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Stoichiometry Worksheet 1 (Mole-Mole Conversions)

1. The combustion of the organic fuel, decane, is outlined in the chemical equation below. You must balance the equation in order to answer the subsequent questions a-c.



a. How many moles of CO₂ are produced if 5.0 moles of C₁₀H₂₂ react with an excess of O₂?

b. How many moles of O₂ react with 0.75 moles of C₁₀H₂₂?

c. How many moles of O₂ would be required to produce 4.0 moles of H₂O?

2. Use the following equation to solve the problems below:



a. If 6.0 moles of SO₂ react, how many moles of S are produced?

b. S are produced?

c. Al₂O₃ are produced?

3. Given the following reaction:



How many moles of nitrogen gas are required to produce 4.00 moles of ammonia?

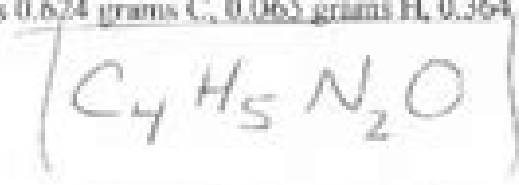
5. A 2.500 gram compound was found that contained only phosphorous and oxygen. The sample contained 1.0918 grams of phosphorous and has a molecular mass of 284 g/mol. Calculate the empirical and molecular formula.



6. A sample of ethylene glycol, antifreeze coolant, was found to contain 99.7 mg of C, 279 mL of H₂ gas at STP and 93 mL of O₂ gas at STP. If the molecular mass is 62 g/mol determine the empirical and molecular formula of ethylene glycol.



7. Caffeine, which occurs in coffee, tea, and kola, is a stimulant for the central nervous system. A 1.261 gram sample of pure caffeine contains 0.624 grams C, 0.065 grams H, 0.364 grams N, and the rest is O. What is the empirical formula of caffeine?



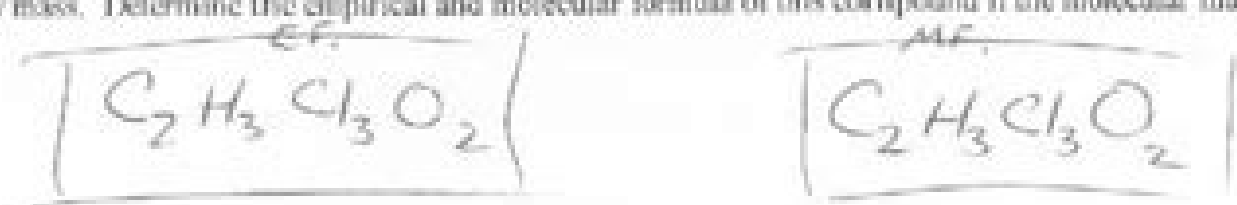
8. Butane gas contains only C and H. If the compound is 82.8% C, what is the empirical formula of the compound? If its molecular mass is 58 g/mol what is the molecular formula?



9. A compound contains 1.4211 grams of C, 2.950 L of H₂ gas at STP, and 0.4211 grams of S. What is the empirical formula of this compound?



10. Chloral hydrate is a drug that is used as a sedative and a hypnotic. It is used in detective stories to make a Mickey Finn. A Mickey is slipped into an unsuspecting victim's drink to knock them out. The compound contains 14.5% C, 1.81% H, 64.4% Cl, and 19.3% O by mass. Determine the empirical and molecular formula of this compound if the molecular mass is 166 g/mol.



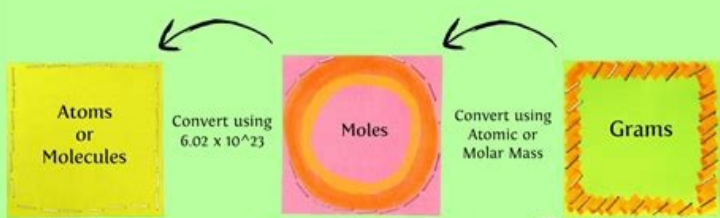
11. The molecular mass of petroleum is 114 g/mol. If a sample of fuel contains 21.05 grams of carbon and 44.24 dm³ of H₂ gas at STP, determine its molecular and empirical formula.



12. An oxide of nitrogen contains 100.0 L of N₂ gas and 50.0 L of O₂ gas at 25°C and 2.3 atm pressure. What is the empirical formula of the compound?

