# StoichiometryLimiting Reagent Examples

chemteam.info/Stoichiometry/Limiting-Reagent.html

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|---|---|
| Limiting Reagent Problems #1-10 | Problems solved using dimensional analysis only |
| Limiting Reagent Problems #11-20 | Stoichiometry Menu |

**First comment before starting:**

Just a bit below, I'm going to tell you (several times) how to determine the limiting reagent in a chemistry problem. I certainly hope it is something you pay attention to and remember. Figuring out which substance is the limiting reagent is an area that many students struggle with.

You will see the word "excess" used in this section and in the problems. It is used several different ways:

> (a) **Compound A reacts with an excess of compound B.** In this case, mentally set compound B aside for the moment. Since it is "in excess," this means there is more than enough of it. The other compound will run out first.
> (b) **20 grams of A and 20 grams of B react. Which is in excess?** What we will do below is find out which substance runs out first (called the limiting reagent). Obviously (I hope), the other compound is seen to be in excess.
>
> (c) **After 20 gm. of A and 20 gm. of B react, how much of the excess compound remains.** To answer this problem, a subtraction will be involved. This is a part of many limiting reagent problems and it causes difficult with students. Expect it to be on your test.

**Second comment before starting: What is the Limiting Reagent?**

It is simply the substance in a chemical reaction that runs out first. It seems to be a simple concept, but it does cause people problems. Let's try a simple non-chemical example.

Reactant A is a test tube. I have 20 of them.
Reactant B is a stopper. I have 30 of them.

Product C is a stoppered test tube.

The reaction is:

> A + B ---> C
> test tube plus stopper gives stoppered test tube.

So now we let them "react." The first stopper goes in, the second goes in and so on. Step by step we use up stoppers and test tubes (the amounts go down) and make stoppered test tubes (the amount goes up).

Suddenly, we run out of one of the "reactants." Which one? We run out of test tubes first. Seems obvious, doesn't it? We had 20 test tubes, but we had 30 stoppers. So when the test tubes are used up, we have 10 stoppers sitting there unused. And we also have 20 test tubes with stoppers firmly inserted.

So, which "reactant" is limiting and which is in excess? The test tubes are limiting (they ran out first) and the stoppers are in excess (we have some left over when the limiting reagent ran out).

There are two techniques for determine the limiting reagent in chemical problems. The first technique is discussed as part of the solution to the first example. Make sure you take a close look at it. The second technique will make its first appearance in Example #6.

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**Example #1:** Here's a nice limiting reagent problem we will use for discussion. Consider the reaction:

> $2Al + 3I_2$ ------> $2AlI_3$

Determine the limiting reagent and the theoretical yield of the product if one starts with:

> (a) 1.20 mol Al and 2.40 mol iodine.
> (b) 1.20 g Al and 2.40 g iodine
> (c) How many grams of Al are left over in part b?

**Solution for part (a):**

We already have moles as the unit, so we use those numbers directly.

1) Here is how to find out the limiting reagent:

> take the moles of each substance and divide it by its coefficient in the balanced equation. The substance that has the smallest answer is the limiting reagent.

2) Let's say that again:

> **to find the limiting reagent, take the moles of each substance and divide it by its coefficient in the balanced equation. The substance that has the smallest answer is the limiting reagent.**
> You're going to need that technique, so remember it.
>
> By the way, did you notice that I bolded the technique to find the limiting reagent? I did this so as to emphasize its importance to you when learning how to do limiting reagent problems.

3) Resuming with the problem solution:

> Aluminum ---> 1.20 / 2 = 0.60
> Iodine ---> 2.40 / 3 = 0.80

4) The lowest number indicates the limiting reagent. Aluminum will run out first in part (a) of the question. Why?

> 1.20/2 means there are 0.60 "groupings" of 2 and 2.40/3 means there are 0.80 "groupings" of 3. If they ran out at the same time, we'd need one "grouping" of each. Since there is less of the "grouping of 2," it will run out first.
> If you're not sure what I just said, that's OK. The technique works, so remember it and use it.

5) The second part of the question "theoretical yield" depends on finding out the limiting reagent. Once we do that, it becomes a stoichiometric calculation.

> Al and $AlI_3$ stand in a one-to-one molar relationship, so 1.20 mol of Al produces 1.20 mol of $AlI_3$. Notice that the amount of $I_2$ does not play a role, since it is in excess.

