a. H_2O

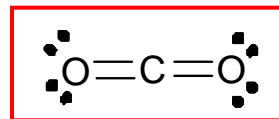


Reason for why the H_2O structure is bent will be explained in section 6.3

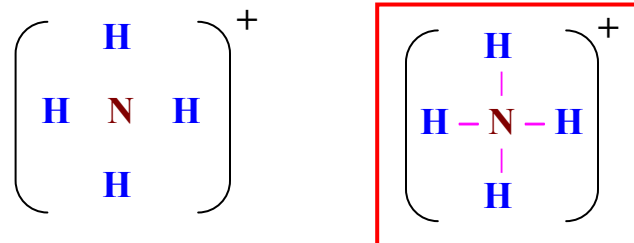
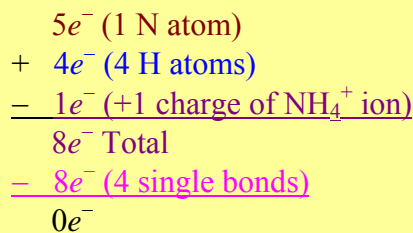
b. CO_2



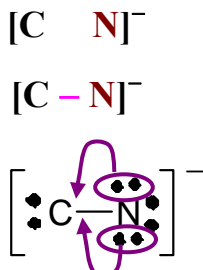
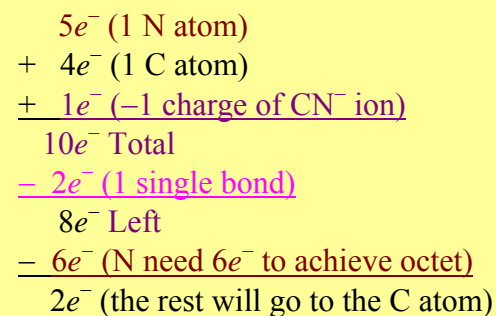
There are only $4e^-$ around C. Oxygen atoms need to contribute unshared e^- pairs (2 lone pairs) to form two double bonds.



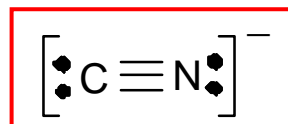
c. NH_4^+



d. CN^-



There are only $4e^-$ around C. Nitrogen atoms need to contribute unshared e^- pairs (2 lone pairs) to form a triple bond.



Resonance: - occurs when more than one Lewis structure exists for a molecule.

Resonance Structure: - different variations of a resonance molecule can be drawn using a double headed arrow, \leftrightarrow , between the diagrams.

Example 4: Draw the Lewis structure for nitrate.

Nitrate = NO_3^-

$5e^-$ (1 N atom)

+ $18e^-$ (3 O atoms)

+ $1e^-$ (-1 charge of NO_3^- ion)

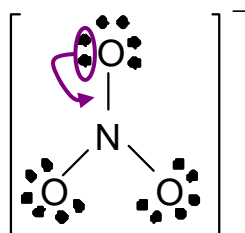
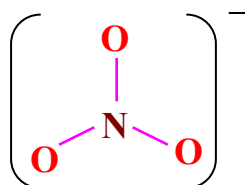
$24e^-$ Total

- $6e^-$ (3 single bond)

$18e^-$ Left

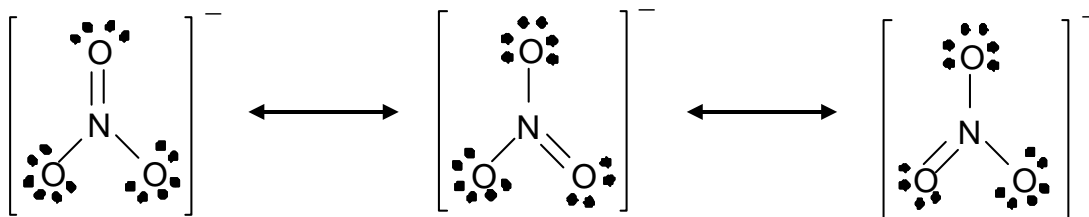
- $18e^-$ (each O atom need $6e^-$ to achieve octet)

$0e^-$



There are only $6e^-$ around N. One oxygen atom needs to contribute unshared e^- pair (1 lone pair) to form a double bond.

There will be three possible structures as the double bond can “rotate” around the central N atom.

