CHEM 1212
General Chemistry I Refresher

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Section: **Significant Figures**

I. **Significant figures** involve the numbers in a measured quantity or value that are known to be correct and one digit that is not known for sure. A common application of significant figures happens nearly every day in lab with reading glassware.

- **a.** To the right is a graduated cylinder. Each mark on the graduated cylinder represents a 0.1 mL increase in volume.
- **b.** With the graduated cylinder to the right, it is known for sure that the true liquid measurement lies between 15.0 mL and 15.1 mL. We are confident in the value of the ones and tenths place, but the digit in the hundredths place is an approximation.
- **c.** An appropriate guess for liquid in this graduated cylinder would be 15.02 mL OR 15.03 mL. Either answer would be fine since the digit in the hundredths place is an approximation.

Figure 1 Work by "Horia Varlan"/Flickr https://creativecommons.org/licenses/by/2.0/legalcode

II. **Rules for Significant Figures**:

- **a.** All non-zero numbers are significant.
- **b.** Zeroes between two other significant digits are significant, including numbers that include a decimal place (e.g. 2.**0**34, 1**0**14**0**3).
- **c.** Zeroes following a non-zero number that are also to the left of a decimal are significant (83**000**.)
- **d.** In numbers containing a decimal, all zeros at the end of the number are significant (0.023**0**)
  - Combining rules 2 and 4 shows us that the 3 zeros at the end of 0.071**000** are significant.
- **e.** Zeroes that do not have either a decimal point or non-zero digit to the right of them are "trailing" zeros and are ***not*** significant. (e.g. 32**0**)
- **f.** Zeroes that occur before any non-zero number are ***not*** significant (**0**.**00**147)

III. **Exceptions to traditional rules:**

- **a.** Counting numbers have an unlimited number of significant figures, meaning that there is no way to make them more precise than they already are. Examples of this would be counting 12 pencils or saying that a molecule is made of three atoms.
- **b.** Known conversions also fall in this category. For example, 1 foot = 12 inches is an example of an errorless conversion.

Example: Determine the number of Significant figures contained within each of the following numbers:

a) 0.0900 ____________________
b) 6.230 ____________________
c) 0.0076 ____________________
d) 4.00028 ____________________
e) 4.335 x10^{-22} ____________________
f) 5600 ____________________
g) 3 Science Guyz Packets ____________________

IV. Rounding following Mathematical Operations:

- **a.** Following any mathematical operation, you must determine how many sig figs will be in your answer using sig fig rules.
- **b.** Once you have deteremined the correct number of sig figs, you must round your answer to the correct number of significant figures by looking to the first non-significant digit in your answer.
- **c.** If this number is 5 or greater, you will round your last significant digit up, otherwise you will round you answer down to the first significant digit to the left of the non significant digit.

V. Sig Figs in Multiplication and Division Problems:

- **a.** When multiplying/dividing two or more numbers, to express your answer in the correct number of sig figs, your answer must contain the same number of sig figs as the value in your equation with the fewest number of sig figs.

Example: Perform the operation below and report your answer to the correct number of significant figures.

(3.2005470) x (30.9) = ____________

VI. Sig Figs in Addition and Subtraction Problems:

- **a.** When adding or subtracting two or more numbers, to express your answer in the correct number of sig figs, your answer must contain the same number of decimal places as the value in the equation with the fewest decimal places.

Example: Perform the operation below and report your answer to the correct number of significant figures.

(321.1896) + (1.98665) + (0.1) = ____________

VII. Sig Figs with Multiple Operations:

- **a.** You will often find yourself performing operations that involve both addition/subtraction rules and multiplication/division rules. Perform operations in order as dictated by PEMDAS and carry the exact numbers that you obtain throughout the operation. We will only consider significant figures at the end. Consider the examples below:

Example: Perform the operation below and report your answer to the correct number of significant figures.

a) (2.9 x 4.719) + 12.710 = ____________

b) (30.0031 + 0.3) (6.211 – 6.185)/ (5.233 x 10^{-2}) = ____________

Section: Common SI Base Units Used in Chemistry

I. **The International System of Units, or SI**, is the scientific system for measurements. Most measurements in Chemistry are made in SI units and it is important to know the SI unit used for each property provided below:

Property	Unit Used
Mass	Kilogram (kg)
Length	Meter (m)
Time	Second (s)
Temperature	Kelvin (K)
Amount of Substance	Mole (mol)

II. **The Metric System (Commit to Memory!):**

Prefix	Abbreviation	Meaning	Example
Giga-	G	10^9 (billion)	1 gigahertz = 1 x 10^9 Hz
Mega-	M	10^6 (million)	1 megaton = 1 x 10^6 tons
Kilo-	k	10^3 (thousand)	1 kilogram (kg) = 1 x 10^3 g
Dec-	d	10^{-1} (tenth)	1 decimeter = (dm) 1 x 10^{-1} m
Centi-	c	10^{-2} (hundredth)	1 centimeter = (cm) 1 x 10^{-2} m
Milli-	m	10^{-3} (thousandth)	1 millimeter = (mm) 1 x 10^{-3} m
Micro	µ	10^{-6} (millionth)	1 micrometer = (µm) 1 x 10^{-6} m
Pico-	p	10^{-12}	1 picometer = (pm) 1 x 10^{-12} m
Femto-	f	10^{-15}	1 femtometer = (fm) 1 x 10^{-15} m

III. **Other Useful Conversions:**

a. Several useful conversions are provided below. Depending on when you are taking this class you may be expected to memorize all conversions, or your instructor may give you certain ones. Ask your instructor for their expectations on the conversions!

1 kilometer = 0.62137 mile	1 mile = 5,280 feet
1 meter = 3.281 feet	1 inch = 2.54 cm
1 cm^3 = 1 mL	1 Angstrom (Å) = 1 x 10^{-10} m
1 pound = 453.59 g = 16 ounces	1 mile = 1760 yards

Example: The mass of a cat is most appropriately measured using the ___________ metric unit.
a) Grams. b) Milligrams. c) Kilograms. d) Decigrams. e) Centigrams.

Example: The length of a cat is most appropriately measured using the ___________ metric unit.
a) Meter. b) Millimeter. c) Kilometer. d) Decimeter. e) Centimeter.

Section: Dimensional Analysis

I. In Chemistry and other STEM courses, you will often have to convert from one unit to another. This is sometimes a multistep process that can become tedious. **Dimensional analysis** provides a strategy to organize these conversions.

1. When unit conversions are needed, first establish what your final units must be in.
2. Consider the numerator/denominator relationship of the final unit.
    a. For example, miles per hour requires miles in the numerator and hours in the denominator (miles / hour). So, the result of your dimensional analysis conversions must reflect this.
3. Choose a value as a starting point for your conversion. This is the first blank in "I am converting ____ to ____." It is usually helpful to start with a given value that only has one unit associated with it if possible.
4. Set up a series of conversion steps that allow you to cancel all unwanted units and leave you only with the final units.
    a. Each conversion step is a fraction. Examples include conversion factors (such as 1 hour / 60 seconds) or properties like density (grams / liter).
    b. Units can be canceled when they are found in the numerator and denominator of a dimensional analysis calculation, *regardless of if they are in consecutive steps.*
5. Once your units have been canceled to your final units, multiply across the numerators and denominators, divide the products, and you are finished.

