CHEM 1212
General Chemistry I
Refresher

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Section: Scientific Notation & Units of Measurement

I. A number is written in **scientific notation** when in the form $a \times 10^{n}$ where $1 \leq |a| < 10$ and n is an integer. Scientific notation makes a large or small number more compact by writing them as a product of the power of 10.

- **a.** In putting a large number into scientific notation, move the decimal place to the **left** until $1 \leq |a| < 10$. With $a \times 10^{n}$, n will be a positive number equal to the number of times you moved the decimal to the left.
- **b.** In putting a small number into scientific notation, move the decimal place to the **right** until $1 \leq |a| < 10$. With $a \times 10^{n}$, n will be a negative number equal to the number of times you moved the decimal to the right.

Example: Convert the following numbers into or out of scientific notation.

a) 0.0000783 = ____________________ b) 985512 = ____________________

c) 9.31×10^{6} = ____________________ d) 7.55×10^{-3} = ____________________

II. **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 shown to the right.

Property	**Unit Used**
Mass	kilogram (kg)
Length	meter (m)
Time	second (s)
Temperature	kelvin (K)
Amount of Substance	mole (mol)
Electric Current	ampere (A)
Luminous Intensity	candela (cd)

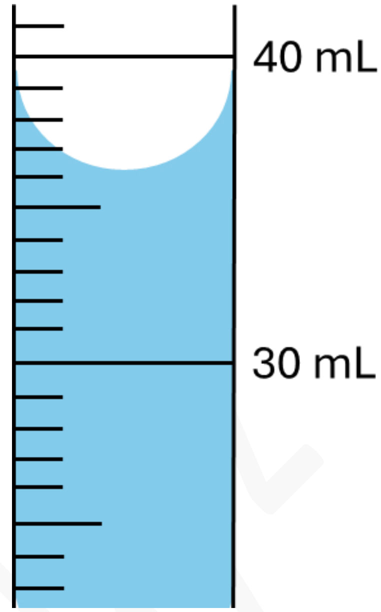
III. **Common SI Prefixes (Commit to Memory!):**

Prefix	Abbreviation	Meaning	Example
Exa-	E	10^{18}	1 exabyte = 1 x 10^{18} bytes
Peta-	P	10^{15}	1 petabyte = 1 x 10^{15} bytes
Tera-	T	10^{12} (trillion)	1 terahertz = 1 x 10^{12} hertz
Giga-	G	10^{9} (billion)	1 gigahertz = 1 x 10^{9} hertz
Mega-	M	10^{6} (million)	1 megaton = 1 x 10^{6} tons
Kilo-	k	10^{3} (thousand)	1 kilogram = 1 x 10^{3} grams
Hecto-	h	10^{2} (hundred)	1 hectopascal = 1 x 10^{2} pascals
Deka-	da	10^{1} (ten)	1 dekameter = 1 x 10^{1} meters
Base Unit	-	10^{0}	________________________
Deci-	d	10^{-1} (tenth)	1 decimeter = 1 x 10^{-1} meters
Centi-	c	10^{-2} (hundredth)	1 centimeter = 1 x 10^{-2} meters
Milli-	m	10^{-3} (thousandth)	1 millimeter = 1 x 10^{-3} meters
Micro-	µ	10^{-6} (millionth)	1 micrometer = 1 x 10^{-6} meters
Nano-	n	10^{-9} (billionth)	1 nanometer = 1 x 10^{-9} meters
Pico-	p	10^{-12} (trillionth)	1 picometer = 1 x 10^{-12} meters
Femto-	f	10^{-15}	1 femtometer = 1 x 10^{-15} meters
Atto-	a	10^{-18}	1 attometer = 1 x 10^{-18} meters

Example: Provide the correct abbreviations for the following units of measurement.

a) micrometer: _____ b) nanogram: _____ c) milliliter: _____ d) centimeter: _____

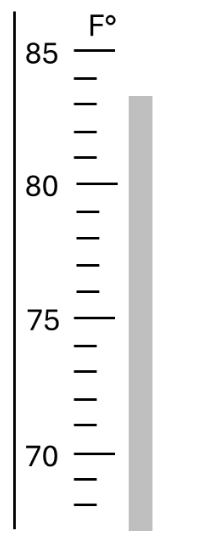
Section: **Significant Figures**



I. **Significant figures** involve the numbers in a measured quantity or value that are known to be correct and one digit that is an estimation. A common application of significant figures is with reading glassware in the lab.

- **a.** To the right is a graduated cylinder. Each mark on the graduated cylinder represents a 1 mL increase in volume.
- **b.** With the graduated cylinder to the right, it is known for sure that the true liquid measurement lies somewhere between 36 mL and 37 mL (always read from the meniscus).
- **c.** We are confident in the value of the ones and tens place, but the digit in the tenth place is an approximation.
- **d.** An appropriate guess for liquid in this graduated cylinder would be something along the lines of 36.1 mL or 36.2 mL. Either of these answers would be acceptable since the digit in the tenth place is an approximation.
- **e.** This graduated cylinder is one example of a measurement tool but understand that different measurement tools will provide a different number of significant figures.

Example: A thermometer filled with mercury is shown to the right. What is the temperature reading for this thermometer, considering significant figures?



II. **Rules for Significant Figures**:

- **a.** All non-zero numbers are significant.
- **b.** Zeroes between two other significant digits are significant, including numbers that have a decimal place (7.**0**51, 6**0**71**0**1).
- **c.** Zeroes following a non-zero number that is also to the left of a decimal are significant (72**000**.).
- **d.** In numbers containing a decimal, all zeros at the end of the number are significant (0.032**0**).
  1. Combining rules c and d shows us that the three 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 (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 infinite number of significant figures because they represent exact quantities and cannot be made more precise.
  - **i.** For instance, counting 12 pencils or stating that a molecule consists of 3 atoms are examples of exact values.
- **b.** Known conversions have an unlimited number of significant figures.
  - **i.** For instance, 1 foot = 12 inches is an example of an errorless conversion.