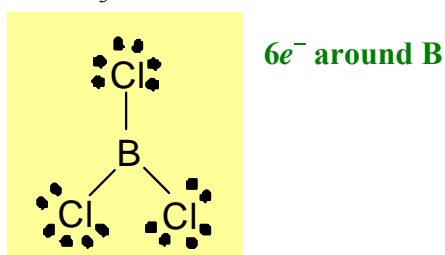


Exceptions to the Octet Rule

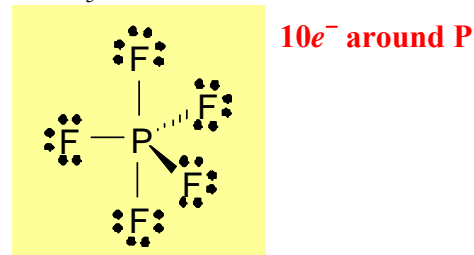
1. In the second row, due to the $2s$ and the $2p$ orbitals, C, N, O, and F should always follow the octet rule. Overall, the second period elements will never exceed the octet rule.
2. Boron (B) follows the $6e^-$ rule.
3. In the third period, heavier atoms towards the right of the Table often satisfy the octet rule. This being said, there are times that they can use the empty $3d$ orbitals to exceed the 8 electrons limit. (Examples: P can have $10e^-$ rule; S and Cl can have $10e^-$ or $12e^-$ rule)
4. Elements in higher row can definitely exceed the octet rule. This sometimes applies to larger noble gases as well. (Examples: Br, I, Xe, Rn can have $10e^-$ or $12e^-$)

Example 5: Draw the Lewis structure for the following molecules.

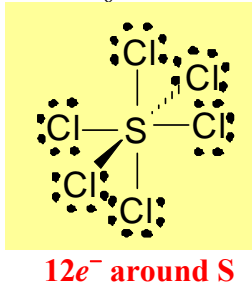
a. BCl_3



b. PF_5



c. SCl_6



Assignment

6.2 pg. 202 #1 & 2 (Practice); pg. 203 #1 & 2 (Practice);
pg. 205 #1 & 2 (Practice); pg. 207 #1 to 13

6.3: Molecular Shape

Valence Shell Electron-Pair Repulsion (VSEPR) Model:


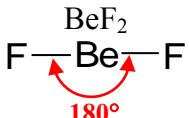
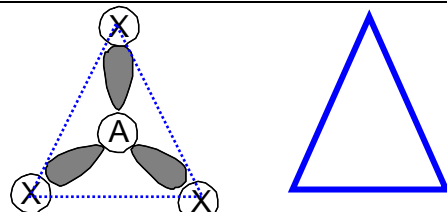
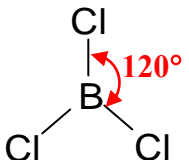
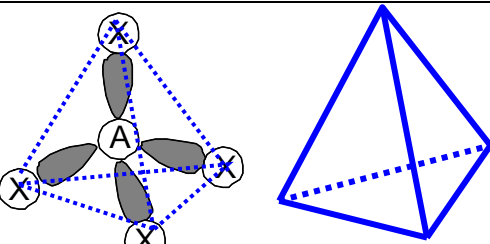
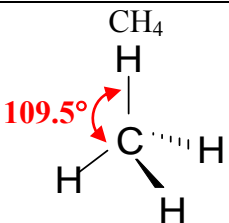
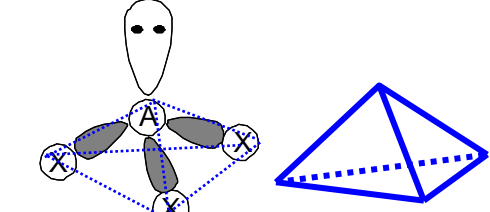
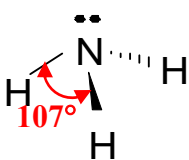
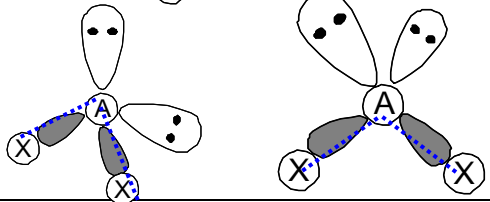
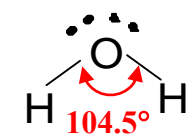
- the best structure for a molecule is one that minimizes electrons lone pairs repulsion.
- most often used to predict molecular structures involving non-metals.