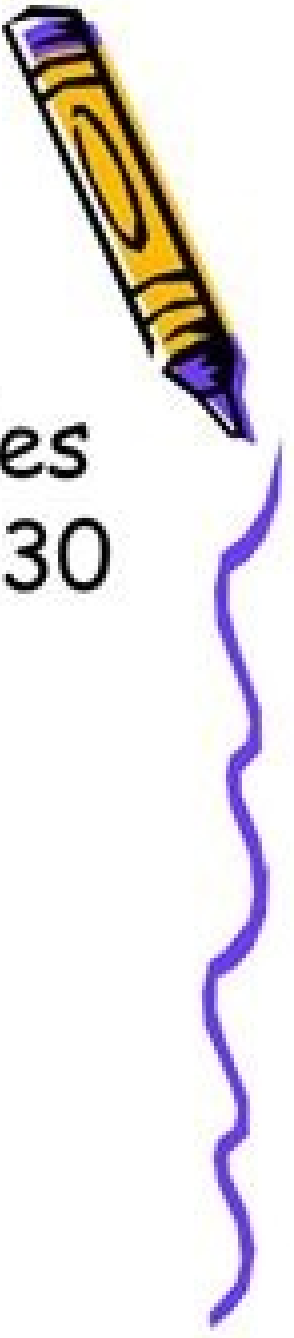
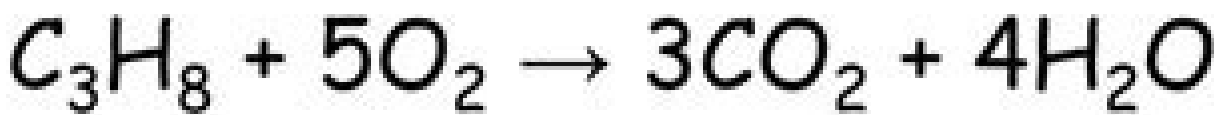


Stoichiometry Cheat Sheet



Practice

- 2) Calculate the number of moles of O₂ required to react with 4.30 mol propane:



Stoichiometry Calculation Practice Worksheet

1. Calculate the number of moles of NaOH that are needed to react with 500.0 g of H₂SO₄ according to the following equation:



2. Calculate the mass of NH₃ that can be produced from the reaction of 125 g of NCl₃ according to the following equation:



3. What mass of CO₂ can be produced from the reaction of 25.0 g of C₃H₈ with 75.0 g of O₂ according to the following equation:



One experimental method for determining the composition of organic compounds is combustion analysis, in which a weighed sample of the compound is burned in excess oxygen. How is a mole like a dozen? Using this integer factor, multiply all the subscripts (including any implied 1's) in the empirical formula to obtain the molecular formula. The atomic weight of oxygen is 16.00 u. [m.w. CO₂ = 44.01 u; m.w. H₂O = 18.02 u] For the burning (combustion) of propane gas the balanced equation is $C_3H_8(g) + 5 O_2(g) \rightarrow 3 CO_2(g) + 4 H_2O(l)$. When we first encountered reaction equations, we thought of this in terms of ratios among reactant and product species; e.g., "For every molecule of $C_3H_8(g)$, five molecules of $O_2(g)$ are required to produce three molecules of $CO_2(g)$ and four molecules of $H_2O(l)$." The relationships between individual reactant and product species and between moles of those species is multiplication by the constant Avogadro's number. Therefore, the ratios among moles of reactants and products are the same as between individual reactant and product species; e.g., "For every mole of $C_3H_8(g)$, five moles of $O_2(g)$ are required to produce three moles of $CO_2(g)$ and four moles of $H_2O(l)$." Thus, we can routinely interpret balanced chemical equations in terms of mole relationships. How do we know which of two or more reactants is limiting? For example, $1.25 \cdot 2.75 = 1\frac{1}{4} \cdot 2\frac{3}{4} = 5 \cdot 1\frac{1}{4}$. Write the empirical formula, using the same whole-number ratios between atoms of each element as the ratios among moles of elements. If the compound is found to have molecular weight of 88.12 u, what is the molecular formula? Find the lowest whole number ratios among the moles of elements. It follows that if we had a mole of atoms of some other element, that sample would weigh its atomic weight in grams. If you get stuck, try asking another group for help. In such cases, one reactant may be present in short supply, while other reactants may be present in abundance. Name: _____ Section: _____

