



ScienceGuyz

CHEM 1212

General Chemistry I Refresher

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Spring 2021

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Introduction:

This workshop is designed to help you review major topics discussed in your General Chemistry 1 course. The workshop will not cover every small detail in General Chemistry 1 but provide an overview of topics that tend to be further explored and built upon in General Chemistry 2.

Learning Objectives:

- Describe major trends associated with the periodic table, including effective nuclear charge, atomic size, ionization energy, electron affinity and electronegativity.
- Draw Lewis structures given a molecular formula.
- Describe the octet rule and identify the exceptions to it.
- Perform formal charge calculations and understand the significance behind them.
- Define what a resonance structure is.
- Identify and produce valid resonance structures if given a structure.
- Define sp , sp_2 , and sp_3 hybridization.
- Identify the hybridization of an atom in a molecule.
- Identify both the molecular and electronic geometry of a molecule.
- Describe how polarity arises and identify a bond as polar or non-polar.
- Explain how a molecule can have a polar bond yet overall be a non-polar molecule.
- Describe what a dipole moment is and how it impacts the polarity of a molecule.
- Rank bonds from most polar to least polar and explain how to come to this conclusion.
- Explain how a dipole moment can create partial charges on atoms within a molecule and be able to predict these partial charges.
- Explain how to use molecular geometry to predict polarity.
- Explain what molarity is and solve problems involving this quantity.
- Be able to balance chemical equations.
- Understand why we use dimensional analysis and the applications of it in various situations.
- Understand the concept of a mole and why it is used.
- Solve stoichiometry problems and be able to identify the limiting and excess reagents.
- Understand the formula for percent yield and apply it if given a problem.
- Memorize the diatomic elements.
- Understand the metric system, including the prefixes and abbreviations and be able to convert to various units within the system.
- Memorize the polyatomic ions.

Atomic Behavior and Periodic Table Trends:

To understand the chemical and physical behavior of the elements of the periodic table, it is important to become comfortable with trends associated with it.

Atomic Size (Main Group Elements):

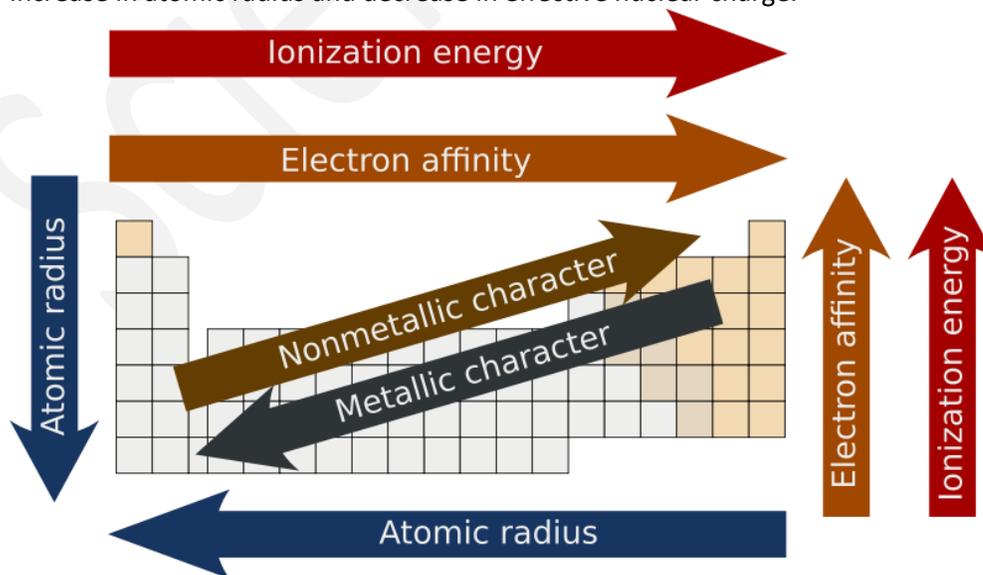
- Moving down in a given group of atoms, the atomic radius will increase. This is caused by electron shielding and the layering of electron orbitals taking up more space as they increase in diameter from the nucleus.
- Moving across a period from **left to right**, the atomic radius will **decrease**. This is caused by the increase in the effective nuclear charge as strength of positive charge increases in the nucleus via growing number of electrons, all while electron shielding does not increase.