Example: How many picograms are in a milligram?

Example: How many millimeters are in 4.20 centimeters?

Example: Convert 5.10 meters per second to micrometers per hour.

Example: Kailee is travelling to Atlanta. One gallon of gas will allow her to travel for 8.5 miles. If it takes her 95 minutes to get to Atlanta and she travels at a speed of 65 miles/hour, how many gallons of gas does she use?

Example: You are a medical assistant and need to administer medication to help a patient with hip pain. The recommended dosage of this medication is 6.71 mg/kg of body mass. What would be the dosage in milligrams for a 225-lb individual? Note: 1 lb = 453.59 g.

Example: Ben is moving up in the world and has decided it is time to install a pool. Ben wants the pool to have the following dimensions: 58.7 ft long and 27.0 ft wide and 9.50 ft deep. Once the pool has been completed, how much water will Ben need to fill the pool (in cubic inches)?

Example: Convert the cubic inches obtained in the previous problem into cubic centimeters. Note: 1 inch = 2.54 cm.

Section: Introduction to Stoichiometry

I. **Stoichiometry** involves using the molar ratios (coefficients) of a balanced chemical equation to calculate and predict things about the chemical reaction represented by the equation.

Example: Barium nitride reacts with water to form barium hydroxide and ammonia. Use this information to answer the following questions.

a) Write a balanced chemical equation from the reaction described.

b) How many moles of water are required to react with 4.0 moles of barium nitride?

Example: Aluminum carbonate reacts with phosphoric acid to produce aluminum phosphate, carbon dioxide, and water. Use this information to answer the following questions.

a) Write a balanced chemical equation from the reaction described.

b) What mass (in grams) of aluminum carbonate is required to completely react with 156.7 g of phosphoric acid?

c) What mass (in grams) of carbon dioxide is produced from the reaction of 200.8 g of phosphoric acid with excess aluminum carbonate?

Example: Ammonia reacts with elemental oxygen to form elemental nitrogen and water. Use this information to answer the following questions.

a) Write a balanced chemical equation from the reaction described.

b) How many molecules of ammonia are required to react with excess oxygen to form 25.64 grams of nitrogen?

c) How many hydrogen atoms are present on the product side when 148.5 grams of oxygen reacts with excess ammonia?

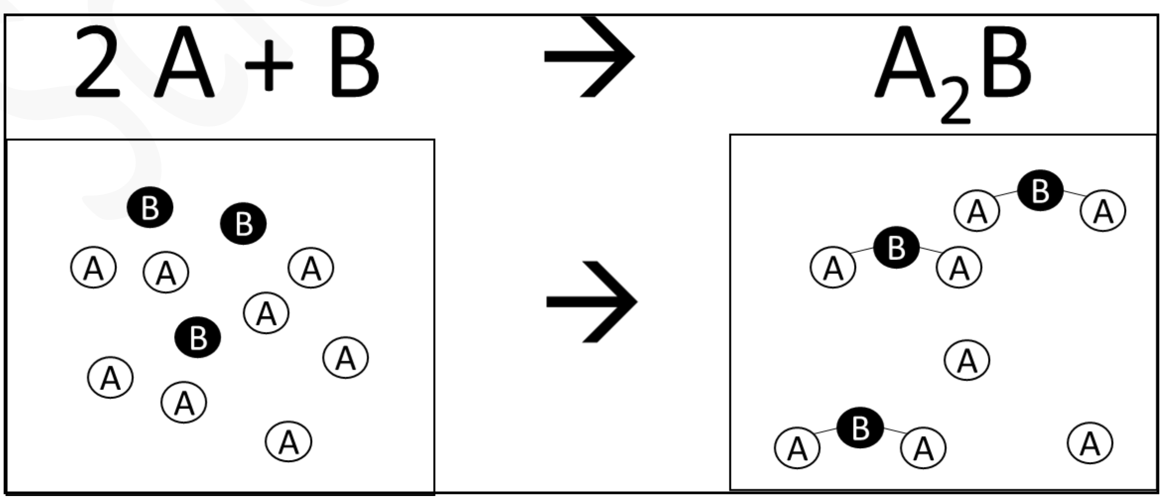
<u>Section</u>: More on Stoichiometry – Limiting and Excess Reactants

I. Up to this point, we have exclusively dealt with stoichiometry problems where one of the reactants is present in excess. Often with stoichiometry problems, you will not be told which reactant will run out first and must determine the limiting reactant and excess reactant.

- **a.** **Limiting Reactants**: This is the reactant that will run out before the other reactants, and therefore *limits* how much of the reaction can take place.
  - **i.** The limiting reactant(s) will always be used up completely, at which point the reaction stops. **Thus, all stoichiometric calculations must be based on the limiting reactant.**
  - **ii.** To determine which reactant is limiting, determine which reactant will produce the smallest quantity of any single product of the reaction. If there are multiple products in the reaction, any product can be used for determining the limiting reactant. However, you must take care to use the same product for each reactant.
  - **iii.** The original amount in grams of each reactant should not be used when determining the limiting reactant!
  - **iv.** There may be cases where all reactants are used up at the same time. In these cases, there would be no limiting reactant.
- **b.** **Excess Reactants**: Once you establish the limiting reactant(s), all other reactants can be considered excess reactants. This means that there will be some of these reactants left over once the reaction has gone to completion.
  - **i.** We can calculate the amount of an excess reactant remaining using the following general formula:

Excess Reactant Remaining = Starting Amount of Excess Reactant - Excess Reactant Consumed

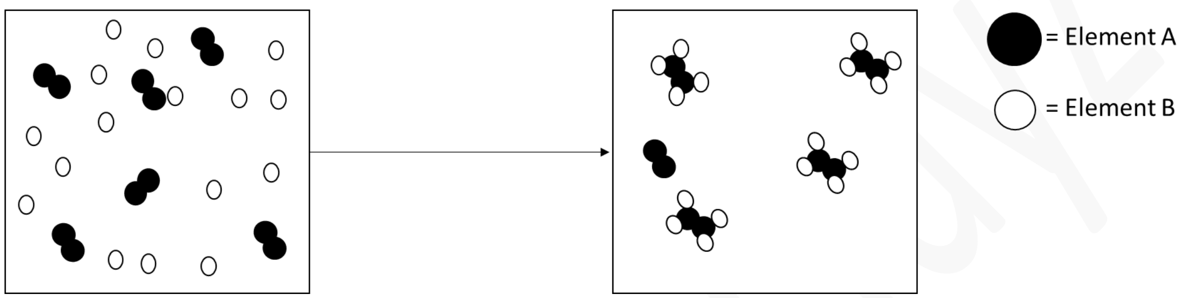
  - **ii.** Use the limiting reactant to determine how much of the excess reactant(s) is consumed.
- **c.** In the diagram below, two atoms of A react with one atom of B to form one molecule of A_2B. All the B is used up first, making it the limiting reactant, and some molecules of A remain, making it the excess reactant.
  - **i.** Without more B, no further reaction can occur.