Student ID#: _____ Work in groups on these problems. What is the mass of a mole of O₂(g)? The mole is sometimes described as "the chemist's dozen". Convert grams of each element into moles of each element, using atomic weights. It is important to realize that the limiting reagent is present in shortest supply on the basis of the stoichiometry of the balanced chemical equation in moles; i.e., the mole ratios implied by the balanced equation. A compound is found to contain 54.52% C, 9.17% H, and 36.31% O. A 2.554-g sample of a certain hydrocarbon is burned in excess oxygen, producing 8.635 g CO₂(g) and 1.768 g H₂O(l). The elemental composition of a compound can be determined experimentally by a variety of techniques. These masses can then be converted into moles, from which the empirical formula can be obtained. Assuming complete reaction, the reactant in shortest supply will be completely consumed, but some amounts of the other reagents will be left over after the reaction is finished. For a compound described by an empirical formula (e.g., ionic compound, network solid, empirical formula unit of a molecular compound), the molar mass in grams is numerically equal to the formula weight in atomic mass units. The strategy for converting analytical data into an empirical formula generally uses the following steps: Convert weight percentages into grams of each element. We then use the number of moles of the limiting reagent (not its calculated number of "sets") as the basis for all our further calculations, such as theoretical yield or amount of non-limiting reagent used. The masses of carbon and hydrogen in the original sample can be calculated from the weights of CO₂ and H₂O. If you have data for the percent composition of a compound, element by element, do you need to know the size of the sample in order to figure out the empirical formula? Define what is meant by the terms limiting reagent and excess reagent. Often it is helpful to assume a sample size of exactly 100 grams; then the given percentages are numerically equal to the number of grams of each element. Using mole and mass relationships, we can calculate the mass of product that should be produced from a given amount of reactant when it is completely consumed in the reaction. If the molecular weight is known, divide the formula weight for the empirical formula into the molecular weight to determine the number of formula units in the molecular formula. In any case, the theoretical yield of product always will be limited by the stoichiometric relationship between the limiting reagent and products. In doing this, we would use six bodies, and we would have 15 - 6 = 9 bodies left over. How much of the excess reagent will be left over, if compete reaction takes place? These ratios are the same as the ratios between the numbers of individual atoms of each elements in the empirical formula. In the complete combustion of propane, how many moles of $H_2O(l)$ are produced per mole of $O_2(g)$? In short, all calculations are based on the moles of the limiting reagent and the stoichiometric relationships implied by the balanced chemical equation. In all cases all the carbon in the compound is converted to CO₂, and all the hydrogen is converted to H₂O, which can be separated from each other and weighed. What is the theoretical yield of $Ca_3(PO_4)_2(s)$ by the reaction $Ca(OH)_2(s) + 2 H_3PO_4(l) \rightarrow Ca_3(PO_4)_2(s) + 3 H_2O(l)$ when 10.00 g $Ca(OH)_2$ and 10.00 g H_3PO_4 are mixed? This may yield integers, or it may yield decimal results that correspond closely to rational fractions. How many O₂ molecules does a mole of O₂(g) contain? Why doesn't a mole of carbon weigh 12 g? How many CO₂ molecules are there in a 15.00-g sample of carbon dioxide? A mole of substance is that amount that contains as many elementary units (atoms, molecules, or formula units, depending on the nature of the substance) as there are atoms in exactly 12 grams of an isotopically pure sample of ¹²C. In running real chemical reactions, it often occurs that less product is obtained than expected, for a variety of reasons. Consider a 15.00-g sample of CO₂ (m.w. = 44.01 u). Next to the atomic theory, the mole concept is the most fundamental unifying idea in all chemistry. Chemists are concerned with mass relationships in chemical reactions, usually run on a macroscopic scale (grams, kilograms, etc.). What is the empirical formula of an oxide of nitrogen whose composition is 25.94% nitrogen? To do this, start by dividing the smallest number of moles into each of the numbers of moles of elements (i.e., set the smallest number to 1). Therefore we have 24/4 = 6 sets of wheels and 15/1 = 15 sets of bodies. For CO₂ (m.w. = 44.01 u), mole CO₂ = 44.01 g CO₂ = 6.022 x 10²³ CO₂ molecules. Because each CO₂ molecule is composed of one carbon atom and two oxygen atoms, we could say that a mole of CO₂ contains one mole of carbon atoms and two moles of oxygen atoms. In such cases, the amount of product obtained is limited by the reactant in shortest supply, which is called the limiting reagent. What is the mass in grams of a single fluorine atom? For an element or compound that is composed of molecules, the molar mass in grams is numerically equal to its molecular weight in atomic mass units. To deal with the very large numbers of atoms and molecules in such samples, chemists developed the unit of the mole (abbreviated mol) and a unit of measure called the molar mass, which has units of g/mol. In the reaction $2 A + 3 B \rightarrow products$, if you have 0.500 mol A and 0.500 mol B, which is the limiting reagent? The results of chemical analysis are usually expressed in terms of weight percentages of each element in the compound, which can be converted into masses of each element for a given sample. Based on the wheels as the "limiting reagent" and their "stoichiometric" relationship to completed