Ionization Energy represents the energy required to remove an electron from an atom, specifically in the gas phase.

- Moving across a period from **left to right**, the ionization energy will **increase**. It is more difficult and energy intensive to remove an electron from a species farther right in a period.
- Moving down in a given group of atoms, the ionization energy will decrease. It is easier to remove electrons from larger species due to decreasing Z^* and increasing distance between valence electrons and their attracting nucleus.

Electron Affinity is the amount of energy released when an electron is added to a neutral atom in the gas phase.

- Moving across a period from **left to right**, the electron affinity of the atom will **increase**. This is caused by increasing effective nuclear charge.
- Moving down in a given group of atoms, the electron affinity will decrease. This is caused by the increase in atomic radius and decrease in effective nuclear charge.



Molecular Structure:

- Lewis Symbols:** Lewis symbols represent elements in which the valence electrons of the element are represented by dots which surround the elemental symbol. To determine the Lewis symbol of an atom, conduct the following:
 - Using the periodic chart, determine how many valence electrons the atom has.
 - Determine the charge of the atom in question.
 - Each negative charge equates to the addition of an electron in the valence shell of the neutral atom.
 - Each positive charge equates to the removal of an electron from the valence shell of the neutral atom.
 - Determine the number of electrons present in the atom or ion given add or remove electrons based on the charge of the atom.
 - Orient one valent electron around the atom, on each of the four sides of the atom. Double the electrons on a side of the atom only after you have added one electron to each of the four sides of the atom.
- The Octet Rule:** The octet rule is a rule that says that for an atom to reach a high degree of stability, the atom must have 8 electrons in its valence shell. Note: there are exceptions to the octet rule.
 - **Exceptions to the Octet Rule:**
 - Compounds where an atom has fewer than eight valence electrons:
 - **Hydrogen:** Hydrogen only needs two electrons to obtain an octet. Hydrogen can also have zero electrons and have an octet.
 - **Beryllium and Boron:** Beryllium can have a stable valence shell with 8 electrons or only 4. Similarly, Boron is stable with 8 or only 6 electrons.
 - Expanded Octets:** Atoms which have more than eight valent electrons
 - **Non-Metals in Period 3 or Higher:** Non-metals in period 3 or higher can have several oxidation states; this means that these elements can violate the octet rule and have more than 8 electrons around them.
- Drawing Lewis Electron Dot Structures**
 - Determine the Central Atom:** The central atom will generally be the atom which can make the greatest number of bonds (common examples include S, P, C, and N).
 - Determine Total # of Valent Electrons:** Sum the number of valence electrons from all atoms shown in the structure. For non-neutral compounds (ions), adjust the total count:
 - Anions (-):** For each negative charge, add an electron to your total valence electron count.
 - Cations (+):** For each positive formal charge, subtract an electron from your total valent electron count.
 - Add Valent Electrons to the Central Atom:** Add the number of valent electrons which the central atom has been designated to have based on the periodic chart, to the central atom.

- 4) **Connect Atoms Via Covalent Bonds:** Connect the chemical symbols of the other atoms in the compound to the central atom. Start by using single bonds wherever atoms are connected by bonds.
- 5) **Assign the rest of the electrons:** Place the rest of the electrons that each constituent atom should have around it using Lewis symbol rules.
- 6) **Create Multiple Bonds:** Often, it is necessary to make multiple bonds to satisfy the octet rule sufficiently for each atom. Apply the octet rule to each atom to make this determination. Remember to consider the exceptions to the rule.
- 7) **Choose the favored structure:** when presented with multiple valid structures, the most favored will be the structure with formal charges closest to zero. When this rule does not yield a clearly favored structure, choose the option that places a negative formal charge(s) on the most electronegative atom(s) present in the compound.