Example: Determine the number of significant figures contained within each of the following values.

a) 0.0900: _________ b) 6.230: _________ c) 0.0076: _________

d) 4.00028: _________ e) 4.335 x 10^{-22}: _________ f) 5600: _________

g) 0.12000 inches: _________ h) 1.20 feet: _________ i) 5600 pieces of paper: _________

j) 5462.0 kilometers: _________ k) 5462.00 kilometers: _________

IV. Rounding Following Mathematical Operations:

- **a.** Following any mathematical operation, you must determine how many significant figures will be in your final answer.
- **b.** Once you have determined the correct number of significant figures, you must round your answer to the correct number of significant figures by looking at the first non-significant digit in your answer.
  - **i.** If this number is 5 or greater, you will round your last significant digit up, otherwise, you will round the last significant digit down.

V. Significant Figures in Multiplication and Division Problems:

- **a.** When multiplying or dividing two or more numbers, your answer must contain the same number of significant figures as the value in your equation with the fewest number of significant figures.

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

(3.2005470) x (30.9) = ____________

VI. Significant Figures in Addition and Subtraction Problems:

- **a.** When adding or subtracting two or more numbers, 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. Significant Figures 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.

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}) = ____________

<u>Section</u>: Dimensional Analysis

I. In Chemistry and other STEM courses, you will often need to convert from one unit to another unit. 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 should reflect this.
3. Choose a value as a starting point for your conversion.
4. Set up a series of conversion steps that allow you to cancel all unwanted units and leave you only with the final desired units. Units can be canceled when the same unit can be found in the numerator of one step and the denominator of another step (or vice versa), *regardless of whether they are in consecutive steps.*
5. Once your units have been canceled to your final units, multiply the numerators and denominators, divide the products, and you are finished. Always make sure you report your final answer to the correct number of significant figures.

<u>Example</u>: How many picograms are in a milligram?

<u>Example</u>: How many angstroms (Å) are in 442 nm? <u>Note</u>: 1 Å = 1 x 10^{-10} m

<u>Example</u>: Convert 5.10 meters per second to micrometers per hour.

<u>Example</u>: Kailee is traveling 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 an average speed of 65 miles/hour, how many gallons of gas does she use?

<u>Example</u>: 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? <u>Note</u>: 1 lb = 453.59 g.

<u>Example</u>: Austin is moving up in the world and has decided it is time to install a pool at his house. Austin wants the pool to have the following dimensions: 37.2 ft long, 15.0 ft wide, and 7.92 ft deep. Once the pool has been completed, how much water will Austin need to fill the pool to the top (in cubic inches)?

<u>Example</u>: Convert the cubic inches obtained in the previous problem into milliliters. <u>Note</u>: 1 inch = 2.54 cm and 1 cm^3 = 1 mL.

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

I. **Solubility** is the extent to which a solute can be dissolved in a solvent under certain conditions. While many ionic compounds dissolve readily in water, not all are fully soluble.

- **a.** 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.
  - **i.** **If the list below 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 include containing silver, lead(II), and mercury(I).
4. Most fluoride salts are soluble. Notable exceptions include magnesium, calcium, strontium, barium, and lead(II).
5. Most sulfate salts are soluble, except for sulfate salts containing barium, lead(II), strontium, silver, and calcium.
6. Most hydroxide salts and oxides are insoluble except for the alkali metal hydroxides, barium hydroxide, and strontium hydroxide.
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 barium are soluble.

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

a) $MgCO_3$: ______________________
b) AgI: ______________________
c) $CaSO_4$: ______________________
d) KCH_3COO: ______________________
e) Ag_3PO_4: ______________________
f) $Mg(ClO_4)_2$: ______________________
g) $PbCl_2$: ______________________
h) MgF_2: ______________________
i) Ammonium carbonate: ______________________
j) Mercury(I) chloride: ______________________
k) Mercury(II) chloride: ______________________
l) Barium sulfide: ______________________
m) Magnesium phosphate: ______________________
n) Lead(IV) sulfate: ______________________
o) Chromium(III) hydroxide: ______________________
p) Aluminum oxalate: ______________________

II. **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) Sodium sulfate + barium chloride	________________________
b) Ammonium chloride + silver nitrate	________________________
c) Sodium phosphate + barium chloride	________________________
d) Lead(II) nitrate + sodium bromide	________________________
e) Potassium chloride + sodium nitrate	________________________

III. A **net ionic equation** shows what chemical reaction is taking place in a solution. Common ions (known as **spectator ions**) present on the reactant and product side of a reaction are omitted since they do not participate in any chemical change. The steps for writing a net ionic equation are provided below.

1. Write out the reaction using "aq" for aqueous (soluble) and "s" for solid (insoluble). These will be based on the solubility rules.
2. Balance the chemical equation.
3. Draw the ionic equation by splitting all soluble (aqueous) compounds into their individual ions. This will represent your **complete/total ionic equation**. Do not break apart compounds or molecules that are solids, liquids, or gases!
4. Cancel any spectator ions that appear on both sides of the reaction arrow.
5. 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: Sodium carbonate is mixed with iron(III) acetate. What is the net ionic equation for the reaction that occurs?

Example: Sodium bromide reacts with mercury(I) nitrate. What is the net ionic equation for the reaction that occurs?