wagons (4 wheels/wagon), we could make only six wagons. The amount obtained is the actual yield. How many oxygen atoms are there in a 15.00-g sample of carbon dioxide. As we assemble wagons the wheels will run out (the "limiting reagent") before the bodies. Applying this approach to chemical reactions, if we take the number of moles of each reactant and divide that by its stoichiometric coefficient in the balanced equation, we will have a number for each that represents its number of reaction "sets". The molar mass based on a formula weight contains an Avogadro's number of formula units of the substance. [f.w. $Ca(OH)_2$] = 74.10 u; m.w. $Ca_3(PO_4)_2$] = 97.99 u; f.w. $Ca_3(PO_4)_2$] = 310.18 u] How many moles of CO₂ are there in this sample? What are the theoretical yields (in grams) of $CO_2(g)$ and $H_2O(l)$ expected from the reaction? You should try to answer the questions without referring to your textbook. This calculated amount of product is called the theoretical yield. If the compound also contains oxygen, its amount can be obtained by subtracting the found masses of carbon and hydrogen from the total mass of the sample. Once we know the numbers of moles, we can use the relationships between moles and molar masses of the various species to calculate masses of reactants and/or products, as needed. For most of our needs, a value of 6.022 x 10²³ will be sufficiently precise. The reagent that has the smallest number by this calculation is the limiting reagent; any other reagent is an excess reagent. For NaCl (f.w. = 58.44 u), mol NaCl = 58.44 g NaCl = 6.022 x 10²³ NaCl formula units. Note: Sodium chloride is an ionic compound so there are no molecules! The atomic weight of carbon is 12.0107 u, so a mole of carbon has a mass of 12.0107 g. In some cases, the limiting reagent may be the substance present with larger absolute amount (either in grams or moles), but used in greater quantity in the balanced equation. All calculations of the theoretical yield for the reaction (or any other stoichiometric calculations) must be based on the amount of the limiting reagent, using the stoichiometric relationships in the balanced chemical equation. The masses of each element can be used to calculate the numbers of moles of each element, from which the lowest whole number ratios between the moles of elements can be determined. How many moles of oxygen atoms does a mole of O₂(g) contain? A 1.638-g sample of propane is burned in excess oxygen. wt. In general, it is useful to think of a mole as just an Avogadro's number of things. Learning Objective Understand the relationship between the mole and Avogadro's number Understand the meaning of molar mass of a substance Understand how the mole concept is applied to determining empirical formulas from analytical data Understand how the mole concept allows prediction of the mass relationships between reactants and products in a chemical reaction Understand the concept of the limiting reagent Success Criteria Convert between numbers of atoms, moles, and mass of sample by using Avogadro's number and the appropriate molar mass Calculate the empirical formula of a compound from percent composition data Calculate mass relationships between reactants and products, based on a balanced chemical equation Be able to determine the limiting reagent in a chemical reaction and calculate mass relationships for a chemical reaction based on it One of the most important ideas in chemistry is the mole concept. For example, suppose we were building toy wagons and had 24 wheels and 15 wagon bodies. We would take the wheels in sets of four and the bodies in sets of one to build each wagon. Very often when we run a reaction between two or more substances, the amounts of reactants are not present in precisely the stoichiometric ratio indicated by the balanced chemical equation. In the case of molecular compounds, that number of molecules has a mass in grams that is numerically equal to the substance's molecular weight. If the molecular weight of the hydrocarbon is found to be 78.11 u, what is its molecular formula? [m.w. C_3H_8] = 44.09 u, m.w. CO_2] = 44.01 u, m.w. H_2O] = 18.02 u] If 4.750 g of $CO_2(g)$ was obtained from the combustion of 1.638 g of propane, what was the percent yield? = 26.981538 u), a sample weighing 26.981538 g would contain a mole of aluminum atoms; i.e., 26.981538 g Al = mole Al atoms The mass in grams of one mole of substance is called its molar mass. Why or why not? One of the most efficient is to see the amounts of each reagent in terms stoichiometric units, or what we might call "sets" (for want of a better term). Fluorine consists of a single isotope, ¹⁹F, with a mass of 19.00 u. In the complete combustion of propane, how many moles of $H_2O(l)$ are produced per mole of $CO_2(g)$? The molar mass of a molecular substance contains an Avogadro's number of molecules of the substance. How is the molecular formula of a molecular compound related to its empirical formula? A comparison of the actual yield to the theoretical yield, expressed as a statement of the percent yield; i.e., $\frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$. In the complete combustion of propane, how many moles of $CO_2(g)$ are produced per mole of $O_2(g)$? There are a number of ways to determine this, 12 g ¹²C (exactly) = mole ¹²C atoms The number of atoms in such a sample defines Avogadro's number (symbol NA), which has been experimentally determined to be NA = 6.0221367 x 10²³. For example, for aluminum (at. What is the empirical formula of the compound? Therefore, in any case where amounts of reactants are specified, determine the moles of each present, and then determine which reactant is the limiting reagent. These mass relationships, made through moles, are called stoichiometry (Gk stoicheon, element + -metry, measure).

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