Mole to mole stoichiometry worksheet

1 atom fe + 1 atom of S ----> 1 molecule of FeS 10 atoms of Fe + 10 atoms of S ----> 10 molecules of FeS 55.8 mg Fe + 32.1 mg S ----> 87.9 mg FeS 5.58 g Fe + 3.21 g S ----> 8.79 g FeS 55.8 g Fe + 32.1 g S ----> 87.9 g FeS Name:

You should try answering the questions without referring to your textbook. If you're stuck, try asking another group for help. Chemists are concerned about mass conditions in chemical reactions, usually running on a macroscopic scale (grams, kilograms, etc.). To deal with the very large number of atoms and molecules in such samples, chemists developed the device of the mole (abbreviated mole) and a unit of measurement called the molar mass, which has units of g / mole. Alongside the atomic theory, the mole concept is the most basic unifying idea in all chemistry. Learning objectives Understanding the relationship between mole and Avogadros number Understanding the importance of molar mass of a substance Understand how the mole concept is used to determine empirical formulas from analytical data Understand how the mole concept allows the prediction of the mass conditions between reactants and products in a chemical reaction Understand the concept of limiting reagent success criteria Convert between the number of atoms, Mol, and mass of sample using Avogadros number and proper molar mass Calculate the empirical formula of a compound from percentage data Calculate mass relationships between reactants and products, based on a balanced chemical equation Could determine the limiting reagent in a chemical reaction and calculate mass relationships for a chemical reaction based on what One of the most important ideas in chemistry is the mole concept. A mole of substance is the amount that contains so many elementary devices (atoms, molecules, or formula units, depending on the nature of the substance) as there are atoms in exactly 12 grams of an isotopic clean sample of 12C. 12 g 12C (exact) = 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 most of our needs, a value of 6.022 x 1023 will be sufficiently precise. It follows that if we had a mole of atoms of another element, that sample would weigh its atomic weight in grams. For example, for aluminum (at. wt. = 26.981538 u), a sample weighing 26.981538 g would contain a mole of aluminum atoms; it will want 26.981538 g Al = mole Al atoms The mass in grams of a mole of substance is called the molar mass. For an element or compound consisting of molecules, the molar mass in grams is numerically similar to its molecular weight in atomic mass units. The molar mass of a molecular substance contains an Avogadros number of molecules of the drug. For CO2 (m.w. = 44.01 u), mole = 44.01 g CO2 = 6.022 x 1023 CO2 molecules Because each CO2 molecule consists of one carbon atom and two oxygen atoms, we can say that a mole of CO2 contains a mole of carbon atoms and two moles of oxygen atoms. In general, it is useful to think of a mole as only an Avogadros number of things. In the case of molecular compounds, it has the number of molecules a mass in grams that is numerically similar to the molecular weight of the substance. 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 similar to the formula weight in atomic mass units. The molar mass based on a formula weight contains an Avogadros number of formula units of the drug. 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. Why doesn't a mole of carbon weigh 12g? The atomic weight of oxygen is 16.00 u. What is the mass of a mole of O2(g)? How many O2 molecules does a mole of O2 (g)contain? How many moles of oxygen atoms does a mole of O2 (g)? The mole is sometimes described as the chemist's dozen. How's a mole like a dozen? Consider a 15.00 g sample of CO2 (m.w. = 44.01 u). How many moles of CO2 are there in this sample? How many CO2 molecules are found in a 15.00g sample of carbon dioxide? How many oxygen atoms there are in a 15.00 g sample of carbon dioxide. Fluoride consists of a single isotope, 19F, with a mass of 19.00 u. What is the mass in grams of a single fluoride atom? The elementary composition of a compound can be determined experimentally by a variety of techniques. The results of chemical analysis are usually expressed in the form of weight percentages of each element of the compound, which can be converted into masses of each element for a given sample. The masses in each element can be used to calculate the number of moles of each element, from which the lowest entire number ratios between the moles of elements can be determined. These conditions are the same as the ratio of the number of individual atoms of each element in the empirical formula. The strategy to convert analytical data to an empirical formula usually uses the following steps: Convert weight percentages to grams of each item. Often it is useful to assume a sample size of exactly 100 grams; then the specified percentages are numerically equal to the number of grams of each item. Convert grams of each element to moles of each element, using atomic scales. Find the lowest full number ratios among the moles of items. To do this, start by dividing the minimum number of moles into each of the number of moles of items (that is, set the minimum number to 1). This can result in integer, or it may produce decimal results that correspond to rational For example,$1,25 : 2,75 = 11/4 :23/4 = 5 :11$ Type the empirical formula, using the same whole number ratios between atoms of each element as the ratio of moles of elements. If the molecular weight is known, divide the formula weight of the empirical formula into the molecular weight to determine the number of formula units in the molecular formula. Using this integer factor, multiply all lowered fonts (including any implied 1's) in the empirical formula to achieve the molecular formula. If you have data for the percentage of a compound, item by item, do you need to know the size of the sample to find out the empirical formula? Why or why not? How is the molecular formula of a molecular compound related to its empirical formula? A compound has been found to contain 54.52% C, 9.17% H, and 36.31% O. What is the empirical formula of the compound? If the compound is found to have molecular weight of 88.12 u, what is molecular formula? What is the empirical formula of a nitric oxide whose composition is 25.94% nitrogen? An experimental method of determining the composition of organic compounds is combustion analysis, in which a weighted sample of the compound is burned in excess oxygen. In all cases, all the carbon