Limiting excess reactants worksheet answers

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That said, the coefficients of the balanced equation have nothing to do with the actual quantity of reactants you start with, as you can mix any amount you choose, but clearly the maximum yield (theoretical yield) must be limited by the reactant that gets consumed up first, the limiting reagent. The reagent that remains is called the excess reagent. This can be easily understood by the analogy of making bicycles, where each bike requires 2 tires and one frame. The "equation" becomes: 1 frames + 2 tires --> 1 bike As you can see, the "balanced equation" simply tells us the ratio of number of frames and tires to the number of bikes made. So let's look at a few case scenarios: A) How many bikes can we theoretically make with 10 frames and 16 tires? Which is the "limiting reagent" and which is the "excess reagent"? With 10 frames, we can make 10 bikes. However, we also need tires to make a bike. With 16 tires, we can make 8 bikes (2 tires per bike). So, in reality we can only make 8 bikes, not 10, because the "limiting reagent" is the tires and the "excess" reagent is the frames. We have 2 frames left over. The theoretical yield of bikes will be 8 bikes (based on the limiting reagent) B) How many bikes can we theoretically make with 10 frames and 20 tires? Which is the "limiting reagent" and which is the "excess reagent"? 10 frames makes 10 bikes. 20 tires also make 10 bikes. So, we have a stoichiometric proportion and there is nothing left over. C) How many bikes can we theoretically make with 10 frames and 30 tires? Which is the "limiting reagent" and which is the "excess reagent"? 10 frames make 10 bikes. 30 tires make 15 bikes. However, we are not making 15 bikes because we ran out of frames after 10 bikes. So we used only 20 tires, with 10 tires left over. So, this time, the limiting reagent is the frames, and the excess reagent is the tires. The theoretical yield of bikes is 10 (based on the limiting reagent). With that being said, let's recap with a few points: Stoichiometric Proportions: Reactants are mixed in the ratios defined by their stoichiometric coefficients. If a reaction proceeds to completion, everything is consumed. NonStoichiometric Proportions: Reactants are mixed in ratios that are different than the stoichiometric coefficients. One species runs out first (Limiting Reagent), while another is not completely consumed (Excess Reagent). Theoretical Yield: the maximum possible yield based on the complete consumption of the limiting reagent Are the limiting reagents always completely consumed? No, only if the reaction goes to completion. There can be many different reasons why the limiting reagent is not completely consumed, and these will be covered in detail in later chapters of this text. The theoretical yield is the maximum amount of product that would be produced through the complete consumption of the limiting reagent. Excess Reagent: The quantity (mole or mass) left over after the complete consumption of the limiting reagent. Quantity Excess = Initial Quantity - Consumed Quantity. Where quantity can be moles or mass. Identify moles of all reactants present. If given mass, divide by formula weight to convert moles (this is the mass to mole step from the section 4.1. Divide moles of each reactant by it's stoichiometric coefficient. This is the denominator of the mole-to-mole step in section 4.1. Smallest number indicates limiting reagent. Multiply by stoichiometric coefficient of species you are solving for, and answer the question . This is the numerator of the mole-to-mole step in section 4.1. If you are after moles, you are finished, if you are after mass, you need to use the molar mass of product to convert moles product to grams mass product, which is the mass-to-mole step in section 4.1. Limiting Reagent Problem Silver tarnishes in the presence of hydrogen sulfide and oxygen due to the following reaction 4Ag + 2H2S + O2 ----> 2Ag2S + 2H2O 6:07 minute YouTube video showing shortcut method for determining limiting reagent and theoretical yield. [corrigenda: after dividing moles Ag by stoichiometric coef. the value is 0.00556, not 0.0556, but this this did not effect the solution as both are more than the value for oxygen of 0.00500, so oxygen is still the limiting reagent). 