Formal Charge: Atoms participating in chemical bonding can exhibit a formal charge. A formal charge is a net positive or negative charge which is localized on an atom within a compound.

$$\text{Formal Charge} = e_{\text{valence}}^- - \frac{1}{2}e_{\text{bonding}}^- - e_{\text{non-bonding}}^-$$

e_{valence}^- : total number of valence electrons associated with the *neutral* atom

e_{bonding}^- : total electrons associated with bonds that involves the atom

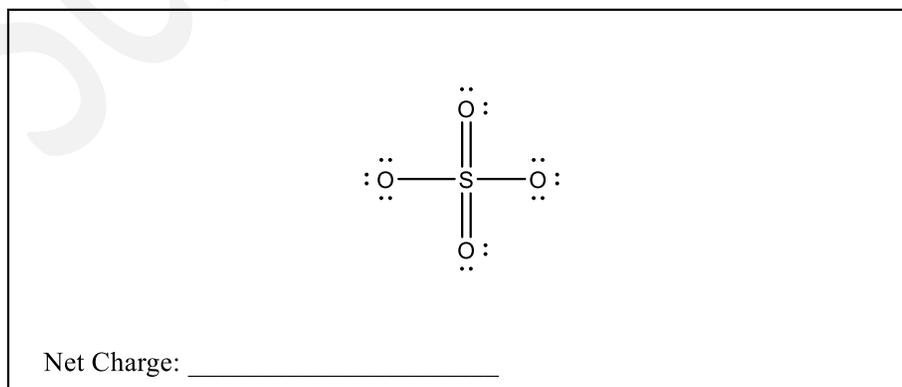
$e_{\text{non-bonding}}^-$: total valence electrons associated with the atom but not with bonds

Note: The sum of the formal charges on individual atoms in a molecule or ion must always equal the total net charge on the molecule or ion.

Hint: An easier way to remember the formal charge formula is shown below:

Formal charge = valence electrons – dots – sticks

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



Example: Give the Lewis structure for the following compounds:



Resonance Structures:

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

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.

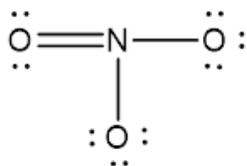
Net Charge MUST remain Constant: In order for two structures to be considered resonance forms, the net charge on the compound must remain constant between two resonance forms.

Movement of Sigma-Bonds: NEVER move sigma-bonds (single bonds) when trying to create new resonance structures.

Resonance Hybrids: for molecules that display resonance, we can draw the actual electron arrangement as a combination of all the possible resonance structures as below.

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 valent electrons.

Example: Draw all the resonance structures of the following molecule. Once all resonance forms have been drawn, draw the resonance hybrid.



Molecular Shapes and Electronic Geometry vs. Molecular Geometry:

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. VSEPR Theory can be used to estimate bond angles and predict polarity. 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.

- **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.
- **Molecular Geometry** ONLY looks at pairs of electrons involved in bonding. It does not consider lone pairs of electrons.

Regions of Electron Density (Steric Number)	Angles	Lone Pairs of e ⁻	Electronic Geometry	Molecular Geometry
2	180°	0	Linear	Linear
		1		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

Example: Draw out the Lewis structures for SnCl_4 and XeF_4 and determine their electronic geometries as well as their molecular geometries.

Hybridization:

Hybridization is the mixing of atomic orbitals from different subshells to form hybrid orbitals that are, in a sense, in a new, hybrid subshell. For instance, in sp hybridization, an s -orbital and a single p -orbital combine to form a hybrid sp subshell with two sp hybrid orbitals.

Solving for Hybridization:

How many regions of electron density does the central atom in my Lewis structure have? (Regions of Electron Density: Every lone pair of electrons; single bond, double bond or triple bond corresponds to a region of electron density.)

Determine the hybridization based on number of regions of electron density. For example, sp^2 hybridization has one s orbital and two p orbitals, so its three total hybridized orbitals correspond to a combined three regions of electron density

Example: For each of the following, draw the Lewis structure, and then indicate the electronic and molecular geometry and the hybridization of the central atom.

