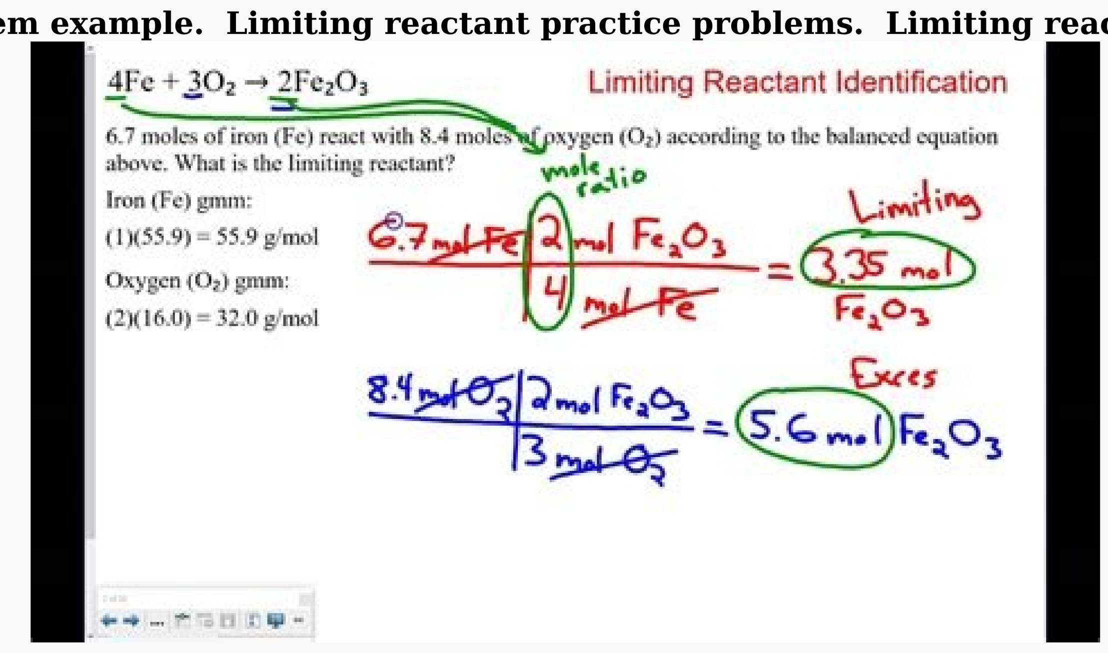
Limiting reactant practice problems with answers pdf

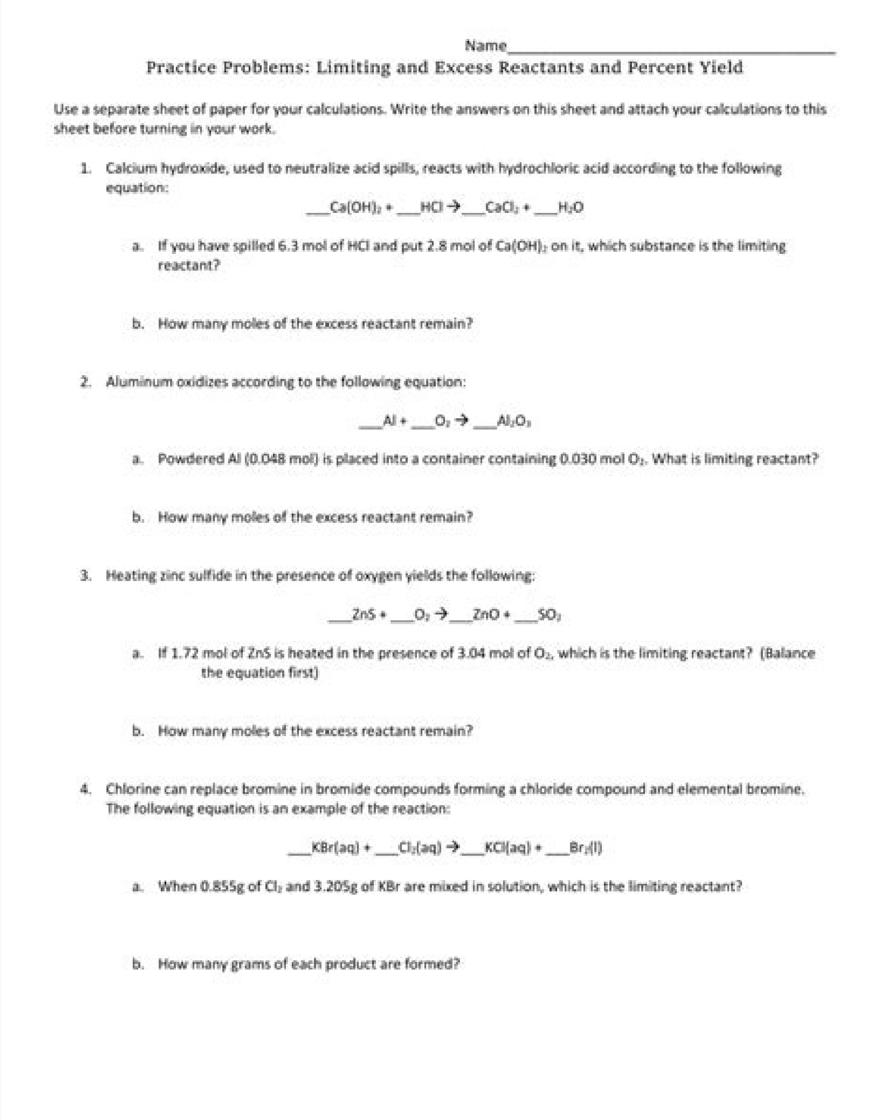
Limiting reactant problem example. Limiting reactant practice problems. Limiting reactant worksheet answers.