II. General Steps for Solving Stoichiometry Problems

1. Write a balanced equation. **Always double-check your equation to make sure it is balanced**.
2. Determine which reactant is the limiting reactant.
3. Convert all the quantities given to you in the problem into moles to use the molar ratios of the balanced equation.
4. Use the limiting reactant and the molar ratios (coefficients) to calculate what is required. Using dimensional analysis, you can calculate the moles, mass, volume, number of molecules, or other things about any substance in the reaction using stoichiometry.

Example: A diagram of a chemical reaction is provided below. Use this diagram to answer the next set of questions.

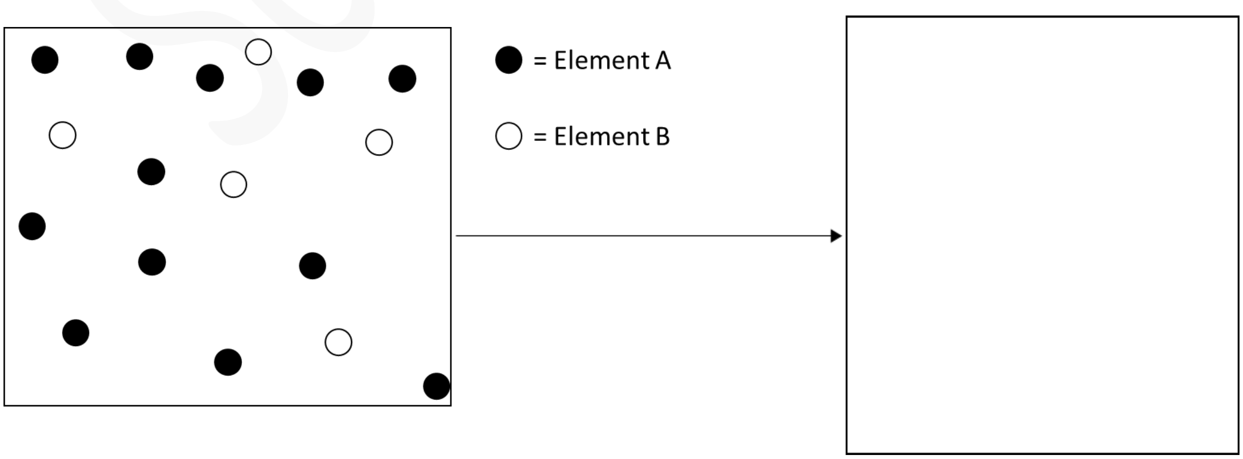


a) Write out a balanced chemical equation for the reaction depicted in the diagram.

b) Are we able to determine the limiting and excess reactant from the diagram? If so, what are they?

c) If you were to react 5.0 moles of element A and 5.0 moles of element B, how much product would you expect to form?

Example: Element A reacts with element B to form the molecule A_2B. A reaction vessel is shown below containing element A and element B before any reaction has occurred. What does the reaction vessel look like after the reaction has gone to completion?



Example: An unbalanced chemical reaction being explored in the lab is provided below. In the reaction, 40.6 g of silver nitrate is combined with 64.1 g of barium chloride to form the products shown. Use this setup to answer the following questions.

$$AgNO_3(aq) + BaCl_2(aq) \rightarrow AgCl(s) + Ba(NO_3)_2(aq)$$

a) What is the **maximum** amount of silver chloride that could be produced from this reaction?

b) What is the excess reactant and what is the amount of the excess reactant remaining after the reaction has gone to completion?

Example: 3.463 grams of solid iron (III) oxide is combined with 2.097 grams of gaseous carbon monoxide to produce solid iron and gaseous carbon dioxide. How many grams of carbon dioxide will be produced and what mass of the excess reactant will remain after the reaction has gone to completion?

III. It is possible to determine the identity of an unknown atom by knowing how the atom reacts in a known reaction and employing the stoichiometry of the known balanced chemical reaction and the law of conservation of mass.

Example: A 3.560 gram sample of $CuSO_4 . x H_2O$ yields 2.276 grams of $CuSO_4$ when heated. Determine the value of x.

Example: You are provided with a 2.680-gram sample of some metal carbonate. The metal within the compound is currently unknown, depicted by "X." The metal carbonate is heated and is observed to give 1.501-grams of a metal oxide. The balanced reaction for the reaction described is provided below. Using this information, determine the identity of the unknown metal.

$$XCO_3(s) + \text{heat} \rightarrow XO(s) + CO_2(g)$$

Section: Percent Yield

I. In perfect conditions, the amount of product formed from a chemical reaction would always equal the maximum amount of product that should form based on the limiting reactant based on the law of conservation of mass.

- **a.** In practice this is not the case. The amount of a product formed during an *actual* chemical reaction is usually *smaller* than the amount predicted by stoichiometry.
  - **i.** **Theoretical yield** is *always* determined by the stoichiometry of a balanced chemical reaction.
  - **ii.** **Actual yield** is *only* determined experimentally and may be provided to you in a problem.
  - **iii.** **Percent yield** describes the efficiency of a reaction and is given by the following equation. This equation must be committed to memory.

$$\boxed{\boldsymbol{Percent\ Yield = \frac{Actual\ Yield}{Theoretical\ Yield} \times 100}}$$

Example: Calcium hydroxide reacts with excess hydrobromic acid to form calcium bromide and water. In a general chemistry lab, a student reacts 130.5 g of calcium hydroxide with excess hydrobromic acid. After the reaction has completed, the student collects 50.60 g of calcium bromide. Determine the percent yield for the reaction.

Example: 5.55 grams of solid aluminum reacts with 5.55 grams of chlorine gas to produce solid aluminum chloride. If 4.98 grams of solid aluminum chloride is isolated once the reaction has gone to completion, what is the percent yield?