BI_3

Hybridization: _____

Electronic Geometry: _____

Molecular Geometry: _____

CO_2

Hybridization: _____

Electronic Geometry: _____

Molecular Geometry: _____

Regions of Electron Density	Hybridization
2	sp
3	sp^2
4	sp^3
5	sp^3d
6	sp^3d^2

Connecting Everything Together:

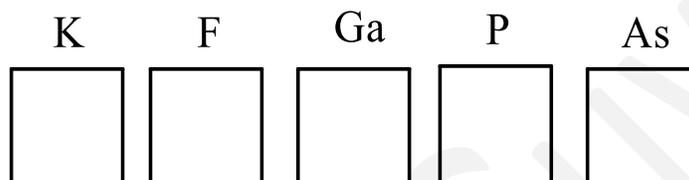
Regions of Electron Density	Hybridization	Lone Pairs of e ⁻	Electronic Geometry	Molecular Geometry
2	sp	0	Linear	Linear
3	sp ²	0	Trigonal Planar	Trigonal Planar
		1		Bent
4	sp ³	0	Tetrahedral	Tetrahedral
		1		Trigonal Pyramidal
		2		Bent
5	sp ³ d	0	Trigonal Bipyramidal	Trigonal Bipyramidal
		1		Seesaw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

Covalent Compounds and Electronegativity:

Electronegativity: Electronegativity is defined as the ability of an atom to attract electrons to itself and is given by the variable, X .

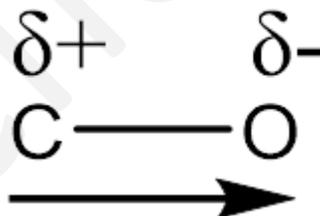
- Electronegativity generally increases from left to right across the period and increases from the bottom to the top of a group.
- Fluorine is the most electronegative element and cesium is the least electronegative element. Atomic size and electronegativity are inversely related.

Example: Rank these atoms in order of increasing electronegativity. 1=least electronegative



Polar Covalent Bonds: Polar covalent bonds arise when two atoms with large differences in electronegativities are bound together. When two atoms having large differences in electronegativities are bound together, the atoms will share electrons within the bond unevenly.

- **Dipole Moments:** When two atoms share electrons within a covalent bond unevenly the bond is said to have a dipole moment and is a **polar covalent bond**.
 - Partial Negative Charges:** In any polar covalent bond the atom which is more electronegative will bear a partially negative charge (δ^-)
 - Partial Positive Charges:** In any polar covalent bond the atom which is more less electronegative will bear a partially positive charge (δ^+)



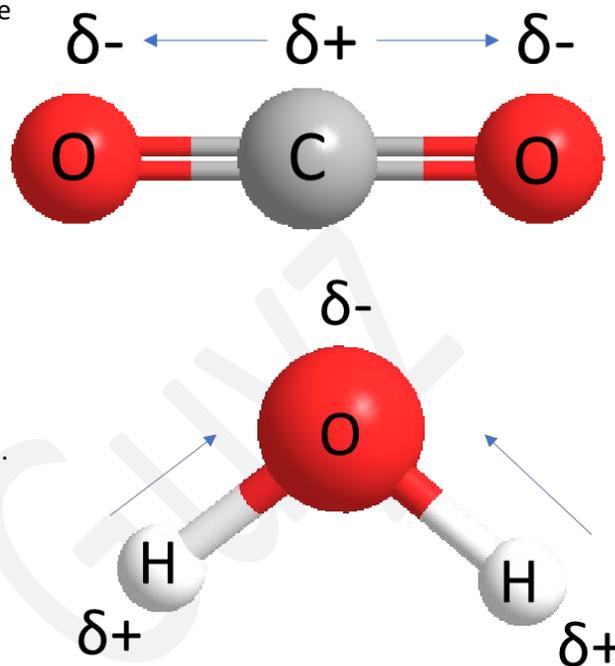
Example: Rank the bonds from least to most polar.