Limiting reactant practice problems with answers.

This is a set of practice problems to help master the concept of limiting reactant which is critical in calculating the amount of product that can be obtained in a chemical reaction. Remember, if the reactants are not in stoichiometric ratio, one of them is the limiting reactant (LR), and the other is in excess.

When solving problems involving a limiting reactant, keep in mind that we cannot look at the moles of the reactants and determine which one is the limiting reactant. The limiting reactant is the one that produces less product and not necessarily the one that is present in fewer amounts. The correct approach is to calculate the moles of both reactants and, using molar conversions, determine which one produces less product. Check Also 1. Which statement about limiting reactant is correct? a) The limiting reactant is the one in a smaller quantity. b) The limiting reactant is the one in greater quantity. c) The limiting reactant is the one producing less product. d) The limiting reactant is the one producing more product. The limiting reactant is the one producing less product than any of the other reactants. 2. Find the limiting reactant for each initial amount of reactants. $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$ a) 2 mol of NH_3 and 2 mol of O_2 b) 2 mol of NH_3 and 3 mol of O_2 c) 3 mol of NH_3 and 3 mol of O_2 d) 3 mol of NH_3 and 2 mol of O_2 Note: This is not a multiple-choice question. Each row represents a separate question where you need to determine the limiting reactant.



Name___________________

Practice Problems: Limiting and Excess Reactants and Percent Yield

Use a separate sheet of paper for your calculations. Write the answers on this sheet and attach your calculations to this sheet before turning in your work.

1. Calcium hydroxide, used to neutralize acid spills, reacts with hydrochloric acid according to the following equation:

   ___$Ca(OH)_2$ + ___HCl → ___$CaCl_2$ + ___H_2O

   a. If you have spilled 6.3 mol of HCl and put 2.8 mol of $Ca(OH)_2$ on it, which substance is the limiting reactant?

   b. How many moles of the excess reactant remain?

2. Aluminum oxidizes according to the following equation:

   ___Al + ___O_2 → ___Al_2O_3

   a. Powdered Al (0.048 mol) is placed into a container containing 0.030 mol O_2. What is limiting reactant?

   b. How many moles of the excess reactant remain?

3. Heating zinc sulfide in the presence of oxygen yields the following:

   ___ZnS + ___O_2 → ___ZnO + ___SO_2

   a. If 1.72 mol of ZnS is heated in the presence of 3.04 mol of O_2, which is the limiting reactant? (Balance the equation first)

   b. How many moles of the excess reactant remain?

4. Chlorine can replace bromine in bromide compounds forming a chloride compound and elemental bromine. The following equation is an example of the reaction:

   ___KBr(aq) + ___Cl_2(aq) → ___KCl(aq) + ___Br_2(l)

   a. When 0.855g of Cl_2 and 3.205g of KBr are mixed in solution, which is the limiting reactant?

   b. How many grams of each product are formed?

a) 2 mol of NH_3 and 2 mol of O_2 b) 2 mol of NH_3 and 3 mol of O_2 c) 3 mol of NH_3 and 3 mol of O_2 d) 3 mol of NH_3 and 2 mol of O_2 The limiting reactant is the one producing less product. So, to determine the limiting reactant for each pair, we need to pick a product (NO or H_2O) and see whether NH_3 or O_2 will produce less product for the given mole ratio. Let's do the calculations based on the amount of NO that can be formed. a)

$$n(NO) = 2\ \cancel{mol\ NH_3} \times \frac{4\ mol\ NO}{4\ \cancel{mol\ NH_3}} = 2\ mol$$

$$n(NO) = 2\ \cancel{mol\ O_2} \times \frac{4\ mol\ NO}{5\ \cancel{mol\ O_2}} = 1.6\ mol$$

So, for (a), O_2 is the limiting reactant.