Example: A reaction to produce silver chloride is provided below. If the reaction were to be executed using 56.78 grams of sodium chloride in the presence of excess silver nitrate, how many grams of silver chloride were collected if the percent yield was 54.67%

$$AgNO_3\ (aq) + NaCl\ (aq) \rightarrow AgCl\ (s) + NaNO_3\ (aq)$$

Example: In your chemistry lab, you perform an experiment involving this reaction and obtain a percent yield of 71.2%. How many grams of H_2S are needed to obtain 63.51 g of SO_2?

$$H_2S\ (aq) + O_2\ (g) \rightarrow H_2O\ (l) + SO_2\ (g)$$

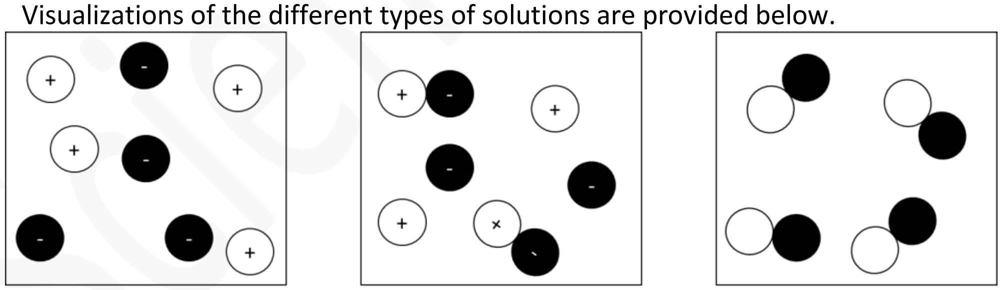
<u>Section</u>: An Introduction to Solutions

I. A **solution** is defined as a homogeneous mixture of two or more substances that come together to form a single phase.

- **a.** A solution consists of a **solute**, which is the actual substance that experiences a phase change after being dissolved in a **solvent**, which is the media that remains in that same phase after a solution is formed.
  - **i.** For example, in a solution of saltwater we would say that the solute is salt, and the solvent is water. The salt was in a solid phase before it was added to the water and the water remained in the liquid phase after the salt water was formed.
  - **ii.** In liquid solutions, the solvent is identified as the substance present in the greatest amount.
  - **iii.** **Aqueous solutions** correspond to solutions in which some solute has been dissolved in water. Many reactions occur in water because it can serve as a non-participating solvent.

II. Solutes can be categorized as electrolytes or non-electrolytes**.**

- **a.** **Non-Electrolytes**: Compounds that dissolve in water but do not **ionize** (split into ions) or compounds that simply do not dissolve in water at all. Molecular substances will typically remain intact in solution and most are considered non-electrolytes.
- **b.** **Electrolytes**: Compounds that are soluble and ionize in water. There are two types of electrolytes that you should be familiar with.
  - **i.** **Strong Electrolytes**: Strong electrolytes are assumed to ionize completely when dissolved in water. Ionic compounds deemed soluble via solubility rules, strong acids, and strong bases are considered strong electrolytes. Solutions formed from strong electrolytes conduct electricity strongly.
  - **ii.** **Weak Electrolytes**: Insoluble compounds which minimally ionize in solution. Weak electrolytes conduct electricity minimally. Weak acids and weak bases are weak electrolytes.
- **c.** Visualizations of the different types of solutions are provided below.



Strong Electrolyte | Weak Electrolyte | Non-Electrolyte

- **d.** Visualizations of what equations look like written out for strong and weak electrolytes are provided below.

$$NaCl_{(aq)} \longrightarrow Na^{\oplus}_{(aq)} + Cl^{\ominus}_{(aq)}$$

Strong Electrolytes

$$HF_{(aq)} \rightleftharpoons F^{\ominus}_{(aq)} + H^{\oplus}_{(aq)}$$

Weak Electrolytes

<u>Example</u>: Write out the equation that **best** represents the dissolution of ammonium bromide in water.

<u>Example</u>: Write out the equation that **best** represents the dissolution of sodium phosphate in water.

<u>Example</u>: Write out the equation that **best** represents the dissolution of mercury(I) acetate in water.

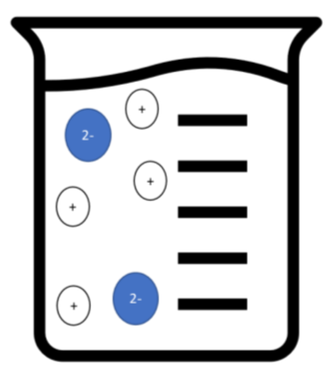
<u>Example</u>: A researcher mixes 100. mL of water with 0.500 L of methanol. Which is the solvent? Explain.

<u>Example</u>: Determine whether each of the following compounds is a strong electrolyte, a weak electrolyte, or a nonelectrolyte.
a) $Ba(NO_3)_2$ ____.
b) NH_3 (weak base)____.
c) CH_3OH ____.
d) CH_4 ____.
e) HCOOH (weak acid) ____.
f) HNO_3 (strong acid) ____.
g) HF (weak acid)____.

<u>Example</u>: Of the options provided below, which is/are **strong** electrolytes? *<u>Select all that apply</u>*.
a) CH_3COOH (weak acid).
b) HBr (strong acid).
c) KCl.
d) CH_3CH_3.
e) NH_4Br.

<u>Example</u>: Which of the following would be the **poorest** conductor of electricity when forming an aqueous solution?
a) Magnesium chloride.
b) Hydrochloric acid.
c) Methanoic acid.
d) Sodium nitrate.
e) Ethanol.

<u>Example</u>: Which compound provided below could be represented by the beaker shown to the right?
a) Magnesium chloride.
b) Sodium sulfide.
c) Potassium bromide.
d) Calcium carbonate.
e) Any of the above options could be represented by the beaker.



<u>Section</u>: Solubility Rules and Precipitation Reactions

I. Many of the ionic compounds that we have explored so far are highly soluble in water. **Solubility** is the extent to which a solute can be dissolved in a solvent under certain conditions. Not all compounds in chemistry are fully soluble in water, however.

- **a.** For example, calcium carbonate ($CaCO_3$), which makes up shells of many marine organisms and is a component of antacid medication, is sparingly soluble in water.

II. It may seem daunting at first to think about all the compounds that you have to memorize the solubility for. Fortunately, there is a general set of rules that you can use when determining whether a compound will be soluble in aqueous solution or not. **Note: If the list above differs from the one given by your instructor, use the rules given by your instructor!**

1. Most nitrate, acetate, chlorate, and perchlorate salts are soluble.
2. Most salts containing the alkali metal ions and the ammonium ion are soluble.
3. Most chloride, bromide, and iodide salts are soluble. Notable exceptions are salts containing silver, lead, and mercury(I).
4. Most fluoride salts are soluble. Notable exceptions include magnesium, calcium, strontium, barium, and lead.
5. Most sulfate salts are soluble, except for sulfate salts containing barium, lead, mercury(I), strontium, silver, and calcium.
6. Most hydroxide salts are insoluble except for the strong bases.
7. Most sulfide, carbonate, chromate, oxalate, and phosphate salts are insoluble, except for those containing the ions mentioned in rule 2. Note that sulfides containing calcium, strontium, and barium are also soluble.

<u>Example</u>: Determine whether each compound provided below is soluble or insoluble in water.

a) $MgCO_3$: ______________________
b) AgI: ______________________
c) $CaSO_4$: ______________________
d) KCl: ______________________
e) Ag_3PO_4: ______________________
f) Ammonium carbonate: ______________________
g) Mercury(I) chloride: ______________________
h) Mercury(II) chloride: ______________________
i) Barium sulfide: ______________________
j) Rubidium sulfate: ______________________

III. **Precipitation reactions** are reactions in which two water-soluble compounds react in a double displacement reaction forming at least one insoluble (solid) product, which then precipitates (or falls out) of the solution.