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

<u>Section</u>: Acid-Base Reactions

I. There are a few different definitions for what an "acid" is and what a "base" is.

a. The Arrhenius Definition:

i. Acids contain hydrogen and increase H^+ (or H_3O^+) concentration in aqueous solutions.

ii. Bases contain hydroxide (OH^-) and increase the concentration of OH^- in aqueous solutions.

b. The Brønsted-Lowry Definition:

i. A **Brønsted-Lowry acid** is a chemical species that donates a proton (H^+) to some accepting chemical species.

ii. A **Brønsted-Lowry base** is a chemical species that accepts a proton (H^+) from some donating chemical species.

iii. Amphoteric species can behave as either an acid or a base depending on the situation. Water is an example of an amphoteric species because it can donate a proton or accept a proton, depending on the conditions.

<u>Example</u>: In each of the reactions below, identify the Brønsted-Lowry acid and the Brønsted-Lowry base on the reactant side of each chemical equation.

a) N_2H_4 (aq) + HSO_4- (aq) → $N_2H_5^+$ (aq) + SO_4^{2-} (aq)

Brønsted-Lowry acid: ___________

Brønsted-Lowry base: ___________

b) C_5H_5N (aq) + H_2O (l) → $C_5H_5NH^+$ (aq) + OH^- (aq)

Brønsted-Lowry acid: ___________

Brønsted-Lowry base: ___________

II. Acids and bases vary in strength. We can either define an acid or a base as strong or weak.

a. Strong acids and **strong bases** correspond to acids and bases that completely ionize in aqueous solution, making their reaction with water strongly product favored at equilibrium.

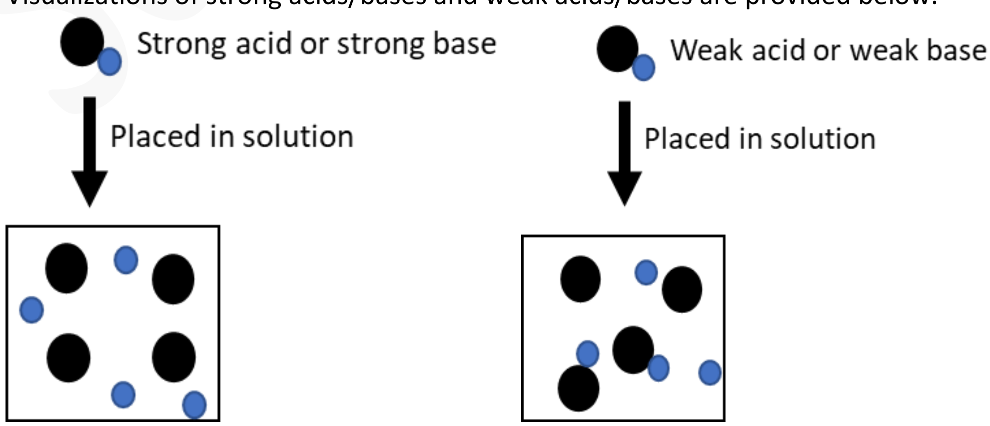
$$HCl\ (aq) + H_2O\ (l) \rightarrow H_3O^+\ (aq) + Cl^-\ (aq) \quad \text{(Complete Ionization)}$$

b. Weak acids and **weak bases** correspond to acids and bases that do not completely ionize in aqueous solution, making their reactions with water reactant favored at equilibrium.

i. Ammonia (NH_3) is a common weak base that you should be able to recognize.

$$HF\ (aq) + H_2O\ (l) \leftrightarrows H_3O^+\ (aq) + F^-\ (aq) \quad \text{(Partial Ionization)}$$

c. Visualizations of strong acids/bases and weak acids/bases are provided below.



d. The list of strong acids and strong bases is provided below.

i. **If this list differs from the one given by your instructor, use the list given by them!**

Strong Acids: HCl, HBr, HI, $HClO_4$, $HClO_3$, HNO_3, and H_2SO_4
Strong Bases: $LiOH$, $NaOH$, KOH, $Ca(OH)_2$, $Sr(OH)_2$, and $Ba(OH)_2$

Example: Which of the following are considered strong acids? *Select all that apply*.
a) Nitrous acid b) Chlorous acid c) Hydroiodic acid d) Hydrosulfuric acid e) Hydrochloric acid

III. **Acid-base reactions** (also called **neutralization reactions**) are reactions in which an acid reacts with a base in a double displacement reaction to form an ionic salt and water (most of the time).

a. We can examine net ionic equations of acid-base reactions.

i. Note that neutralization reactions involving a weak acid with a strong base are the one exception where we do not break up an aqueous substance!

Example: Which of the following compounds will **not** undergo a neutralization reaction with sodium hydroxide? *Select all that apply*.
a) Ammonia b) Sulfuric acid c) Potassium hydroxide d) Acetic acid e) Hydrofluoric acid

Example: Write out the net ionic equation for the reaction of hydrochloric acid with sodium hydroxide.

Example: Write out the net ionic equation for the reaction of acetic acid with potassium hydroxide.

Example: Write out the net ionic equation for the reaction of perchloric acid with iron(III) hydroxide.

Example: Write out the net ionic equation for the reaction of sulfuric acid with barium hydroxide.

<u>Section</u>: Introduction to Stoichiometry

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

<u>Example</u>: Solid aluminum carbonate reacts with aqueous phosphoric acid to produce solid aluminum phosphate, carbon dioxide gas, and liquid water. Use this information to answer the following questions.

a) Write a balanced chemical equation from the reaction described (including states of matter).

b) How many moles of phosphoric acid are required to react with 4.0 moles of aluminum carbonate?