Example: For molecules with a total of 4 e^- pairs, the bond angles decreases from 109.5° as more lone pairs added. (Repulsion of Lone Pair(s) with bond electrons pushed the angle down.)

Effective Electron Pairs:- sometimes refer to as substituents.

- the number of lone pairs on the central atom of a molecule and the number of connections between the central atom with the outer atom(s). Each set of Multiple Bond (like double and triple bonds) count as one connection or one effective electron pair.

Summary of Geometrical Shape of Single Bond Molecules

Around Central Atom		Molecular Structures	Geometrical Shape	Example and Bond Angles
Total # of Eff. e^- Pairs	# of Lone Pairs			
2	0		Linear	BeF_2 
3	0		Trigonal Planar	BCl_3 
4	0		Tetrahedral	CH_4 
4	1		Trigonal pyramid	NH_3 
4	2		V-Shape (Bent)	H_2O 

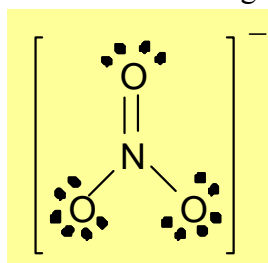
Around Central Atom		Molecular Structures	Geometrical Shape	Example and Bond Angles
Total # of Eff. e^- Pairs	# of Lone Pairs			
5	0		Trigonal bipyramid	PF_5
5	1		See-saw	SF_4
5	2		T-Shape	ClF_3
5	3		Linear	Br_3^-
6	0		Octahedral	SCl_6

Around Central Atom		Molecular Structures	Geometrical Shape	Example and Bond Angles
Total # of Eff. e^- Pairs	# of Lone Pairs			
6	1		Square pyramid	ICl_5
6	2		Square planar	XeF_4

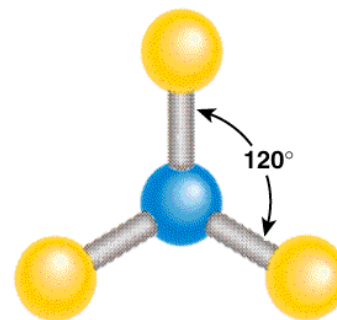
Using VSEPR Model with Multiple Bonds Molecule.

1. Look at multiple bonds as single bonds (only to determine molecule geometry).
2. Using the number of bonds and lone pairs around the central atom, determine the geometry of the molecule from the charts above.

Example 1: From Example 4 of Section 6.2, determine the geometrical shape of NO_3^- . What is the possible bond angle in this complex ion?

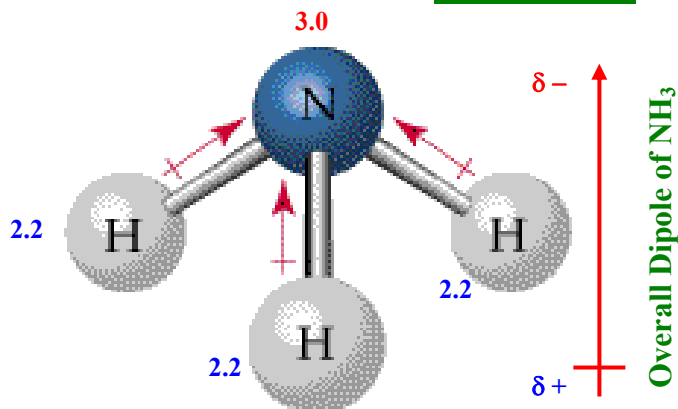


Since there are 3 bonds around nitrogen (we look at the double bond as a single bond when deciding on molecular geometry), and there are no lone pairs around the central atom, the NO_3^- ion is in the shape of trigonal planar with bond angles at 120° .