- **a.** **Double Displacement Reaction**: A chemical reaction where two reactant ionic compounds exchange ions to form two new compounds with the same ions.

$$A^+B^- + C^+D^- \rightarrow A^+D^- + C^+B^-$$

- **b.** An example of a precipitation reaction is provided below:

$$3\ NaOH\ (aq) + FeCl_3\ (aq) \rightarrow Fe(OH)_3\ (s) + 3\ NaCl\ (aq)$$

<u>Example</u>: Lead (II) nitrate and potassium carbonate are placed in a reaction vessel. Predict the products of this reaction (including states) and write a balanced chemical equation for the reaction.

Example: Determine if a precipitate will form in each reaction and if so, what that precipitate will be.

Reactants Mixed	**Precipitate Formation?**
a) Nickel (II) nitrate & sodium carbonate	________________________
b) Sodium sulfate & barium chloride	________________________
c) Ammonium chloride & silver nitrate	________________________
d) Sodium phosphate & barium chloride	________________________
e) Lead (II) nitrate + sodium bromide	________________________
f) Potassium chloride + sodium nitrate	________________________

IV. A **net ionic equation** shows a reaction taking place in a solution. Common ions (known as **spectator ions**) present on the reactant and product side of a reaction are omitted. The steps for writing a net ionic equation are provided below.

1. Draw out the reaction given as a double displacement using the subscripts “aq” for aqueous (soluble) and “s” for solid (insoluble). These will be based on the solubility rules shown previously.
2. Balance the equation.
3. Draw the ionic equation by splitting all soluble compounds into their individual ions. This will represent your **total ionic equation**.
4. Cancel any spectator ions that appear on both sides of the reaction arrow. These ions are not partaking in the chemical reaction.
5. Remember that insoluble products will not be drawn as ions and thus cannot act as spectators.
6. If all reactants and products are soluble, then we say there is no reaction, and thus no net ionic equation. All ions act as spectator ions in this case.

Example: In a reaction vessel, potassium iodide is observed to react with lead (II) nitrate. Write the total ionic equation for this reaction and identify the spectator ions.

Example: Sodium carbonate is mixed with iron (III) acetate. What is the net ionic equation for the reaction that occurs?

Example: Silver nitrate reacts with sodium iodide. What is the sum of the coefficients of the balanced net ionic equation?

Example: Silver acetate reacts with potassium nitrate. What is the net ionic equation for the reaction that occurs?

<u>Section</u>: Specific Heat and Thermal Equilibrium

I. **Specific Heat Capacity (C)** is defined as the energy transferred as heat that is required to raise the temperature of 1 g of a substance by 1 ^oC (or 1 K). Specific heat capacity often has the unit J/g · ^oC (or J/g · K).

- **a.** The greater the heat capacity of a substance, the more energy that is required to change its temperature.
- **b.** The only value for specific heat that you should commit to memory is the specific heat for water, which is 4.184 J/g · ^oC.

<u>Example</u>: You are provided with four unknown metals with known heat capacities, provided in the chart. Each sample of metal has the same mass. Use this setup to answer the following questions.

Metal	Heat Capacity (J/g·^oC)
A	0.478
B	0.202
C	0.300
D	0.375

a) Each of the samples starts at room temperature and all have equal amounts of heat added to them over the course of a few minutes. Rank the metal samples according to the temperature after a few minutes have passed.

b) If each sample of metal is heated to 150. ^oC and placed in a bucket of water, which metal would cause the **greatest** rise in temperature?

II. The energy (as heat) gained or released when changing the temperature of some substance is given by the following equation:

$$\boxed{\begin{array}{c} q = mC\Delta T \\ \Delta T = T_f - T_i \end{array}}$$

- **a.** Where q is heat, m is the mass (usually in g, but pay attention to units of your heat capacity), C is the specific heat capacity, and ΔT is the temperature change.
- **b.** Although two or more substances which are mixed may have different initial temperatures, when the substances reach thermal equilibrium, both substances will have the same final temperature.
- **c.** Note that heat will be negative if heat is lost by the system and positive if heat is gained by the system.

<u>Example</u>: In your general chemistry lab, you are attempting to determine the energy change of a variety of different metals under different conditions.

a) Determine the energy change when 50.0 grams of copper (specific heat capacity is 0.385 J/g·^oC) is decreased from 50.0 ^oC to 30.0 ^oC.

b) Determine the energy change when 0.500 mol of zinc (specific heat capacity is 0.390 J/g·^oC) is increased from 50.0 ^oC to 75.0 ^oC.

III. **Thermal equilibrium** can be understood by the following equation representing several substances exchanging heat to reach thermal equilibrium. Below, q represents the energy associated with the change in thermal energy of each substance.

$$\boxed{q_1 + q_2 \ldots. + q_x = 0}$$

a. The sum of energies is equal to zero due to the first law of thermodynamics stating that energy cannot be created or destroyed.

i. For example, if one system loses energy, the second must gain energy that is equal to the amount the first system lost. This equation is often used when attempting to determine the energy change of two components of a system, such as when a hot metal is dumped into a beaker of cool water. For this example, we would say that:

$$\boxed{q_{metal} + q_{water} = 0 \qquad\qquad q_{water} = -q_{metal}}$$

ii. The above equation makes sense because when the hot metal is placed in the beaker of water, the metal will transfer heat energy (exothermic) to the water, which will then gain the heat energy (endothermic).

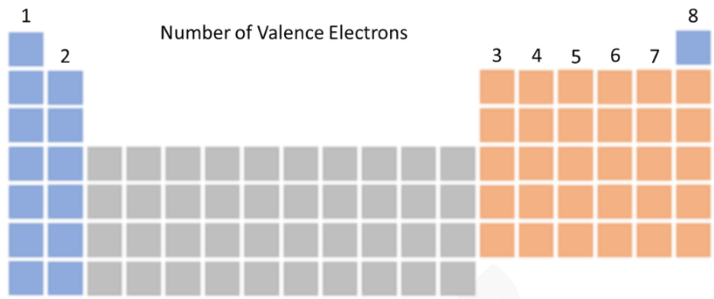
Example: You wake up and are preparing your morning coffee before you head off to your chemistry class. You place a 30.2 g aluminum spoon (specific capacity heat is 0.900 J/g.°C) into your 250 mL (240 g) cup of coffee. The spoon is initially at 25.0 °C and the cup of coffee is originally at 98.0 °C. You place the spoon into the coffee and allow them to reach thermal equilibrium. Determine the temperature (in °C) when the spoon and coffee have reached thermal equilibrium. Note: Consider the specific heat capacity of coffee to be the same as water and that no energy is lost to the surroundings.

Example: In your chemistry lab, you heat a 17.1 g sample of chromium (specific heat is 0.450 J/g.°C) to 96.8 °C. The heated chromium is then placed in 40.00 grams of water that is at 25.0 °C at the start of the experiment. What is the temperature of the water and metal after thermal equilibrium is achieved?