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

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

<u>Example</u>: Aqueous ammonia reacts with oxygen gas to form nitrogen gas and liquid water. How many molecules of ammonia are required to react with excess oxygen to form 25.64 grams of nitrogen?

<u>Example</u>: You are provided with a 7.50 L sample of liquid benzene (C_6H_6) that reacts with excess oxygen according to the balanced equation below. Determine the amount of water vapor produced in this reaction (in grams). <u>Note</u>: The density of benzene is 0.877 g/cm^3.

$$2\ C_6H_6\ (l) + 15\ O_2\ (g) \rightarrow 12\ CO_2\ (g) + 6\ H_2O\ (g)$$

II. It is possible to determine the identity of an unknown atom or quantity of an atom in a compound by knowing how the atom reacts in a known reaction and employing the stoichiometry of the known balanced chemical equation and the **Law of Conservation of Mass**.

<u>Example</u>: A 3.560 gram sample of $CuSO_4 \cdot x\ H_2O$ yields 2.276 grams of $CuSO_4$ when heated, according to the balanced equation below. Determine the value of x.

$$CuSO_4 \cdot x\ H_2O\ (s) + heat \rightarrow CuSO_4\ (s) + X\ H_2O\ (g)$$

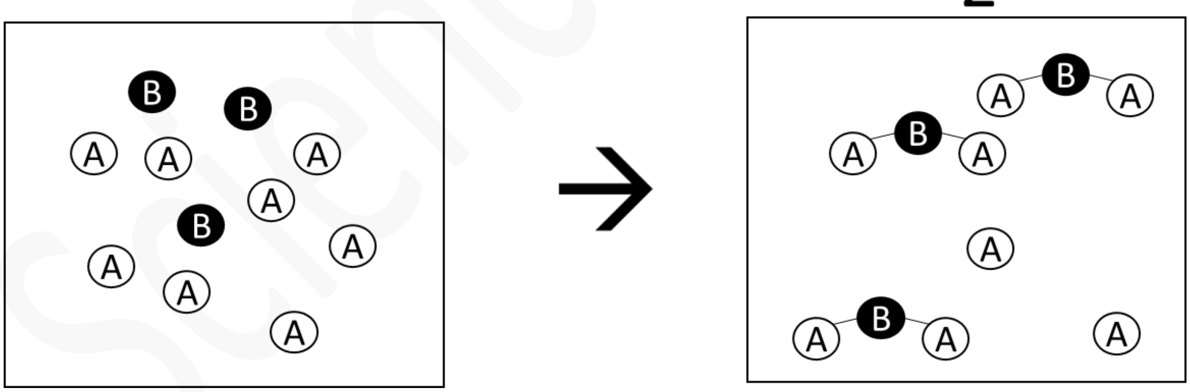
<u>Section</u>: More on Stoichiometry – Limiting and Excess Reactants & Percent Yield

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!

   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. Two different methods can be used to determine the amount of excess reactant remaining. We will explore both methods in an example problem.
      ii. 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 atoms of A remain, making it the excess reactant. Without more B, no further reaction can occur.

$$2\,A + B \rightarrow A_2B$$



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 from the balanced chemical equation.
4. Use the limiting reactant and the molar ratios to calculate what you are being asked.
   i. 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: We are exploring the arbitrary balanced chemical reaction below. If we initially have a reaction vessel composed of 36 molecules of A (g) and 36 molecules of B (g), how many **total** molecules will be present in the vessel after the reaction has gone to completion?

$$6\ A\ (g) + 4\ B\ (g) \rightarrow 2\ C\ (g) + 8\ D\ (g)$$

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.

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

a) What is the **maximum** amount (in grams) 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 (in grams) after the reaction has gone to completion? <u>Note</u>: Show this calculation in two different ways.

<u>Example</u>: In lab, we react 14.8 mL of sodium phosphate (density = 2.54 g/cm^3) and 36.4 mL of silver nitrate (density = 4.35 g/cm^3) to form solid silver phosphate. Using the **unbalanced** equation provided below, determine the **maximum** amount of the solid product (in grams) that could be produced in this reaction.

$$Na_3PO_4\ (aq) \quad + \quad AgNO_3\ (aq) \quad \rightarrow \quad Ag_3PO_4\ (s) \quad + \quad NaNO_3\ (aq)$$

III. In ideal conditions, the amount of product *actually* formed from a chemical reaction would always equal the predicted amount of product from the limiting reactant calculation.

- **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 the limiting reactant calculation.
  - **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.

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

<u>Example</u>: 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?

Section: Molarity and Dilutions

I. **Molarity (M)** is the most common form of concentration defined as the number of moles of a solute divided by the total volume of the solution.

- **a.** Molarity is a temperature dependent quantity.
- **b.** The formula for molarity is provided below.

$$\boxed{\boldsymbol{Molarity\ (M) = \frac{moles\ (solute)}{liters\ (solution)}}}$$

- **c.** The molarity of a solution can be determined in terms of the individual ions of an electrolyte, as well.
  - **i.** In the equation below with MX_2, M represents a cation with a 2+ charge and X represents an anion with a – 1 charge.

MX_2	→	M^{2+}	+	$2\ X^-$
[4 M]		[4 M]		[8 M]

Example: Determine the molarity (M) of a solution made by dissolving 160.5 g of glucose ($C_6H_{12}O_6$) to make a 650. mL solution.

Example: In a general chemistry lab, you are assigned to make a 0.300 M solution of barium nitrate. If you were to start with 120. mL of water, how many grams of barium nitrate is needed to make the solution? Note: Assume no volume change once the barium nitrate is added.

Example: In lab, you dissolve 100. g of magnesium chloride to make a 250. mL solution. Determine the molar concentration (M) of each individual ion in the solution.