b)

$$n(NO) = 2\ \cancel{mol\ NH_3} \times \frac{4\ mol\ NO}{4\ \cancel{mol\ NH_3}} = 2\ mol$$

$$n(NO) = 3\ \cancel{mol\ O_2} \times \frac{4\ mol\ NO}{5\ \cancel{mol\ O_2}} = 2.4\ mol$$

So, for (b), NO is the limiting reactant. c)

$$n(NO) = 3\ \cancel{mol\ NH_3} \times \frac{4\ mol\ NO}{4\ \cancel{mol\ NH_3}} = 3\ mol$$

$$n(NO) = 3\ \cancel{mol\ O_2} \times \frac{4\ mol\ NO}{5\ \cancel{mol\ O_2}} = 2.4\ mol$$

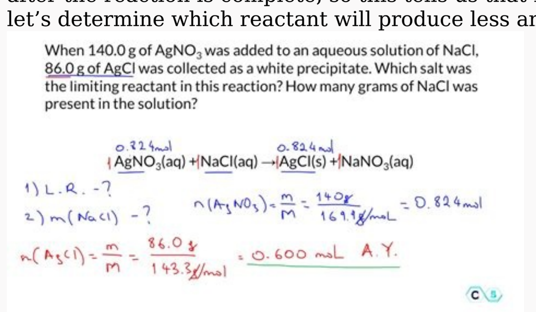
So, for (c), O_2 is the limiting reactant.

d)

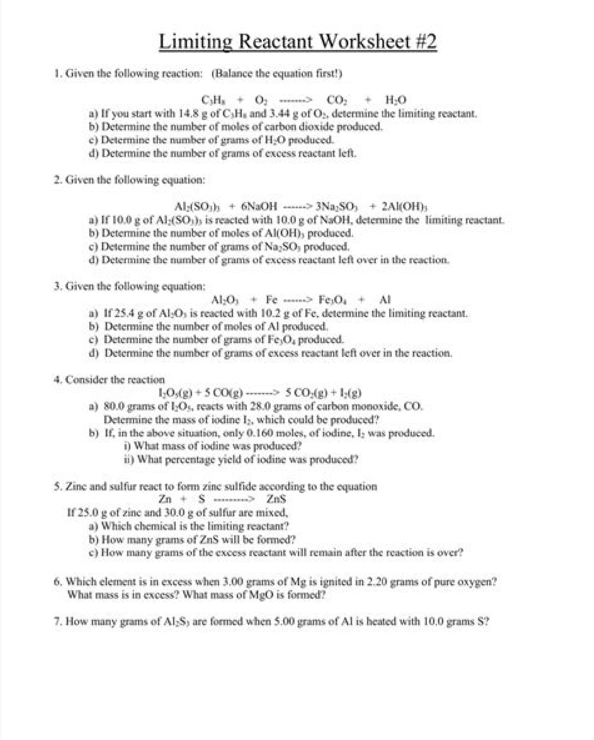
$$n(NO) = 3\ \cancel{mol\ NH_3} \times \frac{4\ mol\ NO}{4\ \cancel{mol\ NH_3}} = 3\ mol$$

$$n(NO) = 2\ \cancel{mol\ O_2} \times \frac{4\ mol\ NO}{5\ \cancel{mol\ O_2}} = 1.6\ mol$$

So, for (d), O_2 is the limiting reactant. The shorter way of determining the limiting reactant is to divide the moles of a given reactant by its coefficient in the chemical reaction. This is the effective quantity of the reactant as it takes into consideration not only the amount of the reactant but also its molar ratio in the reaction. Let's put the equation and the moles here and determine the LR by the ratio of the moles and coefficient: $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$ a) 2 mol of NH_3 and 2 mol of O_2 b) 2 mol of NH_3 and 3 mol of O_2 c) 3 mol of NH_3 and 3 mol of O_2 d) 3 mol of NH_3 and 2 mol of O_2 So, for (a), we are comparing 2/4 mol of NH_3 with 2/5 mol of O_2, and therefore, O_2 is the LR. For (b), we are comparing 2/4 mol of NH_3 with 3/5 mol of O_2, and therefore, NH_3 is the LR. For (c), we are comparing 3/4 mol of NH_3 with 3/5 mol of O_2, and therefore, O_2 is the LR. For (d), 3 mol of NH_3 and 3/4 mol of NH_3 with 2/5 mol of O_2, and therefore, O_2 is the LR. 3. How many g of hydrogen are leftover in producing ammonia when 14.0 g of nitrogen is reacted with 8.0 g of hydrogen? $N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$ As stated in the problem, there is going to be some H_2 left over after the reaction is complete, so this tells us that H_2 is in excess and N_2 is the limiting reactant. Remember, limiting reactant is consumed completely in a chemical reaction. Remember also that stoichiometric calculations need to be done based on the moles of limiting reactant, so let's first determine the limiting reactant. Limiting reactant: Now, let's determine which reactant will produce less ammonia. It would be very beneficial to practice and determine the moles in the chemical reaction without doing mole ratio calculations.