4:36 minute YouTube determining the excess reagents after the complete consumption of the limiting reagent. To Calculate moles of Excess reagent you subtract the amount consumed by the complete consumption of the limiting reagent from the initial quantity of the excess reagent. Click the following link for more practice on limiting reagents. You will have feedback and hints to help guide you. Do the Following Worksheet Limiting Reagent Worksheet Limiting Reagent Worksheet Key When there is not enough of one reactant in a chemical reaction, the reaction stops abruptly. To figure out the amount of product produced, it must be determined which reactant will limit the chemical reaction (the limiting reagent) and which reactant is in excess (the excess reagent). One way of finding the limiting reagent is by calculating the amount of product that can be formed by each reactant; the one that produces less product is the limiting reagent. The following scenario illustrates the significance of limiting reagents. In order to assemble a car, 4 tires and 2 headlights are needed (among other things). In this example, imagine that the tires and headlights are reactants while the car is the product formed from the reaction of 4 tires and 2 headlights. \[4 \, \text{Tires} + 2 \, \text{Headlights} = 1 \, \text{Car} onumber\] If you have 20 tires and 14 headlights, how many cars can be made? With 20 tires, 5 cars can be produced because there are 4 tires to a car. With 14 headlights, 7 cars can be built (each car needs 2 headlights). Although more cars can be made from the headlights available, only 5 full cars are possible because of the limited number of tires available. In this case, the headlights are in excess. Because the number of cars formed by 20 tires is less than number of cars produced by 14 headlights, the tires are the limiting reagent (they limit the full completion of the reaction, in which all of the reactants are used up). This scenario is illustrated below: + = Figure 1: The synthesis reaction of making a car. Images used from Wikipedia with permission. The initial condition is that there must be 4 tires to 2 headlights. The reactants must thus occur in that ratio; otherwise, one will limit the reaction. There are 20 tires and 14 headlights, so there are two ways of looking at this problem. For 20 tires, 10 headlights are required, whereas for 14 headlights, 28 tires are required. Because there are not enough tires (20 tires is less than the 28 required), tires are the limiting "reactant." The limiting reagent is the reactant that is completely used up in a reaction, and thus determines when the reaction stops. From the reaction stoichiometry, the exact amount of reactant needed to react with another element can be calculated. If the reactants are not mixed in the correct stoichiometric proportions (as indicated by the balanced chemical equation), then one of the reactants will be entirely consumed while another will be left over. The limiting reagent is the one that is totally consumed; it limits the reaction from continuing because there is none left to react with the in-excess reactant. There are two ways to determine the limiting reagent. One method is to find and compare the mole ratio of the reactants used in the reaction (approach 1). Another way is to calculate the grams of products produced from the given quantities of reactants; the reactant that produces the smallest amount of product is the limiting reagent (approach 2). How to Find the Limiting Reagent: Approach 1 Find the limiting reagent by looking at the number of moles of each reactant. Determine the balanced chemical equation for the chemical reaction. Convert all given information into moles (most likely, through the use of molar mass as a conversion factor). Calculate the mole ratio from the given information. Compare the calculated ratio to the actual ratio. Use the amount of limiting reactant to calculate the amount of product produced. If necessary, calculate how much is left in excess of the non-limiting reagent. How to Find the Limiting Reagent: Approach 2 Find the limiting reagent