Example: Which solution provided below would have the **greatest** concentration of Cl- ions present in the solution?

a) 1.0 M magnesium chloride.
b) 1.0 M sodium chloride.
c) 1.0 M aluminum chloride.
d) 1.0 M lead(IV) chloride.
e) Each has the same concentration of chloride ions.

II. We can combine concepts that we learned in stoichiometry with molarity.

Example: An **unbalanced** chemical reaction is provided below. An unknown amount of potassium is observed to react with an excess amount of water. This produces a 0.255 M solution of potassium hydroxide that has a final volume of 233 mL. Determine the unknown amount of potassium (in grams) that reacted with the water.

$$K\ (s) \quad + \quad H_2O\ (l) \quad \rightarrow \quad KOH\ (aq) \quad + \quad H_2\ (g)$$

Example: Determine the volume (in mL) of 0.550 M potassium oxalate that is needed to react completely with 45.0 mL of 0.250 M iron(III) nitrate using the balanced chemical equation below.

$$2\ Fe(NO_3)_3\ (aq) \quad + \quad 3\ K_2C_2O_4\ (aq) \quad \rightarrow \quad Fe_2(C_2O_4)_3\ (s) \quad + \quad 6\ KNO_3\ (aq)$$

III. **Dilutions** are a common lab technique used to go from a stock solution of higher concentration and add some volume of solvent to make a solution of lower concentration.

- **a.** The general equation for making a dilution is given below.

$$\boxed{M_1V_1 = M_2V_2\ or\ c_iV_i = c_fV_f}$$

  - **i.** Where M_1 or Ci is the initial concentration, V_1 or Vi is the initial volume, M_2 or Cf is the final concentration, and V_2 or Vf is the final volume.
- **b.** You can solve for concentrations and volumes before and after dilution, depending on what the problem gives you.
- **c.** Some dilutions may be multi-step.
- **d.** You must also conceptually understand what is going on when we are diluting a solution.
  - **i.** When we dilute a solution, the amount of *solute* does not change but the amount of *solvent* increases.

Example: What volume (in L) of 15.0 M nitric acid is needed to prepare 5.00 L of 1.50 M nitric acid?

Example: You evaporate water from 175 mL of a 0.500 M hydroiodic acid solution. Water stops evaporating once the volume decreases by 100. mL. What is the molarity (M) of the hydroiodic solution at this final volume?

<u>Section</u>: Hess's Law

I. When the enthalpy change of a reaction cannot be directly determined experimentally, we can estimate it through changes in enthalpy of other known reactions. **Hess's Law** states that if a reaction can be represented as the sum of two or more reactions, then the enthalpy of reaction ($\Delta H_r°$) for the overall process will be equivalent to the sum of the $\Delta H_r°$ values of those reactions.

II. **Applying Hess's Law**: The goal is to arrange the reactions such that you are left with *exactly the overall reaction* for which you are estimating the $\Delta H_r°$ for.

1. Determine the overall reaction for which you are trying to determine the $\Delta H_r°$.
2. Align the other individual reactions provided to you in the problem to the overall reaction given to you.
3. **Invert** reactions as needed to get the correct substances on the correct side, using the overall reaction as your guide. When reactions are inverted, you must also change the sign of the $\Delta H_r°$ associated with the reaction.
4. Some reactions will need to be **multiplied** by some factor to help you get the desired reaction. If you multiply a reaction by a factor, you must also multiply the $\Delta H_r°$ by the same factor.
5. **Cancel** all reactants and products that appear on opposite sides of different reactions. In some cases, you will cancel out partially. **Add** together reactants and products that appear on the same side of different reactions.
6. After all reaction manipulations, you should be left with the overall reaction for which you are trying to determine the $\Delta H_r°$ for. Add together all the $\Delta H_r°$ for each of the individual reactions, adjusting these values as needed based on inverting them or multiplying them by some factor. Once you have done this, you will have the $\Delta H_r°$ of the desired reaction.

<u>Example</u>: Given the standard enthalpy changes for the following two reactions, determine the standard enthalpy change for the overall reaction. <u>Note</u>: Report answer to four significant figures.

2 A (s) + B (g) →2 C (s) ΔH = -604.0 kJ/mol
2 D (l) + B (g) →2 E (s) ΔH = -180.0 kJ/mol

Overall Reaction: C (s) + D (l) →A (s) + E (s)

<u>Example</u>: The reaction of calcium hydroxide and hydrochloric acid is provided below with accompanying thermochemical data. Using this data, calculate ΔHr. <u>Note</u>: Report answer to four significant figures.

$Ca(OH)_2$ (aq) + 2 HCl (aq)→ $CaCl_2$ (aq) + 2 H2O (l)

<u>Useful Thermochemical Data</u>:

CaO (s) + 2 HCl (aq) → $CaCl_2$ (aq) + H2O (l) ΔH = -186.0 kJ/mol

CaO (s) + H_2O (l) → $Ca(OH)_2$ (s) ΔH = -65.1 kJ/mol

$Ca(OH)_2$ (s) → $Ca(OH)_2$ (aq) ΔH = - 12.6 kJ/mol

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

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

- **a.** Atoms that have fewer than eight valence electrons:
  - **1.** Hydrogen and helium only need 2 electrons to obtain a full outer shell.
  - **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 that have more than 8 valence electrons:
  - **1.** Non-metals in periods 3 and beyond can exceed the octet rule and maintain more than 8 electrons around their nuclei (ex. Cl, Br, S, P, Kr, Xe, etc).
- **c.** We can also encounter molecules that contain an odd number of electrons, and these are called **free radicals**.