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. **Always check net dipole.**

Please note that it is possible for a molecule to contain a polar covalent bond, observe the example of CO_2 below:

- CO_2 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 CO_2 were not symmetrical, it would be polar.
- H_2O 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 CO_2 . 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. In the next section, we will explore how you can use molecular shapes to predict polarity.



Using Molecular Geometry to Predict Polarity:

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 pyramidal	Always	Square Planar	Not always
Trigonal Bipyramidal	Not always		

What does “not always” depend on?

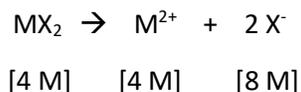
- If same atoms surrounding central atom = NON-POLAR
- If different atoms surrounding central atom = POLAR

Molarity:

Molarity (M) is a form of concentration defined as the quantity of moles of a solute (the thing being dissolved) divided by the total volume of the solution that it is dissolved in. Molarity is a temperature dependent quantity.

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

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



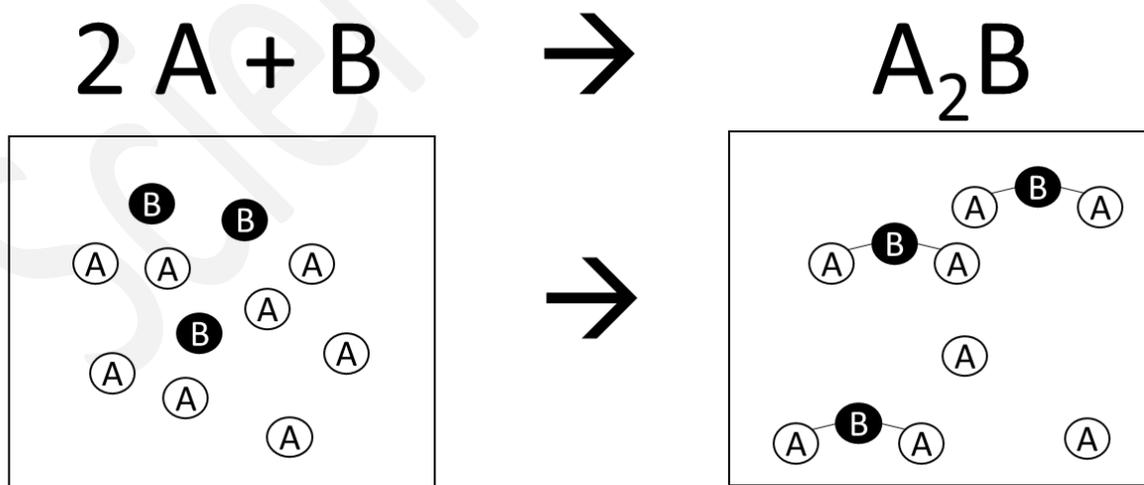
Example: Determine the molarity of a solution made by dissolving 300 g of magnesium acetate ($\text{Mg}(\text{CH}_3\text{COO})_2$) in enough water to make 750 mL of solution. In addition, determine the concentrations of the individual ions in solution.

Example: In a general chemistry lab, you are assigned to make a 0.300 M solution of barium nitrate ($\text{Ba}(\text{NO}_3)_2$) If you were to start off with 120 mL of water, how many grams of barium nitrate are needed to make the solution?

Stoichiometry:

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.

- Limiting Reactants:** In chemical reactions, there is always a limiting reactant. This is the reactant that will run out before the other reactants, and therefore *limit* how much of the reaction can take place. It is possible to have multiple limiting reactants.
 - 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.
 - To determine which reactant is limiting, determine which reactant will produce the smallest molar 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 reagent. However, you must take care to use the same product for each reactant.
 - It is important to note that only *molar quantities* can be compared when determining the limiting reactant. The amounts in grams should NOT be used when determining the limiting reactant!
- 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.
 - We can calculate the amount of an excess reactant remaining using the following general formula: $\text{Excess Reactant Remaining} = (\text{Starting Amount of Excess Reactant}) - (\text{Excess Reactant Consumed})$
 - Use the limiting reactant to determine how much of the excess reactant(s) is consumed.