by calculating and comparing the amount of product each reactant will produce. Balance the chemical equation for the chemical reaction. Convert the given information into moles. Use stoichiometry for each individual reactant to find the mass of product produced. The reactant that produces a lesser amount of product is the limiting reagent. The reactant that produces a larger amount of product is the excess reagent. To find the amount of remaining excess reactant, subtract the mass of excess reagent consumed from the total mass of excess reagent given. Example \(\PageIndex{1}\): Photosynthesis Consider respiration, one of the most common chemical reactions on earth. \[\ce{ C6H_{12}O6 + 6 O_2 \rightarrow 6 CO2 + 6 H2O} + \rm{energy}\] What mass of carbon dioxide forms in the reaction of 25 grams of glucose with 40 grams of oxygen? Solution When approaching this problem, observe that every 1 mole of glucose (\(C_6H_{12}O_6\)) requires 6 moles of oxygen to obtain 6 moles of carbon dioxide and 6 moles of water. Step 1: Determine the balanced chemical equation for the chemical reaction. The balanced chemical equation is already given. Step 2: Convert all given information into moles (most likely, through the use of molar mass as a conversion factor). \[\mathrm{25\:g \times \dfrac{1\: mol}{180.06\:g} = 0.1388\: mol\: C_6H_{12}O_6} onumber\] \[\mathrm{40\:g \times \dfrac{1\: mol}{32\:g} = 1.25\: mol\: O_2} onumber\] Step 3: Calculate the mole ratio from the given information. Compare the calculated ratio to the actual ratio. a. If all of the 1.25 moles of oxygen were to be used up, there would need to be \(\mathrm{1.25 \times \dfrac{1}{6}}\) or 0.208 moles of glucose. There is only 0.1388 moles of glucose available which makes it the limiting reactant. \[1.25 \; \rm{mol} \; O_2 \times \dfrac{ 1 \; \rm{mol} \; C_6H_{12}O_6}{6\; \rm{mol} \; O_2}= 0.208 \; \rm{mol} \; C_6H_{12}O_6 onumber\] b. If all of the 0.1388 moles of glucose were used up, there would need to be 0.1388 x 6 or 0.8328 moles of oxygen. Because there is an excess of oxygen, the glucose amount is used to calculate the amount of the products in the reaction. \[0.1388\; \rm{ mol}\; C_6H_{12}O_6 \times \dfrac{6 \; \rm{mol} \;O_2}{1 \; \rm{mol} \; C_6H_{12}O_6} = 0.8328 \; \rm{mol}\; O_2 onumber\] If more than 6 moles of O2 are available per mole of C6H12O6, the oxygen is in excess and glucose is the limiting reactant. If less than 6 moles of oxygen are available per mole of glucose, oxygen is the limiting reactant. The ratio is 6 mole oxygen per 1 mole glucose, OR 1 mole oxygen per 1/6 mole glucose. This means: 6 mol O2 / 1 mol C6H12O6 . Therefore, the mole ratio is: (0.8328 mol O2)/(0.208 mol C6H12O6) This gives a 4.004 ratio of \(\ce{O2}\) to \(\ce{C6H12O6}\). Step 4: Use the amount of limiting reactant to calculate the amount of CO2 or H2O produced. For carbon dioxide produced: \(\mathrm{0.1388\: moles\: glucose \times \dfrac{6}{1} = 0.8328\: moles\: carbon\: dioxide}\). Step 5: If necessary, calculate how much is left in excess. 1.25 mol - 0.8328 mol = 0.4172 moles of oxygen left over Example \(\PageIndex{2}\): Oxidation of Magnesium Calculate the mass of magnesium oxide possible if 2.40 g \(\ce{Mg}\) reacts with 10.0 g of \(\ce{O_2}\) \[\ce{ Mg +O_2 \rightarrow MgO} onumber\] Solution Step 1: Balance equation \[\ce{2 Mg + O_2 \rightarrow 2 MgO} onumber\] Step 2 and Step 3: Converting mass to moles and stoichiometry \[\mathrm{2.40\:g\: Mg \times \dfrac{1.00\: mol\: Mg}{24.31\:g\: Mg} \times \dfrac{2.00\: mol\: MgO}{2.00\: mol\: Mg} \times \dfrac{40.31\:g\: MgO}{1.00\: mol\: MgO} = 3.98\:g\: MgO} onumber\] \[\mathrm{10.0\:g\: O_2\times \dfrac{1\: mol\: O_2}{32.0\:g\: O_2} \times \dfrac{2\: mol\: MgO}{1\: mol\: O_2} \times \dfrac{40.31\:g\: MgO}{1\: mol\: MgO} = 25.2\: g\: MgO} onumber\] Step 4: The reactant that produces a smaller amount of product is the limiting reagent \(\ce{Mg}\) produces less \(\ce{MgO}\) than does \(\ce{O2}\) (3.98 g MgO vs. 25.2 g MgO), therefore Mg is the limiting reagent in this reaction. Step 5: The reactant that produces a larger amount of product is the excess reagent O2 produces more amount of MgO than Mg (25.2g MgO vs. 3.98 MgO), therefore O2 is the excess reagent in this reaction. Step 