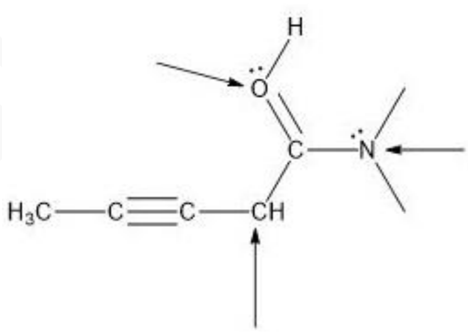
<u>Example</u>: Which of the atoms provided below can have an **expanded** octet? *<u>Select all that apply</u>*.
a) Hydrogen. b) Boron. c) Carbon. d) Xenon. e) Phosphorus.

II. **Formal charge** indicates the charges on the individual atoms as well as the overall charge present on the molecule.

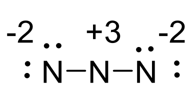
Formal Charge = (# valence electrons) - (# of nonbonding electrons) - 1/2(# of bonding electrons)
Formal Charge = (# valence electrons) - (# of dots) - (# of lines)

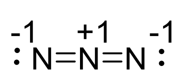
- **a.** Double and triple bonds can be used to fulfill the octet rule and/or to minimize formal charge where possible.
  - **i.** Formal charges should be as close to zero as possible and if we do have to have a formal charge, the most electronegative atom should possess the negative formal charge.
  - **ii.** An atom in period 3 and beyond can expand its octet to minimize its formal charge in the best Lewis structure.

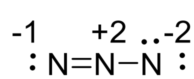
<u>Example</u>: Determine the missing formal charges on the indicated atoms in the molecule below.

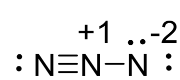


<u>Example</u>: Which Lewis structure is the **best** for the azide anion (N_3^-)?









○ ○ ○ ○

III. **Lewis Structures** are structural representations of a molecule that show 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.
- **b. Determine the central atom of the molecule**: This is typically the least electronegative atom, or the atom that can make the greatest number of bonds.
- **c. Add valence electrons to the central atom:** Add the number of valence electrons that the central atom has been designated to have based on the periodic table.
- **d. Connect atoms via covalent bonds:** Make bonds connecting the central atom to other atoms such that each atom gains an octet.
- **e. Assign formal charges:** Assign a formal charge to each atom in the structure. Some formulas may yield multiple valid structures. The most favored structure will be the one that minimizes formal charges overall.

<u>Example</u>: Draw the **best** Lewis structure for each of the following compounds and indicate the number of single, double, and triple bonds, as well as the number of lone pairs of electrons.

a) CF_4	b) XeF_4	c) SiH_2S	d) SO_2
Single Bonds: ______	Single Bonds: ______	Single Bonds: ______	Single Bonds: ______
Double Bonds: ______	Double Bonds: ______	Double Bonds: ______	Double Bonds: ______
Triple Bonds: ______	Triple Bonds: ______	Triple Bonds: ______	Triple Bonds: ______
Lone Pairs: ______	Lone Pairs: ______	Lone Pairs: ______	Lone Pairs: ______

e) HCN	f) ClF_4^-	g) PCl_3	h) NH_4^+
Single Bonds: ______	Single Bonds: ______	Single Bonds: ______	Single Bonds: ______
Double Bonds: ______	Double Bonds: ______	Double Bonds: ______	Double Bonds: ______
Triple Bonds: ______	Triple Bonds: ______	Triple Bonds: ______	Triple Bonds: ______
Lone Pairs: ______	Lone Pairs: ______	Lone Pairs: ______	Lone Pairs: ______

<u>Example</u>: Draw the **best** Lewis structure for each of the following carbon compounds and indicate the number of single, double, and triple bonds, as well as the number of lone pairs of electrons.

a) CH_2CH_2	b) NH_2CCCH_2I	c) $HOCH_2OCH_2OH$	d) $CH_3CHCHCO_2H$
Single Bonds: ______	Single Bonds: ______	Single Bonds: ______	Single Bonds: ______
Double Bonds: ______	Double Bonds: ______	Double Bonds: ______	Double Bonds: ______
Triple Bonds: ______	Triple Bonds: ______	Triple Bonds: ______	Triple Bonds: ______
Lone Pairs: ______	Lone Pairs: ______	Lone Pairs: ______	Lone Pairs: ______

IV. It is possible for a molecule to have a total number of valence electrons that is odd, meaning it will have unpaired electrons, known as **free radicals**.

- **a.** Compounds that exist as free radicals are very reactive.

Example: Two hypothetical elements are provided in a table with their corresponding number of valence electrons. Which of the hypothetical compounds formed from the elements would you expect to be free radicals? *Select all that apply*.
a) AB b) A_2B_2 c) AB_2 d) AB_3 e) A_3B

Element	Valence Electrons
A	4
B	5

Example: Which of the following compounds would be the **most** reactive?
a) CBr_4 b) F_2 c) NO_2^- d) ClO_2 e) HF

Example: Draw the **best** Lewis structure for nitrogen dioxide. In addition, assign formal charges and oxidation states to each atom in the molecule.

Example: Which of the following would violate the octet rule in their **best** Lewis structure? *Select all that apply*.
a) O_2- b) $BeCl_2$ c) OCl_2 d) IO_4- e) $SiCl_4$

V. **Resonance structures** occur when there is more than one possible valid Lewis structure for a compound. We indirectly explored this topic earlier in our discussions of formal charge.