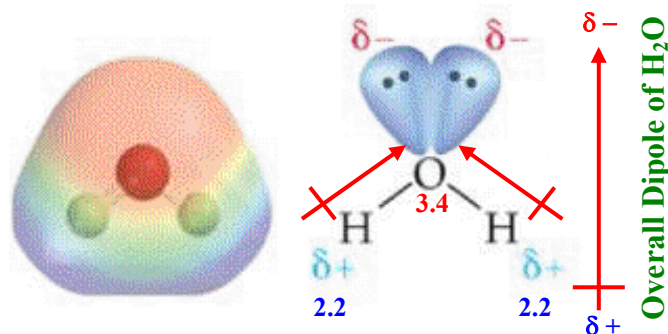


Polarity of Polyatomic Molecules: - when dipoles of each bond are treated like vectors and are added up:

- a **polar molecule** will result when **dipoles do NOT cancel** each other out.

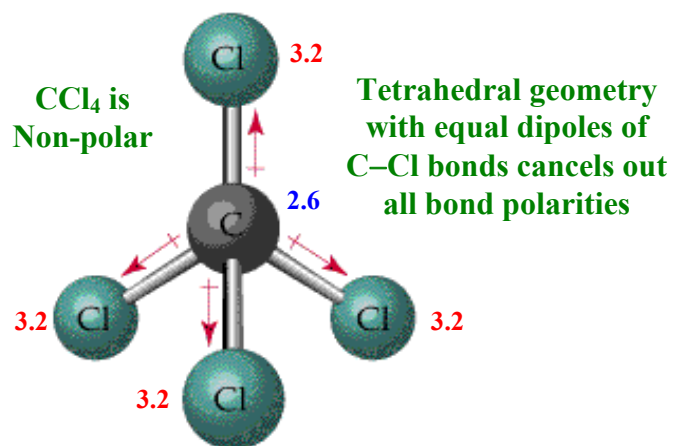
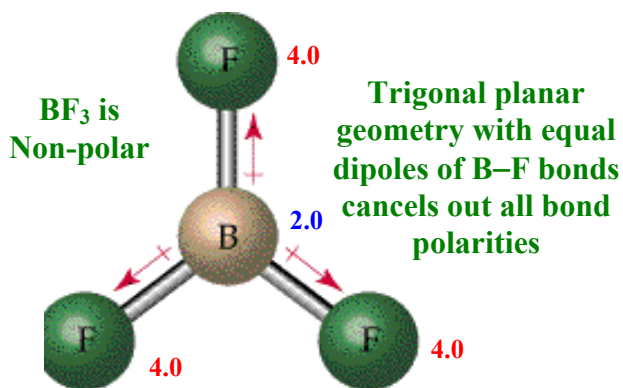


Trigonal pyramid geometry with all N–H bond polarities pointing upward causes an overall dipole for NH₃, making it a polar molecule