Example: A 25.5-gram sample of some metal X is heated to 150. ^{o}C and then dropped into 4.00 mols of water at 25.0 ^{o}C. Once thermal equilibrium has been achieved the final temperature is measured to be 30.1 ^{o}C. Determine the specific heat capacity of metal X.

<u>Section</u>: Introduction to Lewis Structures

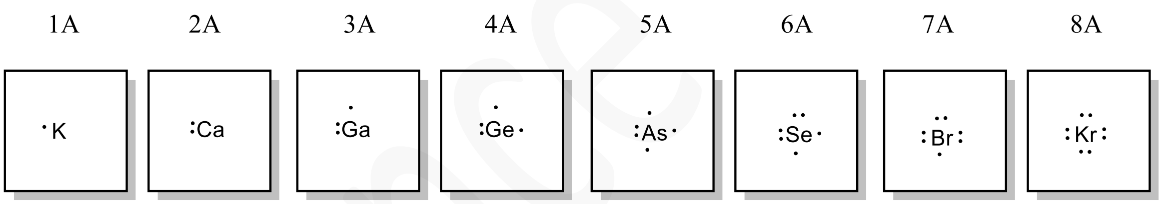
I. **Valence electrons** correspond to the electrons in the outer shell of an atom. These are the electrons that are involved in chemical bonding. The number of valence electrons for all main group elements can be determined from the periodic chart.



<u>Example</u>: Determine the number of valence electrons that each of the following elements possesses:

Chlorine Aluminum Carbon Xenon

II. **Lewis symbols** show elements with their valence electrons. Lewis symbol examples for all main group elements are provided below.



<u>Example</u>: The Lewis symbol of an unknown element has 6 paired electrons and 1 unpaired electron; how many bonds is this atom likely to form in a covalent compound?

III. **The Octet Rule** states that for an atom to be stable it must have 8 electrons in its valence shell. Each bond and lone pair correspond to two electrons. However, the octet rule has a few notable exceptions.

- **a.** Atoms which have fewer than eight valence electrons:
  - **1.** Hydrogen only needs two electrons to obtain a full outer shell, and the element can also be stable with zero electrons.
  - **2.** Boron and beryllium can be stable with fewer than 8 electrons: 4 in the case of beryllium, and 6 in the case of boron.
- **b.** Atoms which have more than eight valence electrons:
  - **1.** Non-metals in Period 3 or higher can exceed the octet rule and maintain more than 8 electrons around their nuclei (ex. Cl, Br, S, P, Ar, Kr, Sn, Sb, Xe).

<u>Example:</u> Which of the following elements can have an expanded octet? *<u>Select all that apply</u>*.

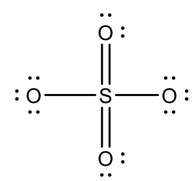
a) H. b) Be. c) Cl. d) O. e) F.

IV. **Formal charge** indicates the charges on the individual atoms as well as the overall charge present on the molecule. Use the equation below to calculate formal charge.

Formal Charge = (# valence electrons) - (# of nonbonding electrons) - 1/2(# of bonding electrons)

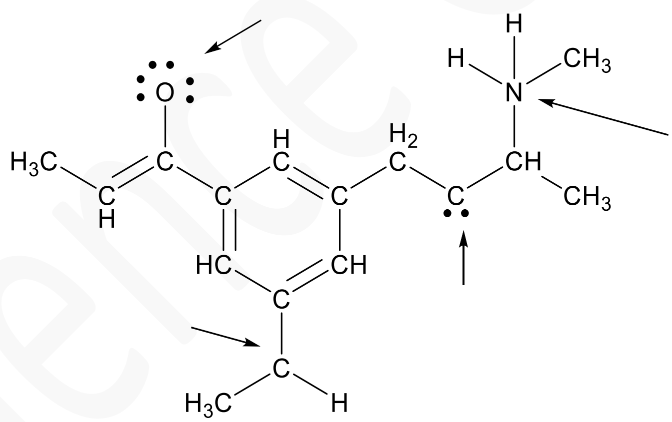
Formal Charge = (# valence electrons) - (# of dots) - (# of lines)

Example: Determine the formal charge on each atom in the compound below. Determine the net charge on the overall compound.



Net Charge: ______________________

Example: Determine the formal charge of each atom in the structure provided below.



V. **Lewis Structures** are structural representation of a molecule/atom/ion that shows all valence electrons and charges. There are several **steps we can take for drawing Lewis structures:**

a. **Determine the total number of valence electrons.** Using the molecular formula, sum the total number of valence electrons present in the molecule. For non-neutral compounds (ions), adjust the total count:

1. Anions (-): For each negative charge, add one electron.
2. Cations (+): For each positive charge, subtract one electron.

b. **Determine the central atom(s) of the compound**. This is typically the most electropositive atom, or the atom that can make the greatest number of bonds. In organic molecules, the central atom will always be carbon.

c. **Add valence electrons to the central atom.** Add the number of valence electrons which the central atom has been designated to have based on the periodic chart, to the central atom.

d. **Connect atoms via covalent bonds.** Make bonds connecting the central atom to other atoms such that each atom gains an octet. The number of valence electrons present in the connecting atoms can be determined from the periodic chart as well.
   1. **Carbon:** 4 bonds, no lone pairs
   2. **Nitrogen:** 3 bonds, 1 lone pair
   3. **Oxygen**: 2 bonds, 2 lone pairs
   4. **Halogens:** 1 bond, 3 lone pairs

e. **Insert multiple bonds.** Double or triple bonds sometimes may be made to satisfy the octet rule. Insert these one at a time, as needed.

f. **Assign formal charges.** Assign formal charge to each atom in the structure that has a formal charge different than zero.

g. **Choose the favored structure.** Some formulas may yield multiple valid structures. The most favored structure will be the one that minimizes formal charges overall.

Example: Draw the best Lewis structure for each of the following.

a) CH_4. b) H_2O. c) BH_3.

d) NH_3. e) ClO_4-. f) $SbCl_4$+

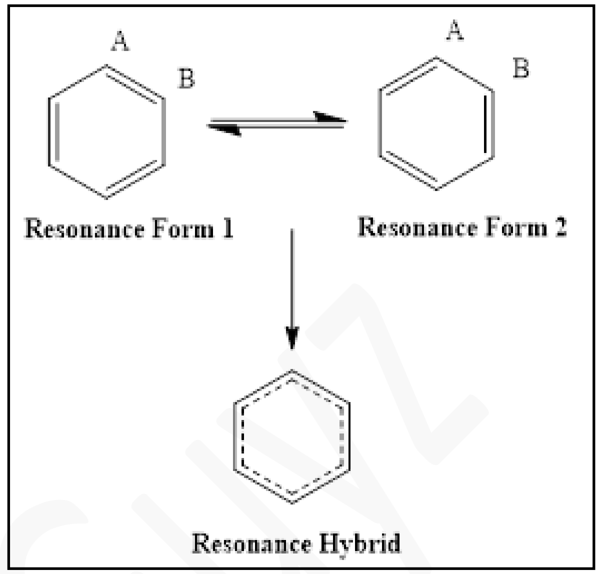
h) HNO_3. i) CO_2. g) SF_4.