6: Find the amount of remaining excess reactant by subtracting the mass of the excess reagent consumed from the total mass of excess reagent given. Mass of excess reagent calculated using the limiting reagent: \[\mathrm{2.40\:g\: Mg \times \dfrac{1.00\: mol\: Mg}{24.31\:g\: Mg} \times \dfrac{1.00\: mol\: O_2}{2.00\: mol\: Mg} \times \dfrac{32.0\:g\: O_2}{1.00\: mol\: O_2} = 1.58\:g\: O_2} onumber\] OR Mass of excess reagent calculated using the mass of the product: \[\mathrm{3.98\:g\: MgO \times \dfrac{1.00\: mol\: MgO}{40.31\:g\: MgO} \times \dfrac{1.00\: mol\: O_2}{2.00\: mol\: MgO} \times \dfrac{32.0\:g\: O_2}{1.00\: mol\: O_2} = 1.58\:g\: O_2} onumber\] Mass of total excess reagent given – mass of excess reagent consumed in the reaction 10.0g – 1.58g = 8.42g O2 is in excess. Example \(\PageIndex{3}\): Limiting Reagent What is the limiting reagent if 76.4 grams of \(C_2H_3Br_3\) were reacted with 49.1 grams of \(O_2\)? \[\ce{4 C_2H_3Br_3 + 11 O_2 \rightarrow 8 CO_2 + 6 H_2O + 6 Br_2} onumber\] Solution Using Approach 1: A. \[\mathrm{76.4\:g \times \dfrac{1\: mole}{266.72\:g} = 0.286\: moles\: of\: C_2H_3Br_3} onumber\] \[\mathrm{49.1\: g \times \dfrac{1\: mole}{32\:g} = 1.53\: moles\: of\: O_2} onumber\] B. Assuming that all of the oxygen is used up, \(\mathrm{1.53 \times \dfrac{4}{11}}\) or 0.556 moles of C2H3Br3 are required. Because there are only 0.286 moles of C2H3Br3 available, C2H3Br3 is the limiting reagent. Using Approach 2: \[\mathrm{76.4\:g\: C_2H_3Br_3 \times \dfrac{1\: mol\: C_2H_3Br_3}{266.72\:g\: C_2H_3Br_3} \times \dfrac{8\: mol\: CO_2}{4\: mol\: C_2H_3Br_3} \times \dfrac{44.01\:g\: CO_2}{1\: mol\: CO_2} = 25.2\:g\: CO_2} onumber\] \[\mathrm{49.1\:g\: O_2 \times \dfrac{1\: mol\: O_2}{32\:g\: O_2} \times \dfrac{8\: mol\: CO_2}{11\: mol\: O_2} \times \dfrac{44.01\:g\: CO_2}{1\: mol\: CO_2} = 49.1\:g\: CO_2} onumber\] Therefore, by either method, C2H3Br3is the limiting reagent. Example \(\PageIndex{4}\): Limiting Reagent What is the limiting reagent if 78 grams of \(\ce{Na2O2}\) were reacted with 29.4 grams of \(\ce{H2O}\)? Solution Using Approach 1 A. \[\mathrm{78\:g \times \dfrac{1\: mol}{77.96\:g} = 1.001\: moles\: of\: Na_2O_2} onumber\] \[\mathrm{29.4\:g \times \dfrac{1\: mol}{18\:g}= 1.633\: moles\: of\: H_2O} onumber\] B. Assume that all of the water is consumed, \(\mathrm{1.633 \times \dfrac{2}{2}}\) or 1.633 moles of Na2O2 are required. Because there are only 1.001 moles of Na2O2, it is the limiting reactant. Using Approach 2: \\[\mathrm{78\:g\: Na_2O_2 \times \dfrac{1\: mol\: Na_2O_2}{77.96\:g\: Na_2O_2} \times \dfrac{4\: mol\: NaOH}{2\: mol\: Na_2O_2} \times \dfrac{40\:g\: NaOH}{1\: mol\: NaOH} = 80.04\:g\: NaOH} onumber\] Using either approach gives Na2O2 as the limiting reagent. Example \(\PageIndex{5}\): Excess Reagent How much the excess reagent remains if 24.5 grams of CoO is reacted with 2.58 grams of O2? \[\ce{4 CoO + O_2 \rightarrow 2 Co_2O_3} onumber\] Solution A. \[\mathrm{24.5\:g \times \dfrac{1\: mole}{74.9\:g}= 0.327\: moles\: of\: CoO} onumber\] \[\mathrm{2.58\:g \times \dfrac{1\: mole}{32\:g}= 0.0806\: moles\: of\: O_2} onumber\] B. Assuming that all of the oxygen is used up, \(\mathrm{0.0806 \times \dfrac{4}{1}}\) or 0.3225 moles of \(CoO\) are required. Because there are 0.327 moles of CoO, CoO is in excess and thus O2 is the limiting reactant. C. 0.327 mol - 0.3224 mol = 0.0046 moles left in excess. Example \(\PageIndex{6}\): Identifying the Limiting Reagent Will 28.7 grams of \(SiO_2\) react completely with 22.6 grams of \(H_2F_2\)? If not, identify the limiting reagent. \[\ce{SiO_2+ 2 H_2F_2 \rightarrow SiF_4+ 2 H_2O} onumber\] Solution A. \[\mathrm{28.7\:g \times \dfrac{1\: mole}{60.08\:g} = 0.478\: moles\: of\: SiO_2} onumber\] \[\mathrm{22.6\:g \times \dfrac{1\: mole}{39.8\:g} = 0.568\: moles\: of\: H_2F_2} onumber\] B. There must be 1 mole of SiO2 for every 2 moles of H2F2 consumed. Because the ratio is 0.478 to 0.568, 28.7 grams of SiO2 do not react with the H2F2. C. Assuming that all of the silicon dioxide is used up, \(\mathrm{0.478 \times \dfrac{2}{1}}\) or 0.956 moles of H2F2 are required. Because there are only 0.568 moles of H2F2, it is the limiting reagent.

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