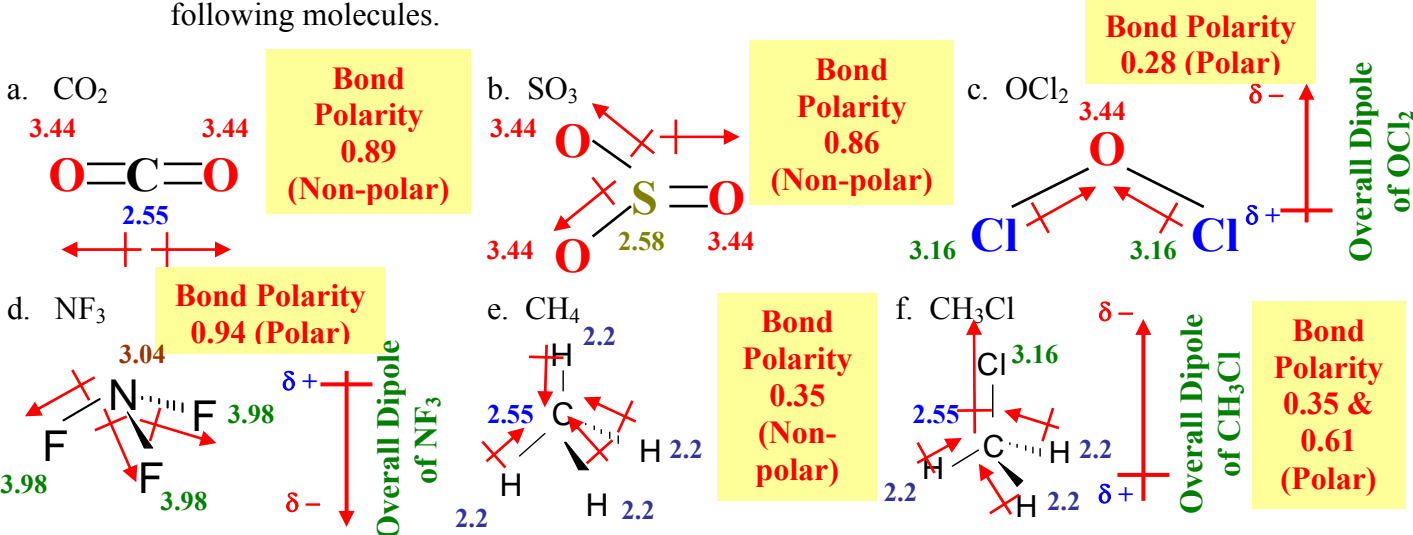


V-shape or Bent geometry with both O–H bond polarities pointing upward causes an overall dipole for H₂O, making it a polar molecule

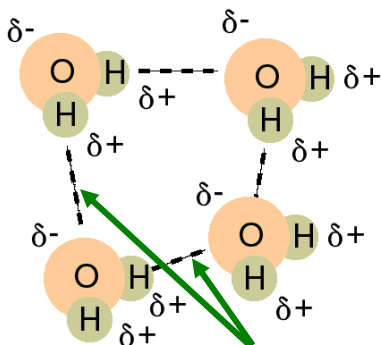
- when the **dipoles cancel** each other out, the **molecule is non-polar**.



Example 2: Using Figure 6.20 of the textbook, determine the bond polarity and dipole (if any) of the following molecules.



Polarity Affects Boiling Point: - the stronger the polarity of a molecule, the higher its boiling point.
 - this is because the **polar molecules** form positive and negative regions so there is **more attraction between molecules (intermolecular forces)**.

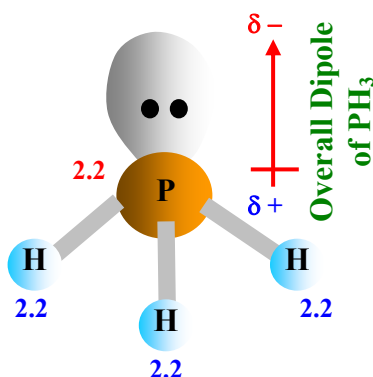


Intermolecular Forces

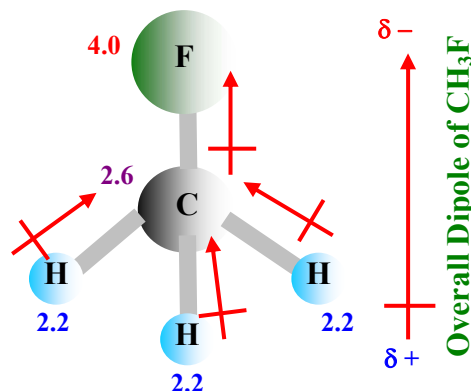
Hence, **more energy (higher temperature) is required to pull these molecules from one another** to change from the liquid phase to the gaseous phase.

Boiling Point of Polar Molecules		Boiling Point of Non-Polar Molecules	
Water (H ₂ O)	100°C	Carbon dioxide (CO ₂)	-78.5°C
Chloroform (CHCl ₃)	62.3°C	Carbon tetrafluoride (CF ₄)	-130°C
Ammonia	-33°C	Fluorine (F ₂)	-188°C

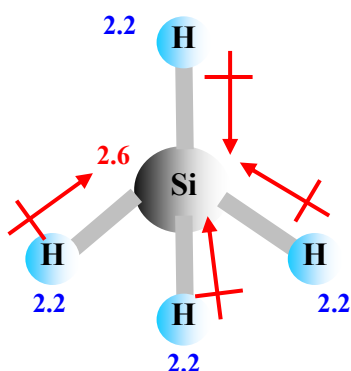
Example 3: Order the boiling points from the least to greatest for the following compounds, PH₃, CH₃F and SiH₄.



PH₃ has a trigonal pyramid geometry (VSEPR) and is Polar. Even though the P-H bonds have no polarity (electronegativities of P and H are the same), the lone pair on one end of the P atom causes an uneven distribution of electrons.



CH₃F has a tetrahedral geometry and is very polar. The C-F bond along with the C-H bonds have strong polarity. The overall dipole for the molecule has electrons around the F atom.