a) Alternative to dimensional analysis. For example, the chemical equation indicates that from every one mole of N_2 there is 2 moles of NH_3 forming. This means that every 2 moles of N_2 will produce 4 moles of NH_3, every 3 moles of N_2 will produce 6 moles of NH_3, every 10 moles of N_2 will produce 20 moles of NH_3 etc. Ammonia is always two times more. So, in this case, 0.50 mol N_2 will produce 1.0 mol NH_3. On the other hand, every 3 moles of H_2 produces 2 moles of NH_3, so ammonia is always going to be 1.5 times less than the H_2. We can put the moles on top of H_2 and mark the moles of NH_3 as X. 4.0 mol X mol $N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$ Then cross multiply to solve for the X. We can put the results together into one equation: 0.50 mol 1.0 mol 4.0 mol 2.7mol $N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$ Based on the moles of NH_3, N_2 is the limiting reactant in this reaction. b) Dimensional analysis. The method shown above is a quicker way of doing stoichiometry calculations but of course you can also use the dimensional analysis approach: Based on N_2, there will be: Based on H_2, there will be: As expected, both approaches lead to the same answer and N_2 is the limiting reactant. We now need to calculate how much H_2 is left over after the reaction is complete and in order to do it, we will first calculate how much of it is consumed in the reaction based on the moles of the limiting reactant. This is how much H_2 reacted when the limiting reactant (N_2) was consumed completely. To find how much is left, we need to subtract the initial and consumed amounts: n (H_2) = 4 mol (initial) – 1.5 mol (reacted) = 2.5 mol H_2 left over mass H_2 = 2.5 mol x 2.0 g/mol = 5.0 g 4. When 140.0 g of $AgNO_3$ was added to an aqueous solution of NaCl, 86.0 g of AgCl was collected as a white precipitate. Which salt was the limiting reactant in this reaction? How many grams of NaCl were present in the solution? $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$ NaCl was the limiting reactant and 35.0 g of it was present in the solution when $AgNO_3$ was added. We need to find the moles of AgCl which will allow us to calculate how much $AgNO_3$ and NaCl had reacted in the reaction.



Limiting Reactant Worksheet #2

1. Given the following reaction: (Balance the equation first!)

   C_3H_8 + O_2 ------> CO_2 + H_2O

   a) If you start with 14.8 g of C_3H_8 and 3.44 g of O_2, determine the limiting reactant.
   b) Determine the number of moles of carbon dioxide produced.
   c) Determine the number of grams of H_2O produced.
   d) Determine the number of grams of excess reactant left.

2. Given the following equation:

   $Al_2(SO_3)_3$ + 6NaOH ------> $3Na_2SO_3$ + $2Al(OH)_3$

   a) If 10.0 g of $Al_2(SO_3)_3$ is reacted with 10.0 g of NaOH, determine the limiting reactant.
   b) Determine the number of moles of $Al(OH)_3$ produced.
   c) Determine the number of grams of Na_2SO_3 produced.
   d) Determine the number of grams of excess reactant left over in the reaction.

3. Given the following equation:

   Al_2O_3 + Fe ------> Fe_3O_4 + Al

   a) If 25.4 g of Al_2O_3 is reacted with 10.2 g of Fe, determine the limiting reactant.
   b) Determine the number of moles of Al produced.
   c) Determine the number of grams of Fe_3O_4 produced.
   d) Determine the number of grams of excess reactant left over in the reaction.

4. Consider the reaction

   $I_2O_5(g)$ + 5 CO(g) ------> 5 $CO_2(g)$ + $I_2(g)$

   a) 80.0 grams of I_2O_5 reacts with 28.0 grams of carbon monoxide, CO. Determine the mass of iodine I_2, which could be produced?
   b) If, in the above situation, only 0.160 moles, of iodine, I_2 was produced.
      i) What mass of iodine was produced?
      ii) What percentage yield of iodine was produced?

5. Zinc and sulfur react to form zinc sulfide according to the equation

   Zn + S ------> ZnS

   If 25.0 g of zinc and 30.0 g of sulfur are mixed,
   a) Which chemical is the limiting reactant?
   b) How many grams of ZnS will be formed?
   c) How many grams of the excess reactant will remain after the reaction is over?

6. Which element is in excess when 3.00 grams of Mg is ignited in 2.20 grams of pure oxygen? What mass is in excess? What mass of MgO is formed?