- **a.** In resonance structures, the arrangement of atoms does not change, only delocalized electrons can move.
- **b.** When a molecule can resonate, the actual structure of the molecule is represented by an average of all the resonance structures, known as the **resonance hybrid**.
  - **i.** Please note that different resonance structures may not all contribute equally to the resonance hybrid; there is a difference between equivalent and non-equivalent resonance structures. The more favorable the resonance structure, the more it contributes to the resonance hybrid.
  - **ii.** Molecules that exhibit resonance **do not** rapidly switch between different resonance contributors; **this is a common misconception**.

Example: Draw all **equivalent** resonance structures of the phosphate ion. In addition, draw the resonance hybrid.

Example: Draw all possible resonance structures for the SNO- ion and identify the **biggest** resonance contributor. Hint: In all Lewis structures, nitrogen is the central atom.

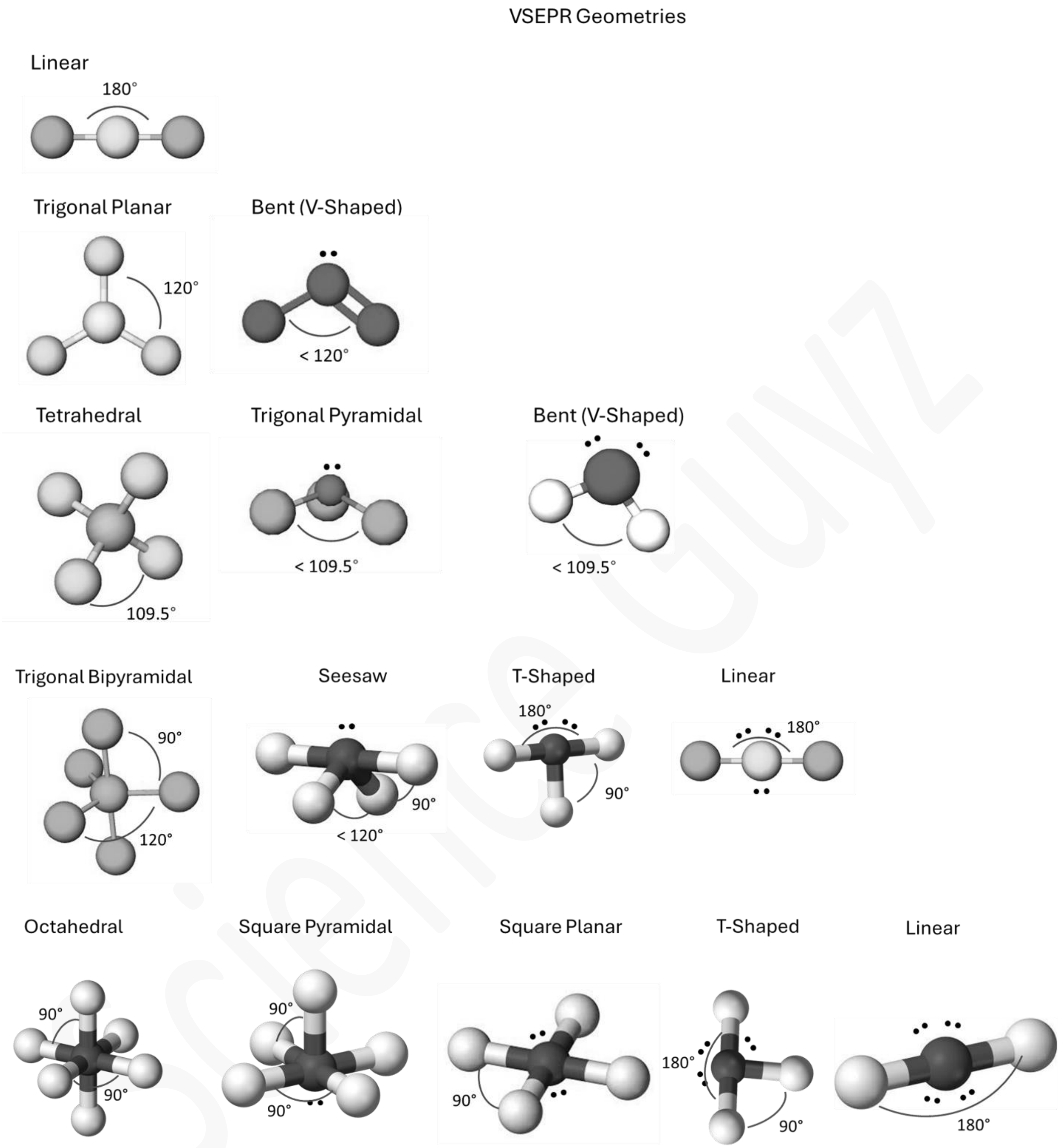
Section: VSEPR Theory

I. **Valence shell electron-pair repulsion (VSEPR) Theory** is a model which allows us to predict the 3-dimensional shape of covalent molecules and polyatomic ions. The model is based on 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. 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.

a. **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.

b. **Molecular Geometry** ONLY looks at the 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
	< 120°	1		Bent (or V-Shaped)
4	109.5°	0	Tetrahedral	Tetrahedral
	< 109.5°	1		Trigonal Pyramidal
	< 109.5°	2		Bent (or V-Shaped)
5	90° and 120°	0	Trigonal Bipyramidal	Trigonal Bipyramidal
	90° and < 120°	1		Seesaw
	90° and 180°	2		T-Shaped
	180°	3		Linear
6	90°	0	Octahedral	Octahedral
	90°	1		Square Pyramidal
	90°	2		Square Planar
	90° and 180°	3		T-Shaped
	180°	4		Linear



Example: Draw the **best** Lewis structures for the following molecules and determine their electronic geometries (EG) and molecular geometries (MG) and approximate bond angles (BA) observed.