SiH₄ has a tetrahedral geometry with equal dipoles of Si-H bonds cancels out all bond polarities. Hence, SiH₄ is Non-Polar.

Since non-polar molecule have no dipole interactions, SiH₄ should have the lowest boiling point. PH₃ is less polar than CH₃F due to the difference in electronegativities between P-H bond and C-F with C-H bonds. Therefore, CH₃F must have the highest boiling point.

Boiling Point: SiH₄ < PH₃ < CH₃F

Assignment

6.3 pg. 211 (Practice) #1; pg. 213 #1 to 11 and
 Worksheet: Molecular Geometry and VSEPR Theory
 Chapter 6 Review pg. 216-218 #11 to 40, 42, 44 to 46, 48

Worksheet: Molecular Geometry and VSEPR Theory

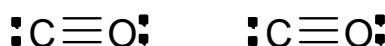
1. For the following molecules and polyatomic ions,
- draw the Electron (Lewis) Dot Diagram (2-Dimensional), show all resonance structures if any;
 - draw the 3-Dimensional Lewis Dot Diagram (with wedged and dotted bond notations);
 - state the Number of Effective Electron Pairs and the Number of Lone Pairs around the central atom;
 - determine the Shape of the Structure and state the bond angles;
 - evaluate (give reasons why) the overall Polarity of the molecule or polyatomic ion.

a. Br ₂	b. CO	c. BF ₃	d. NO ₂ ⁻	e. CF ₄
f. CHBr ₃	g. NF ₃	h. SiO ₃ ²⁻	i. CS ₂	j. H ₂ Te
k. PBr ₅	l. SeI ₄	m. IF ₃	n. BrCl ₂ ⁻	o. SBr ₆
p. ClF ₅	q. XeBr ₄			

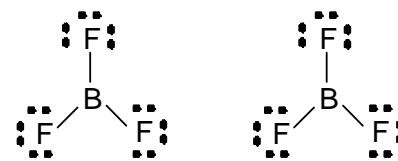
Answersa. Br₂

No central atom; Linear (diatomic molecules); Bond Angle = 180°; Non-polar (both atoms have equal electronegativity values)

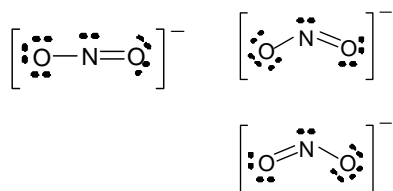
b. CO



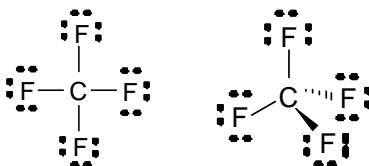
No central atom; Linear (diatomic molecules); Bond Angle = 180°; Polar (both atoms have different electronegativity values)

c. BF₃

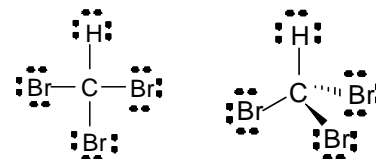
Around central atom B (3 eff e⁻ pairs total, no lone pair); Trigonal Planar; Bond Angles = 120°; Non-polar (no lone pair on central atom **and** all bonds are equivalent)

d. NO₂⁻

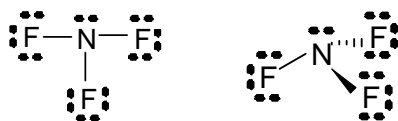
Around central atom N (3 eff e⁻ pairs total, 1 lone pair); Bent; Bond Angles = less than 120° (about 117.5°); Polar (lone pair on central atom)

e. CF₄

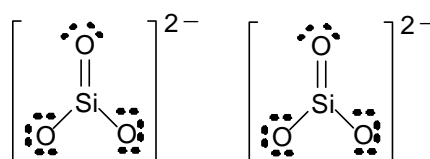
Around central atom C (4 eff e⁻ pairs total, no lone pair); Tetrahedral; Bond Angles = 109.5°; Non-polar (no lone pair on central atom **and** all bonds are equivalent)

f. CHBr₃

Around central atom C (4 eff e⁻ pairs total, no lone pair); Tetrahedral; Bond Angles = 109.5°; Polar (no lone pair on central atom **but not** all bonds are equivalent)