Example: Determine the number of single bonds, double bonds, triple bonds, and lone pairs present in the Lewis structure for HCN.

Example: Which of the following Lewis structures would **violate** the octet rule? *Select all that apply*.

a) CCl_4. b) XeF_4. c) PO_4^{3-}. d) $BrCl_4$+. e) SF_4.

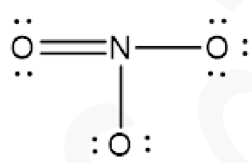
<u>Section</u>: More on Lewis Structures and Bonds



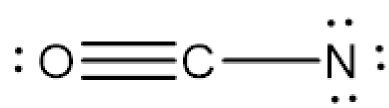
I. **Resonance structures** are structures that are identical in terms of molecular structure and energy, but different in the location of electrons.

   a. **Forming Resonance Structures:** The different resonance structures of a molecule can be drawn by (a) moving a pi bond (one of the bonds in a double or triple bond), or (b) by making a lone pair of electrons into a pi bond, making a pi-bond into a lone pair of electrons, or a combination of both.
   b. **Net Charge MUST remain Constant**: For two structures to be considered resonance forms, the net charge on the compound must remain constant between two resonance forms.
   c. **Movement of Sigma-Bonds**: NEVER move sigma-bonds (single bonds) when trying to create new resonance structures.
   d. **Resonance Hybrids**: For molecules that display resonance, we can draw the actual electron arrangement as a combination of all the possible resonance structures.
   e. **More Resonance / Greater Stability**: A greater number of resonance structures results in a more stable molecule from a reactivity standpoint. This is due to what is known as delocalization of the valence electrons.
      i. **Delocalized** electrons are electrons that *can* move.
      ii. **Localized** electrons are electrons that cannot move.

<u>Example</u>: Draw all the resonance structures of the following molecule.



<u>Example</u>: Resonance and resonance structures are essential concepts to understand in your journey through chemistry. One resonance form of cyanate is provided to the right.



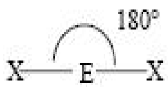
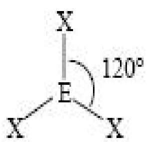
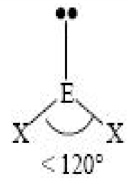
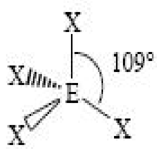
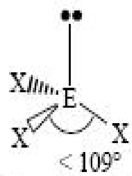
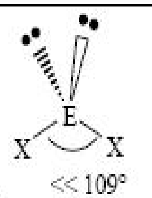
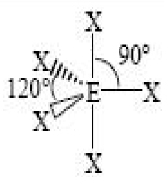
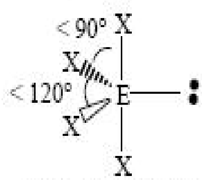
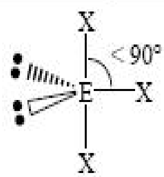
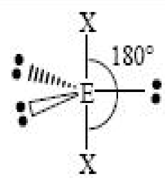
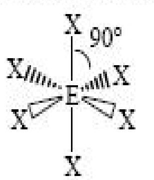
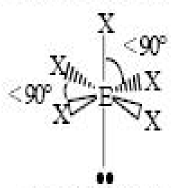
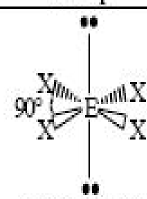
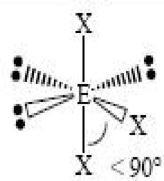
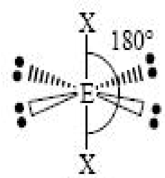
a) Draw the other possible resonance forms of cyanate.

b) Which resonance structure of cyanate is **most** likely to form? Explain your reasoning.

<u>Section</u>: Molecular Shapes and Electronic Geometry vs. Molecular Geometry

I. **Valence shell electron-pair repulsion (VSEPR) Theory**: A model which allows us to predict the 3-dimensional shape of covalent molecules and polyatomic ions. The model is based off the idea that bonds and lone electron pairs in the valence shell of an element repel each other and thus want to be as far away from each other as possible.
   - **a.** VSEPR Theory can be used to estimate bond angles and predict polarity.
   - **b.** By counting the regions of electron density (bonds *plus* electron pairs) surrounding the central atom, and then determining the number of lone pairs of electrons, the *electronic* and *molecular geometry* can be determined.
     - **i. Electronic Geometry** looks at ALL pairs of electrons surrounding the central atom of a compound. This includes electrons involved in bonding AND lone pairs of electrons.
     - **ii. Molecular Geometry** ONLY looks at pairs of electrons involved in bonding. It does not consider lone pairs of electrons.

Regions of Electron Density	Angles	Lone Pairs of e^-	Electronic Geometry	Molecular Geometry
2	**180°**	**0**	**Linear**	**Linear**
3	**120°**	**0**	**Trigonal Planar**	**Trigonal Planar**
		1		**Bent**
4	**109.5°**	**0**	**Tetrahedral**	**Tetrahedral**
		1		**Trigonal Pyramidal**
		2		**Bent**
5	**90° or 120°**	**0**	**Trigonal Bipyramidal**	**Trigonal Bipyramidal**
		1		**Seesaw**
		2		**T-Shaped**
		3		**Linear**
6	**90° or 180°**	**0**	**Octahedral**	**Octahedral**
		1		**Square Pyramidal**
		2		**Square Planar**
		3		**T-Shaped**
		4		**Linear**

VSEPR Geometries					
Steric No.	Basic Geometry 0 lone pair	1 lone pair	2 lone pairs	3 lone pairs	4 lone pairs
2	 Linear				
3	 Trigonal Planar	 Bent or Angular			
4	 Tetrahedral	 Trigonal Pyramid	 Bent or Angular		
5	 Trigonal Bipyramid	 Sawhorse or Seesaw	 T-shape	 Linear	
6	 Octahedral	 Square Pyramid	 Square Planar	 T-shape	 Linear

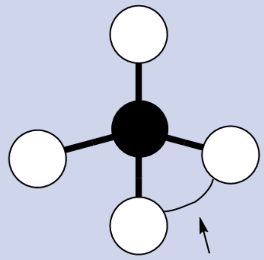
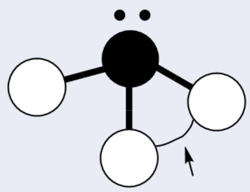
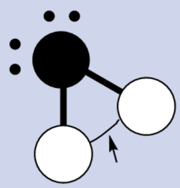
<u>Example</u>: Draw out the Lewis structures for each of the following molecules and determine their electronic and molecular geometries.

a) BeF_2. b) SO_2. c) NO_2- d) BH_3.

e) TeO_3. f) SCl_4. g) GeO_2. h) $XeCl_2$.