a) CO_3^{2-} b) OF_2 c) XeF_2 d) BH_2^-

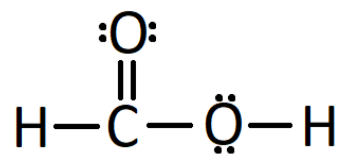
EG: _______________	EG: _______________	EG: _______________	EG: _______________
MG: _______________	MG: _______________	MG: _______________	MG: _______________
BA: _______________	BA: _______________	BA: _______________	BA: _______________

Example: For which of the following molecules do the electronic and molecular geometries **differ**? *Select all that apply*.

a) BCl_3 b) NCl_3 c) PO_4^{3-} d) SO_4^{2-} e) NO_2- f) SF_6 g) OBr_2

Example: Use the molecule to the right to indicate the following bond angles.

a) H-C=O: _________ b) O=C-O: _________ c) C-O-H: _________



II. 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, the molecule is polar.

- **a.** **Always check the net dipole to determine the polarity of a molecule.**
- **b.** **Beware:** A molecule may contain a polar covalent bond yet be a non-polar molecule overall.
- **c.** Using molecular geometry, it is possible to predict whether a molecule will be polar or non-polar. Look at 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 pyramidal	**Always**	Square Planar	Not always
Trigonal Bipyramidal	Not always		

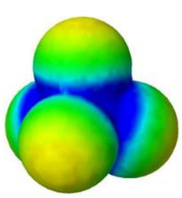
- **d.** What does "not always" depend on?
  - **i.** If the same atoms surrounding the central atom = **NON-POLAR**
  - **ii.** If different atoms surrounding the central atom = **POLAR**

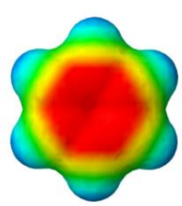
Example: Draw the following Lewis structures and determine if the molecules are polar or non-polar.

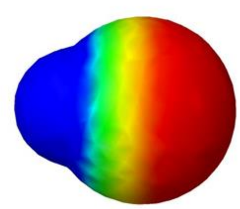
a) CCl_4 b) NO_3^- c) CH_2Br_2 d) IBr_5

III. We can use **electrostatic potential maps** to visualize electron distribution in a molecule and visualize a dipole. Utilizing the entirety of one of these maps, we can visualize whether a molecule will experience a net dipole or not.

Example: In each of the images provided below, determine if the molecule will experience a net dipole or not.







<u>Section</u>: Valence Bond Theory and Hybridization

I. **Valence Bond Theory** tells us that s, p, or d atomic orbitals can hybridize to form **hybrid orbitals**. Hybrid orbitals only exist in covalently bonded atoms and are both an average of the shapes and energy levels of the atomic orbitals that form them.

- **a.** **sp^3 hybridization** arises when a central atom has four regions of electron density.
- **b.** **sp^2 hybridization** arises when a central atom has three regions of electron density.
- **c.** **sp hybridization** when a central atom has two regions of electron density.
- **d.** To have more than four regions of electron density, we must now consider the d orbital. Molecules with five regions of electron density are known to be sp^3d hybridized and molecules with six regions of electron density are sp^3d^2 hybridized.

Regions of Electron Density	Hybridization	Lone Pairs of e^-	Electronic Geometry	Molecular Geometry
2	**sp**	**0**	**Linear**	**Linear**
3	**sp^2**	**0**	**Trigonal Planar**	**Trigonal Planar**
		1		**Bent (or V-Shaped)**
4	**sp^3**	**0**	**Tetrahedral**	**Tetrahedral**
		1		**Trigonal Pyramidal**
		2		**Bent (or V-Shaped)**
5	**sp^3d**	**0**	**Trigonal Bipyramidal**	**Trigonal Bipyramidal**
		1		**Seesaw**
		2		**T-Shaped**
		3		**Linear**
6	**sp^3d^2**	**0**	**Octahedral**	**Octahedral**
		1		**Square Pyramidal**
		2		**Square Planar**
		3		**T-Shaped**
		4		**Linear**

<u>Example</u>: In each molecule provided below, indicate the electronic geometry (EG), the molecular geometry (MG), the approximate bond angles (BA) observed, the hybridization of the central atom (H), and if the molecule is polar or non-polar (P or NP).

a) NH_3 b) SF_6 c) PBr_5 d) SH_2

EG: _______________ EG: _______________ EG: _______________ EG: _______________
MG: _______________ MG: _______________ MG: _______________ MG: _______________
BA: _______________ BA: _______________ BA: _______________ BA: _______________
H: _______________ H: _______________ H: _______________ H: _______________
P or NP: ___________ P or NP: ___________ P or NP: ___________ P or NP: ___________

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H 1.01																	2 He 4.00
3 Li 6.94	4 Be 9.01											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	23 V 50.94	24 Cr 51.99	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.63	33 As 74.92	34 Se 78.97	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.95	43 Tc 98.91	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	57-71	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.09	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po [209]	85 At [210]	86 Rn [222]
87 Fr [223]	88 Ra [226]	89-103	104 Rf [267]	105 Db [268]	106 Sg [269]	107 Bh [270]	108 Hs [269]	109 Mt [277]	110 Ds [281]	111 Rg [282]	112 Cn [285]	113 Nh [286]	114 Fl [290]	115 Mc [290]	116 Lv [293]	117 Ts [294]	118 Og [294]

57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm [145]	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.05	71 Lu 174.97
89 Ac [277]	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np [237]	94 Pu [244]	95 Am [243]	96 Cm [247]	97 Bk [247]	98 Cf [251]	99 Es [252]	100 Fm [257]	101 Md [258]	102 No [259]	103 Lr [262]