g. NF₃

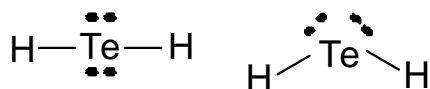
Around central atom N (4 eff e⁻ pairs total, 1 lone pair); Trigonal Pyramid; Bond Angles = 107°; Polar (lone pair on central atom)

h. SiO₃²⁻

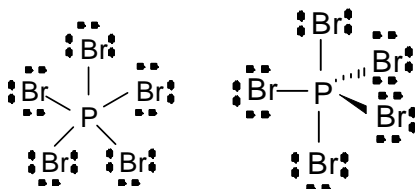
Around central atom Si (3 eff e⁻ pairs total, no lone pair); Trigonal Planar; Bond Angles = 120°; Non-polar (no lone pair on central atom **and** all bonds are equivalent)

i. CS₂

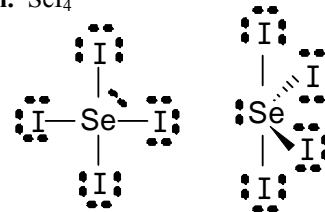
Around central atom C (2 eff e⁻ pairs total, no lone pair); Linear; Bond Angles = 180°; Non-polar (no lone pair on central atom **and** all bonds are equivalent)

j. H_2Te 

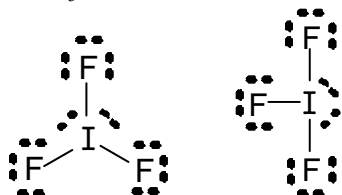
Around central atom Te (4 eff e^- pairs total, 2 lone pairs); V-Shaped or Bent; Bond Angles = 105° ; Polar (lone pairs on central atom)

k. PBr_5 

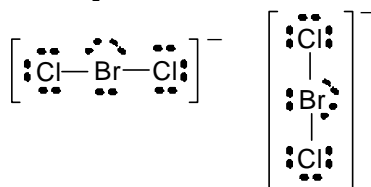
Around central atom P (5 eff e^- pairs total, no lone pair); Trigonal Bipyramid; Bond Angles = 90° and 120° ; Non-polar (no lone pair on central atom **and** all bonds are equivalent)

l. SeI_4 

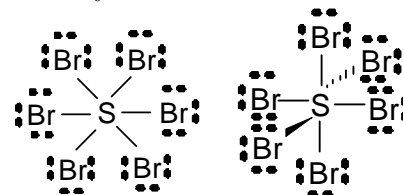
Around central atom Se (5 eff e^- pairs total, 1 lone pair); See-Saw; Bond Angles = 88.5° and 118.5° ; Polar (lone pair on central atom)

m. IF_3 

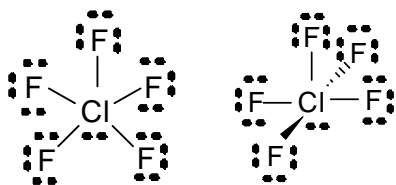
Around central atom I (5 eff e^- pairs total, 2 lone pairs); T-shaped; Bond Angles = 88.5° and 183° ; Polar (lone pairs on central atom)

n. BrCl_2^- 

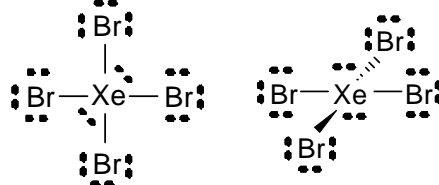
Around central atom Br (5 eff e^- pairs total, 3 lone pairs); Linear; Bond Angles = 180° ; Non-polar (even with lone pairs on central atom, the lone pairs cancelled out with each other **and** all other bonds are equivalent)

o. SBr_6 

Around central atom S (6 eff e^- pairs total, 0 lone pair); Octahedral; Bond Angles = 90° ; Non-polar (no lone pair on central atom **and** all bonds are equivalent)

p. ClF_5 

Around central atom Cl (6 eff e^- pairs total, 1 lone pair); Square-Planar; Bond Angles = 88.5° ; Polar (lone pair on central atom)

q. XeBr_4 

Around central atom Xe (6 eff e^- pairs total, 2 lone pairs); Square Planar; Bond Angles = 90° ; Non-polar (even with lone pairs on central atom, the lone pairs cancelled out with each other **and** all other bonds are equivalent)