7. How many grams of Al_2S_3 are formed when 5.00 grams of Al is heated with 10.0 grams S?

$$n(AgCl) = 86.0\ g \times \frac{1\ mol}{143.3\ g} = 0.600\ mol$$

The stoichiometric ratio of all the reactants and products is1:1, and therefore, there were 0.600 mol of $AgNO_3$ and 0.600 mol of NaCl participating in the reaction. Now, let's calculate the moles of 140.0 g of $AgNO_3$ and see whether it corresponds to 0.600 mol. If it is more than that, then $AgNO_3$ was added in excess and NaCl must be the limiting reactant:

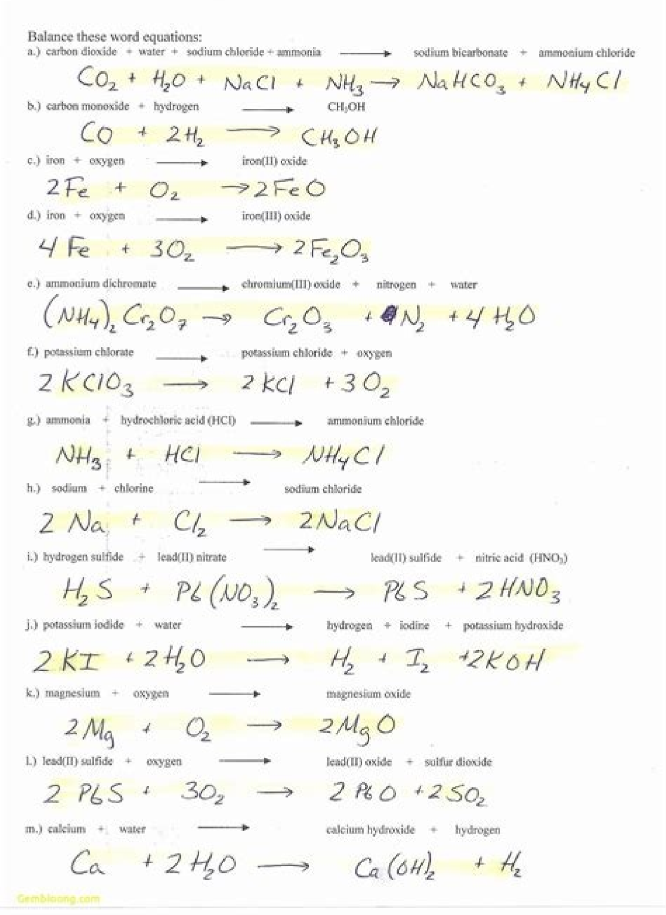
$$n(AgNO_3) = 140.0\ g \times \frac{1\ mol}{169.9\ g} = 0.824\ mol$$

So, 0.824 mol of $AgNO_3$ was added to the solution, but only 0.600 mol had reacted which means that is how much NaCl was present in the solution, and it was not enough to react with all the $AgNO_3$.

And this confirms that $AgNO_3$ was added in excess and NaCl is the limiting reactant. In the last step, we determine the mass of 0.600 mol NaCl that was in the solution when $AgNO_3$ was added:

$$m(NaCl) = 0.600\ mol \times \frac{58.4\ g}{1\ mol} = 35.0\ g$$

5. How many grams of precipitate will be formed when 37.3 g KCl is reacted with 132.5 g $Pb(NO_3)_2$ according to the following chemical equation? $2\ KCl(aq) + Pb(NO_3)_2(aq) \rightarrow PbCl_2(s) + 2\ KNO_3(aq)$ In order to find the amount of products formed or reactants consumed in any chemical equation, we need to first calculate the moles from the corresponding masses. There are exceptions where the calculations can be done based on the volumes but we will not discuss them here. Moles can be calculated from the following formula: Second, we need to identify the precipitate and a solubility table will tell us that $PbCl_2$ is the precipitate which is also indicated by the (s – solid) notation in the chemical equation. At this point we have the moles of both reactants, but in order to determine how much product is formed, we need to always use the limiting reactant.


Balance these word equations:

a.) carbon dioxide + water + sodium chloride + ammonia ⟶ sodium bicarbonate + ammonium chloride

$CO_2 + H_2O + NaCl + NH_3 \rightarrow NaHCO_3 + NH_4Cl$

b.) carbon monoxide + hydrogen ⟶ CH_3OH

$CO + 2H_2 \rightarrow CH_3OH$

c.) iron + oxygen ⟶ iron(II) oxide

$2Fe + O_2 \rightarrow 2FeO$

d.) iron + oxygen ⟶ iron(III) oxide

$4Fe + 3O_2 \rightarrow 2Fe_2O_3$

e.) ammonium dichromate ⟶ chromium(III) oxide + nitrogen + water

$(NH_4)_2Cr_2O_7 \rightarrow Cr_2O_3 + N_2 + 4H_2O$

f.) potassium chlorate ⟶ potassium chloride + oxygen

$2KClO_3 \rightarrow 2KCl + 3O_2$

g.) ammonia + hydrochloric acid (HCl) ⟶ ammonium chloride

$NH_3 + HCl \rightarrow NH_4Cl$

h.) sodium + chlorine ⟶ sodium chloride

$2Na + Cl_2 \rightarrow 2NaCl$

i.) hydrogen sulfide + lead(II) nitrate ⟶ lead(II) sulfide + nitric acid (HNO_3)