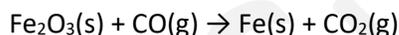


In the diagram above, two atoms of A react with one atom of B to form one molecule of A₂B. All the B is used up first, making it the limiting reactant and some molecules of A remain, making it the excess reactant. Without more B, no further reaction can occur.

General Steps for solving Stoichiometry Problems:

1. Write the balanced equation. **Always double check your equation to make sure it is balanced.**
2. Determine which reactant is limiting.
3. Convert into moles all the quantities given to you in the problem, in order to use the molar ratios of the balanced equation.
4. Use the limiting reagent 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: In the reaction provided below, 3.463 grams of iron (III) oxide is combined with 2.097 grams of carbon monoxide to produce the products shown. How many grams of CO₂ will be produced and what mass of the excess reagent will remain after the reaction has gone to completion?



Example: Atmospheric nitrogen and atmospheric hydrogen are placed in a reaction vessel and form ammonia (NH₃). Determine how much atmospheric nitrogen (in grams) is required to react completely with 19.3 g of atmospheric hydrogen and how much ammonia can be produced if all the atmospheric hydrogen were to react.

Percent Yield:

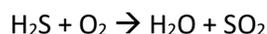
In perfect conditions, the amount of product formed from a chemical reaction would always equal the maximum amount of product that should form based on the limiting reactant based on the Law of Conservation of Mass. In practice this is not the case. The amount of a product formed during an *actual* chemical reaction is usually smaller than the amount predicted by stoichiometry.

- **Theoretical yield** is always determined by the stoichiometry of a balanced chemical reaction.
- **Actual yield** is ONLY determined experimentally and may be provided to you in a problem.
- **Percent yield** describes the efficiency of a reaction, and is given by the following equation:

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

Example: Calcium hydroxide reacts with excess hydrobromic acid to form calcium fluoride and water. In a general chemistry lab, a student reacts 130.5 g of the calcium hydroxide with excess hydrobromic acid. After the reaction has completed, the student collects 50.6 g of calcium fluoride. Determine the percent yield for the reaction.

Example: Reactions involving sulfur are common. An example of a reaction involving sulfur is shown below:



In your chemistry lab, you perform an experiment involving this reaction and obtain a percent yield of 71.2%. How many grams of H₂S is needed to obtain 63.51 g of SO₂?

Miscellaneous Charts:

Diatomic Elements: Elements which exist in nature as a molecule containing two of the same element. You **MUST** memorize each diatomic element.

- **Diatomic Elements:** Hydrogen, Nitrogen, Oxygen, Fluorine, Chlorine, Bromine, and Iodine. (Have No Fear Of Ice Cold Beer)

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

Names and Formulas of Common Polyatomic Ions			
Ammonium	NH_4^+	Borate	BO_3^{3-}
Hydroxide	OH^-	Nitrite	NO_2^-
Cyanide	CN^-	Nitrate	NO_3^-
Acetate	CH_3CO_2^-	Chlorite	ClO_2^-
Sulfite	SO_3^{2-}	Chlorate	ClO_3^-
Sulfate	SO_4^{2-}	Carbonate	CO_3^{2-}
Hydrogen sulfate	HSO_4^-	Hydrogen carbonate	HCO_3^-
Thiosulfate	$\text{S}_2\text{O}_3^{2-}$	Oxalate	$\text{C}_2\text{O}_4^{2-}$
Phosphate	PO_4^{3-}	Dichromate	$\text{Cr}_2\text{O}_7^{2-}$
Hydrogen phosphate	HPO_4^{2-}	Chromate	CrO_4^{2-}
Dihydrogen phosphate	H_2PO_4^-	Permanganate	MnO_4^-
Thiocyanate	SCN^-	Iodate	IO_3^-
Cyanate	OCN^-	Bromate	BrO_3^-

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