Science Guyz, LLC | 250 W Broad St STE 102 Athens, GA 30601 | Phone: 706-723-8436
Email: tutor@scienceguyz.com | Website: www.scienceguyz.com |

II. **Effect of Lone Pairs on Molecular Geometry:** When memorizing the bond angles of molecular structures, it is important to note the effect of lone pairs. Observe the graphic below:

Molecular Geometry	Diagram	Approximate Bond Angle
Tetrahedral		109.5°
Trigonal Pyramidal		107.5°
Bent		104.5°

a. The lone pairs push the constituent atoms towards each other, reducing the bond angle between the non-central atoms. This effect is stronger when there are two lone pairs. Different compounds will experience this effect to different extents.

   i. The take home message is that the more lone pairs that are present, the smaller the bond angles will be.

<u>Example</u>: Draw the Lewis structure for each of the molecules below and predict which would have the smallest bond angles and which would have the largest.

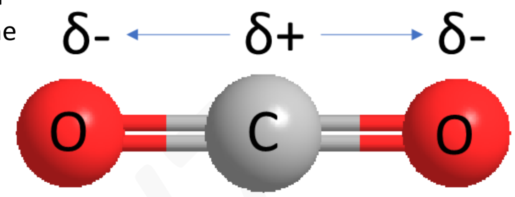
BF_3 BrF_3 NF_3

<u>Example</u>: Which of the following would have the **smallest** bond angles between nearest neighbors?
a) NF_3. b) BCl_3. c) SF_3^-. d) $XeBr_2$. e) CH_4.

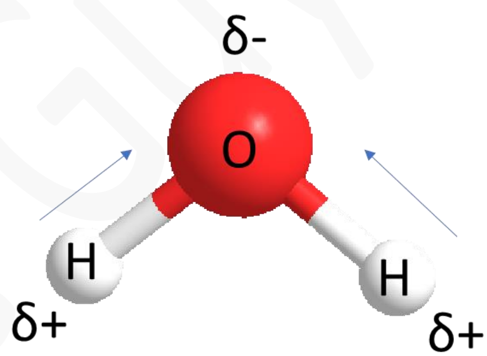
<u>Section</u>: Molecules and Polarity

I. A molecule can be described as polar if the molecule possesses a net dipole, or an uneven distribution of charge (electrons). If the individual bond dipoles within the molecule do not completely cancel each other out (see below), the molecule is polar.

- **a. Always check net dipole.**
- **b.** Please note that it is possible for a molecule to contain a polar covalent bond and not be polar overall, observe the example of carbon dioxide.



  - **i.** Carbon dioxide is a non-polar molecule. Both bonds within the molecule are polar, however the dipoles present within the molecule are of equal strength and in opposite directions. Because the dipoles cancel each other out, the molecule is non-polar. Note that if the two non-central atoms were not the same or if carbon dioxide were not symmetrical, it would be polar.
- **c.** Water is a polar molecule. Both bonds within the molecule are polar, have the same strength, and do not cancel each other out like we saw with carbon dioxide. This means that there is an overall net dipole within the molecule. This bent shape is caused by the two lone pairs sitting on top of the oxygen molecule.
- **d.** We will explore how you can use molecular shapes to predict polarity.



II. Using molecular geometry, it is possible to predict whether a molecule will be polar. Observe the chart below:

Molecular Geometry	**Polar?**	**Molecular geometry**	**Polar?**
Linear	Not always	Seesaw	**Always**
Bent	**Always**	T-Shaped	**Always**
Trigonal Planar	Not always	Octahedral	Not always
Tetrahedral	Not always	Square Pyramidal	**Always**
Trigonal pryamidal	**Always**	Square Planar	Not always
Trigonal Bipyramidal	Not always		

a. What does "not always" depend on?
   i. If same atoms surrounding central atom = **NON-POLAR**
   ii. If different atoms surrounding central atom = **POLAR**

<u>Example</u>: Draw the Lewis structures for the following molecules and determine whether they are polar or non-polar.

a) CH_2Br_2

b) H_2S

H 1 1.01																	He 2 4.00
Li 3 6.94	Be 4 9.01											B 5 10.81	C 6 12.01	N 7 14.01	O 8 16.00	F 9 19.00	Ne 10 20.18
Na 11 22.99	Mg 12 24.31											Al 13 26.98	Si 14 28.09	P 15 30.97	S 16 32.07	Cl 17 35.45	Ar 18 39.95
K 19 39.10	Ca 20 40.08	Sc 21 44.96	Ti 22 47.87	V 23 50.94	Cr 24 51.99	Mn 25 54.94	Fe 26 55.85	Co 27 58.93	Ni 28 58.69	Cu 29 63.55	Zn 30 65.38	Ga 31 69.72	Ge 32 72.63	As 33 74.92	Se 34 78.97	Br 35 79.90	Kr 36 83.80
Rb 37 85.47	Sr 38 87.62	Y 39 88.91	Zr 40 91.22	Nb 41 92.91	Mo 42 95.95	Tc 43 98.91	Ru 44 101.07	Rh 45 102.91	Pd 46 106.42	Ag 47 107.87	Cd 48 112.41	In 49 114.82	Sn 50 118.71	Sb 51 121.76	Te 52 127.6	I 53 126.90	Xe 54 131.29
Cs 55 132.91	Ba 56 137.33	57-71	Hf 72 178.49	Ta 73 180.95	W 74 183.84	Re 75 186.21	Os 76 190.23	Ir 77 192.22	Pt 78 195.09	Au 79 196.97	Hg 80 200.59	Tl 81 204.38	Pb 82 207.2	Bi 83 208.98	Po 84 [208.98]	At 85 209.99	Rn 86 222.02
Fr 87 223.02	Ra 88 226.03	89-103	Rf 104 [261]	Db 105 [262]	Sg 106 [266]	Bh 107 [264]	Hs 108 [269]	Mt 109 [278]	Ds 110 [281]	Rg 111 [280]	Cn 112 [285]	Nh 113 [286]	Fl 114 [289]	Mc 115 [289]	Lv 116 [293]	Ts 117 [294]	Og 118 [294]

La 57 138.91	Ce 58 140.12	Pr 59 140.91	Nd 60 144.24	Pm 61 144.91	Sm 62 150.36	Eu 63 151.96	Gd 64 157.25	Tb 65 158.93	Dy 66 162.50	Ho 67 164.93	Er 68 167.26	Tm 69 168.93	Yb 70 173.06	Lu 71 174.97
Ac 89 227.03	Th 90 232.04	Pa 91 231.04	U 92 238.03	Np 93 237.05	Pu 94 244.06	Am 95 243.06	Cm 96 247.07	Bk 97 247.07	Cf 98 251.08	Es 99 [254]	Fm 100 257.10	Md 101 258.1	No 102 259.10	Lr 103 [262]