$H_2S + Pb(NO_3)_2 \rightarrow PbS + 2HNO_3$

j.) potassium iodide + water ⟶ hydrogen + iodine + potassium hydroxide

$2KI + 2H_2O \rightarrow H_2 + I_2 + 2KOH$

k.) magnesium + oxygen ⟶ magnesium oxide

$2Mg + O_2 \rightarrow 2MgO$

l.) lead(II) sulfide + oxygen ⟶ lead(II) oxide + sulfur dioxide

$2PbS + 3O_2 \rightarrow 2PbO + 2SO_2$

m.) calcium + water ⟶ calcium hydroxide + hydrogen

$Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$

Let’s determine independently how much PbCl2 could be formed from each reactant. Based on KCl: Based on Pb(NO3)2: The limiting reactant in this reaction is KCl and based on the mole ratio, 0.25 mole of PbCl2 will be formed. The last step is to convert the moles to grams: From the formula we can use cross multiplication and find that the mass is equal to: m = n x M so, m (PbCl2) = 0.250 mol x 278.1 g/mol = 69.5 g PbCl2 Alternatively, you can use dimensional analysis to calculate the mass: 6. How many grams of PCl3 will be produced if 38.5 g Cl2 is reacted with 56.4 g P4 according to the following equation? 6Cl2(g) + P4(s) → 4PCl3(l) When the amounts of both reactants are given, you must determine the LR and then, based on the moles of the LR, determine the amount of product(s) that can be formed. The moles of Cl2 and P4 are: \[{\rm{n}}\;{\rm{(C}}{{\rm{l}}_{\rm{2}}}{\rm{)}}\;{\rm{ = }}\;{\rm{38}}{\rm{.5}}\;{\rm{g}}\;{\rm{ \times }}\;\frac{{{\rm{1}}\;{\rm{mol}}}}{{{\rm{70}}{\rm{.9}}\;{\rm{g}}}}\;{\rm{ = }}\;{\rm{1}}{\rm{.84}}\;{\rm{mol}}\] \[{\rm{n}}\;{\rm{(}}{{\rm{P}}_{\rm{4}}}{\rm{)}}\;{\rm{ = }}\;{\rm{56}}{\rm{.4}}\;{\rm{g}}\;{\rm{ \times }}\;\frac{{{\rm{1}}\;{\rm{mol}}}}{{{\rm{123}}{\rm{.9}}\;{\rm{g}}}}\;{\rm{ = }}\;{\rm{0}}{\rm{.455}}\;{\rm{mol}}\] Next, we calculate how much PCl3 can be formed from each reactant: \[{\rm{n}}\;\left({{\rm{PC}}{{\rm{l}}_{\rm{3}}}} \right)\;{\rm{ = }}\;{\rm{1}}{\rm{.84 }}\cancel{{{\rm{mol}}\;{\rm{C}}{{\rm{l}}_{\rm{2}}}}}\;{\rm{ \times }}\;\frac{{{\rm{4}}\;{\rm{mol}}\;{\rm{PC}}{{\rm{l}}_{\rm{3}}}}}{{{\rm{6}}\;\cancel{{{\rm{mol}}\;{\rm{C}}{{\rm{l}}_{\rm{2}}}}}}}\;{\rm{ = }}\;{\rm{1}}{\rm{.23}}\;{\rm{mol}}\] \[{\rm{n}}\;\left({{\rm{PC}}{{\rm{l}}_{\rm{3}}}} \right)\;{\rm{ = }}\;{\rm{0}}{\rm{.455 }}\cancel{{{\rm{mol}}\;{{\rm{P}}_{\rm{4}}}}}\;{\rm{ \times }}\;\frac{{{\rm{4}}\;{\rm{mol}}\;{\rm{PC}}{{\rm{l}}_{\rm{3}}}}}{{{\rm{1}}\;\cancel{{{\rm{mol}}\;{{\rm{P}}_{\rm{4}}}}}}}\;{\rm{ = }}\;{\rm{1}}{\rm{.82}}\;{\rm{mol}}\] Cl2 gives less PCl3 and therefore, it is the LR and 1.23 mol of PCl3 can be produced in this reaction. The mass of PCl3 is: \[{\rm{m}}\;{\rm{(PC}}{{\rm{l}}_{\rm{3}}}{\rm{)}}\;{\rm{ = }}\;{\rm{1}}{\rm{.23}}\;{\rm{mol}}\;{\rm{ \times }}\;\frac{{{\rm{137}}{\rm{.3}}\;{\rm{g}}}}{{{\rm{1}}\;{\rm{mol}}}}\;{\rm{ = }}\;{\rm{169}}\;{\rm{g}}\] 7. How many grams of sulfur can be obtained if 12.6 g H2S is reacted with 14.6 g SO2 according to the following equation?