in the connection to CO2 is converted, and all hydrogen is converted into H2O, which can be separated from each other and weighed. The masses of carbon and hydrogen in the original sample can be calculated based on the weights of CO2 and H2O. If the compound also contains oxygen, the amount can be obtained by pulling the found masses of carbon and hydrogen from the total mass of the sample. These masses can then be converted into moles, from which the empirical formula can be obtained. A 2,554g sample of a specific hydrocarbon is burned in excess oxygen, producing 8,635 g of CO2(g) and 1,768 g H2O(l). If the molecular weight of the hydrocarbon is found to be 78.11 u, what is the molecular formula? [m.w. CO2 = 44.01 u; m.w. H2O = 18.02 u] For burning (combustion) of propane gas is the balanced equation $C_3H_8(g) + 5 O_2(g) \rightarrow 3 CO_2(g) + 4 H_2O(l)$ When we first detected reaction equations, we thought about this in relation to reaktant and product species; Eg. For each 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 relationship between individual reactant and product species and between moles of these species multiplication of the constant Avogadros number. Therefore, the relationships between moles of reactants and products are the same as between individual reactante and product species; Eg. For each 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 when it comes to mole relationships. When we know the number of moles, we can use between moles and molar masses of the different species to calculate masses of reactants and / or products, as needed. These mass conditions, made through moles, are called stoichiometry (Gk stoicheon, element + -metry, target). With the help of mole and mass conditions, we can calculate the mass of the product to be produced from a given amount of reactant when completely consumed in the reaction. This calculated amount of product is called the theoretical return. When driving real chemical reactions, it often occurs that smaller product is obtained than expected, for a variety of reasons. The amount achieved is the actual return. A comparison of the actual return on the theoretical return, expressed as a percentage, is a statement of the percentage return; ie $\text{yield} = \dfrac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$ In complete propane combustion, how many moles of $CO_2(g)$ are produced per mole by $O_2(g)$)? In complete combustion of propane, how many moles of $H_2O(l)$ are produced per mole of $O_2(g)$)? In complete combustion of propane, how many moles of $H_2O(l)$ are produced per mole of $CO_2(g)$)? A 1,638 g sample propane is burned in excess oxygen. What are the theoretical yields (in grams) of $CO_2(g)$ and $H_2O(l)$ expected from the reaction? [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 $CO_2(g)$ was extracted from the combustion of 1,638 g of propane, what was the percentage yield? Very often when we run a reaction between two or more drugs, the amounts of reactants are not present in exactly the stoichiometric ratio indicated by the balanced chemical equation. In such cases, a reactant may be present in the deficiency of the commodity, while other reactants may be present in abundance. Assuming complete reaction, the reactant in the shortest supply will be completely consumed, but some amounts of the other reagents will remain after the reaction is finished. In such cases, the amount of the product obtained is limited by the reactant in the shortest supply, which is called the limiting reagent. It is important to realize that the limiting reagent is present in the shortest supply on the basis of stoichiometry of balanced chemical equation in moles; it will want the mole conditions to be implied by the balanced equation. In some cases, the limiting reagent may be the drug found with a greater absolute amount (either in grams or moles), but is used in greater quantity in the balanced equation. In any case, the theoretical return of the product will always be limited by the stoichiometric ratio between limiting reagent and products. Therefore, in any case, where amounts of reactants are specified, the moles of each present determine, and then determine which reaktant is the limiting reagent. All calculations of the theoretical yield for the reaction (or other stoichiometric calculations) must be based on the amount of the limiting relationships in the balanced chemical equation. How do we know which of two or more reactants are limiting? There are a number of ways to determine this. One of the most effective is to see the amounts of each reagent in the form of stoichiometric units, or what we might call sets (to want a better term). For example, suppose we built toy carts and had 24 wheels and 15 carriage bodies, we would take the wheels in sets of four and the bodies in the set of one to build each carriage. Therefore, we have 24/4 = 6 sets of wheels and 15/1 = 15 sets of organs. When we mount the carts the wheels will run out (limiting reagent) before the corpses. Based on the wheels as limiting reagent and their stoichiometric ratio to completed wagons (4 wheels / carriage), we could make only six wagons. By doing this, we would use six bodies, and we would have 15 - 6 = 9 organs left. By applying this approach to chemical reactions, if we take the number of moles of each reagent and divide it by its stoichiometric coefficient in the balanced equation, we will have a number for each representing the number of reaction sets. The reagent that has the smallest number in this calculation is the limiting reagent; other reagents are a superfluous reagent. We then use the number of moles of the limiting reagent (not the calculated number seen) as the basis for all our further calculations, such as theoretical yield or amount of non-limiting reagent used. In short, all calculations are based on the moles of the limiting reagent and the stoichiometric relationships suggested by the balanced chemical equation. Define what is meant by the terms that limit reagent and excess reagent. In the reaction \(2 A + 3 B \rightarrow products\), if you have 0.500 mol A and 0.500 moles B, which is limiting reagent? How much of excess reagent will be left, if the competitive reaction takes place? What is the theoretical return of \(Ca_3(PO_4)_2(s)\) of the reaction \[\ce{3 Ca(OH)2(s) + 2 H3PO4(l) \rightarrow Ca3(PO4)2(s) + 3 H2O(l)}\] when 10.00 g \(\ce{Ca(OH)2}\) and 10.00 g \(\ce{H3PO4}\) are mixed? [f.w. \(\ce{Ca(OH)2}\) = 74.10 u; m.w. \(\ce{H3PO4}\) = 97.99 u; f.w. \(\ce{Ca3(PO4)2}\) = 310.18 u] u] u]