2H2S(g) + SO2(g) → 3S(s) + 2H2O(g) When the amount of both reactants is given, you must determine the LR and then, based on the moles of the LR, determine the amount of product(s) that can be formed. The moles of H2S and SO2 are: \[{\rm{n}}\;{\rm{(}}{{\rm{H}}_{\rm{2}}}{\rm{S)}}\;{\rm{ = }}\;{\rm{12}}{\rm{.6}}\;{\rm{g}}\;{\rm{ \times }}\;\frac{{{\rm{1}}\;{\rm{mol}}}}{{{\rm{34}}{\rm{.1}}\;{\rm{g}}}}\;{\rm{ = }}\;{\rm{0}}{\rm{.370}}\;{\rm{mol}}\] \[{\rm{n}}\;{\rm{(S}}{{\rm{O}}_{\rm{2}}}{\rm{)}}\;{\rm{ = }}\;{\rm{14}}{\rm{.6}}\;{\rm{g}}\;{\rm{ \times }}\;\frac{{{\rm{1}}\;{\rm{mol}}}}{{{\rm{64}}{\rm{.1}}\;{\rm{g}}}}\;{\rm{ = }}\;{\rm{0}}{\rm{.228}}\;{\rm{mol}}\] Next, we calculate how much PCl3 can be formed from each reactant: \[{\rm{n}}\;\left({\rm{S}} \right)\;{\rm{ = }}\;{\rm{0}}{\rm{.370 }}\cancel{{{\rm{mol}}\;{{\rm{H}}_{\rm{2}}}{\rm{S}}}}\;{\rm{ \times }}\;\frac{{{\rm{3}}\;{\rm{mol}}\;{\rm{S}}}}{{{\rm{2}}\;\cancel{{{\rm{mol}}\;{{\rm{H}}_{\rm{2}}}{\rm{S}}}}}}\;{\rm{ = }}\;{\rm{0}}{\rm{.555}}\;{\rm{mol}}\] \[{\rm{n}}\;\left({\rm{S}} \right)\;{\rm{ = }}\;{\rm{0}}{\rm{.228 }}\cancel{{{\rm{mol}}\;{\rm{S}}{{\rm{O}}_{\rm{2}}}}}\;{\rm{ \times }}\;\frac{{{\rm{3}}\;{\rm{mol}}\;{\rm{S}}}}{{{\rm{1}}\;\cancel{{{\rm{mol}}\;{\rm{S}}{{\rm{O}}_{\rm{2}}}}}}}\;{\rm{ = }}\;{\rm{0}}{\rm{.684}}\;{\rm{mol}}\] H2S gives less S and therefore, it is the LR and 0.555 mol of S can be produced in this reaction. The mass of S is: \[{\rm{m}}\;{\rm{(S)}}\;{\rm{ = }}\;{\rm{0}}{\rm{.555}}\;{\rm{mol}}\;{\rm{ \times }}\;\frac{{{\rm{32}}{\rm{.1}}\;{\rm{g}}}}{{{\rm{1}}\;{\rm{mol}}}}\;{\rm{ = }}\;{\rm{17}}{\rm{.8}}\;{\rm{g}}\] 8. The following equation represents the combustion of octane, C8H18, a component of gasoline: 2C8H18(g) + 25O2(g) → 16CO2(g) + 18H2O(g) Will 356 g of oxygen be enough for the complete combustion of 954 g of octane? 356 g of oxygen is not enough for the complete combustion of 954 g of octane. First, calculate the moles of the reactants: \[{\rm{n}}\;{\rm{(}}{{\rm{O}}_{\rm{2}}}{\rm{)}}\;{\rm{ = }}\;{\rm{356}}\;{\rm{g}}\;{\rm{ \times }}\;\frac{{{\rm{1}}\;{\rm{mol}}}}{{{\rm{32}}{\rm{.0}}\;{\rm{g}}}}\;{\rm{ = }}\;{\rm{11}}{\rm{.4}}\;{\rm{mol}}\] \[{\rm{n}}\;{\rm{(}}{{\rm{C}}_{\rm{8}}}{{\rm{H}}_{{\rm{18}}}}{\rm{)}}\;{\rm{ = }}\;{\rm{954}}\;{\rm{g}}\;{\rm{ \times }}\;\frac{{{\rm{1}}\;{\rm{mol}}}}{{{\rm{114}}{\rm{.2}}\;{\rm{g}}}}\;{\rm{ = }}\;{\rm{8}}{\rm{.35}}\;{\rm{mol}}\] And now let’s see how many moles of O2 are needed to react with 8.35 mol of C8H18: \[{\rm{n}}\;\left({{{\rm{O}}_{\rm{2}}}} \right)\;{\rm{ = }}\;{\rm{8}}{\rm{.35 }}\cancel{{{\rm{mol}}\;{{\rm{C}}_{\rm{8}}}{{\rm{H}}_{{\rm{18}}}}}}\;{\rm{ \times }}\;\frac{{{\rm{25}}\;{\rm{mol}}\;{{\rm{O}}_{\rm{2}}}}}{{{\rm{2}}\;\cancel{{{\rm{mol}}\;{{\rm{C}}_{\rm{8}}}{{\rm{H}}_{{\rm{18}}}}}}}}\;{\rm{ = }}\;{\rm{104}}\;{\rm{mol}}\] 104 moles of O2 are needed to react with only 8.35 mol of C8H18 and this is because of the 25:2 mole ratio. So, 356 g of oxygen is not enough for the complete combustion of 954 g of octane. 9. Consider the reaction between MnO2 and HCl: MnO2 + 4HCl → MnCl2 + Cl2 + 2H2O What is the theoretical yield of MnCl2 in grams when 165 g of MnO2 is added to a solution containing 94.2 g of HCl? 0.645 mol or 81.1 g MnCl2 The theoretical yield is the amount of product that can be formed, so this problem is similar to the ones we have been working on. The amounts of both reactants are given, so we determine the LR and then, based on the moles of the LR, determine the amount of product(s) that can be formed. The moles of MnO2 and HCl are: \[{\rm{n}}\;{\rm{(Mn}}{{\rm{O}}_{\rm{2}}}{\rm{)}}\;{\rm{ = }}\;{\rm{165}}\;{\rm{g}}\;{\rm{ \times }}\;\frac{{{\rm{1}}\;{\rm{mol}}}}{{{\rm{86}}{\rm{.9}}\;{\rm{g}}}}\;{\rm{ = }}\;{\rm{1}}{\rm{.90}}\;{\rm{mol}}\] \[{\rm{n}}\;{\rm{(HCl)}}\;{\rm{ = }}\;{\rm{94}}{\rm{.2}}\;{\rm{g}}\;{\rm{ \times }}\;\frac{{{\rm{1}}\;{\rm{mol}}}}{{{\rm{36}}{\rm{.5}}\;{\rm{g}}}}\;{\rm{ = }}\;{\rm{2}}{\rm{.58}}\;{\rm{mol}}\] Next, we calculate how much MnCl2 can be formed from each reactant: \[{\rm{n}}\;\left({{\rm{MnC}}{{\rm{l}}_{\rm{2}}}} \right)\;{\rm{ = }}\;{\rm{1}}{\rm{.90 }}\cancel{{{\rm{mol}}\;{\rm{Mn}}{{\rm{O}}_{\rm{2}}}}}\;{\rm{ \times }}\;\frac{{{\rm{1}}\;{\rm{mol}}\;{\rm{MnC}}{{\rm{l}}_{\rm{2}}}}}{{{\rm{1}}\;\cancel{{{\rm{mol}}\;{\rm{Mn}}{{\rm{O}}_{\rm{2}}}}}}}\;{\rm{ = }}\;{\rm{1}}{\rm{.90}}\;{\rm{mol}}\] \[{\rm{n}}\;\left({{\rm{MnC}}{{\rm{l}}_{\rm{2}}}} \right)\;{\rm{ = }}\;{\rm{2}}{\rm{.58 }}\cancel{{{\rm{mol}}\;{\rm{HCl}}}}\;{\rm{ \times }}\;\frac{{{\rm{1}}\;{\rm{mol}}\;{\rm{MnC}}{{\rm{l}}_{\rm{2}}}}}{{{\rm{4}}\;\cancel{{{\rm{mol}}\;{\rm{HCl}}}}}}\;{\rm{ = }}\;{\rm{0}}{\rm{.645}}\;{\rm{mol}}\] HCl gives less product so, it is the LR, and therefore, 0.645 mol MnCl2 is the theoretical yield of the reaction. If you are asked to find the theoretical yield in grams, then convert the moles to grams using the molar mass of MnCl2: \[{\rm{m}}\;{\rm{(MnC}}{{\rm{l}}_{\rm{2}}}{\rm{)}}\;{\rm{ = }}\;{\rm{0}}{\rm{.645}}\;{\rm{mol}}\;{\rm{ \times }}\;\frac{{{\rm{125}}{\rm{.8}}\;{\rm{g}}}}{{{\rm{1}}\;{\rm{mol}}}}\;{\rm{ = }}\;{\rm{81}}{\rm{.1}}\;{\rm{g}}\]