**Solution for part (b):**

1) Since we have grams, we must first convert to moles. The we solve just as we did in part (a) just above. For the mole calculation:

> Aluminum ---> 1.20 g / 26.98 g mol$^{-1}$ = 0.04477 mol
> Iodine ---> 2.4 g / 253.8 g mol$^{-1}$ = 0.009456 mol

2) To determine the limiting reagent:

> Aluminum ---> 0.04477 / 2 = 0.02238
> Iodine ---> 0.009456 / 3 = 0.003152
> The lower number is iodine, so we have identified the limiting reagent.

3) Finally, we have to do a calculation and it will involve the iodine, NOT the aluminum.

> $I_2$ and $AlI_3$ stand in a three-to-two molar relationship, so 0.009456 mol of $I_2$ produces 0.006304 mol of $AlI_3$. Again, notice that the amount of Al does not play a role, since it is in excess.
> From here figure out the grams of $AlI_3$ and you have your answer.

**Solution for part (c):**

> Since we have moles, we calculate directly and then convert to grams.
> Al and $I_2$ stand in a two-to-three molar relationship, so 0.009456 mol of $I_2$ uses 0.006304 mol of Al.
>
> Convert this aluminum amount to grams and subtract it from 1.20 g and that's the answer.

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Just above was some discussion on a way to determine the limiting reagent in a chemistry problem. This particular thing (determine the limiting reagent) is a real stumbling block for students. Be aware!

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**Example #2:** 15.00 g aluminum sulfide and 10.00 g water react until the limiting reagent is used up. Here is the balanced equation for the reaction:

> $Al_2S_3 + 6H_2O \text{ ---> } 2Al(OH)_3 + 3H_2S$
> (a) Which is the limiting reagent?
> (b) What is the maximum mass of $H_2S$ which can be formed from these reagents?
> (c) How much excess reagent remains after the reaction is complete?

Some comments first:

> The key to this problem is the limiting reagent, part (a). Once you know that, part (b) becomes "How much $H_2S$ can be made from the limiting reagent?" Part (c) becomes two connected questions: first, "How much $Al_2S_3$ is used up when reacting with the limiting reagent?" then second, "What is 15.00 minus the amount in the first part?"
> Make sure you note that second part. The calculation o be performed gives you the answer to "How much reacted?" but the question is "How much remained?" Lots of students forget to do the second part (the 15 minus part) and so get graded down.
>
> Note: I'm carrying a guard digit or two through the calculations. The final answers will appear with the proper number of significant figures.

**Solution for limiting reagent, part (a):**

1) Determine the moles of $Al_2S_3$ and $H_2O$

> Aluminum sulfide ---> 15.00 g ÷ 150.158 g/mol = 0.099895 mol
> Water ---> 10.00 g ÷ 18.015 g/mol = 0.555093 mol

2) Divide each mole amount by equation coefficient:

> Aluminum sulfide ---> 0.099895 mol ÷ 1 mol = 0.099895
> Water --> 0.555093 mol ÷ 6 mol = 0.0925155

3) The water is the lesser amount; it is the limiting reagent.

**Solution for mass of $H_2S$ formed, part (b)**

Now that we know the limiting reagent is water, this problem becomes "How much $H_2S$ is produced from 10.00 g of $H_2O$ and excess aluminum sulfide?"

1) Determine moles of 10.00 g of $H_2O$

> Water ---> 10.00 g ÷ 18.015 g/mol = 0.555093 mol

2) Use molar ratios to determine moles of $H_2S$ produced from above amount of water.

> (a) the $H_2O/H_2S$ molar ratio is 6/3, a 2/1 ratio.
> (b) water is associated with the two. This means the $H_2S$ amount is one-half the water value = 0.2775465 mol.

3) Convert moles of $H_2S$ to grams.

> (0.2775465 mol) (34.0809 g/mol) = 9.459 g

**Solution for excess reagent remaining, part (c)**

We will use the amount of water to calculate how much $Al_2S_3$ reacts, then subtract that amount from 15.00 g.

1) Determine moles of 10.00 g of $H_2O$

> Water ---> 10.00 g ÷ 18.015 g/mol = 0.555093 mol

2) Use molar ratios to determine moles of $Al_2S_3$ that reacts with the above amount of water.

> (a) the $Al_2S_3/H_2O$ ratio is 1/6
> (b) water is associated with the 6. This means the $Al_2S_3$ amount is one-sixth the water value = 0.09251447 mol

3) Convert moles of $Al_2S_3$ to grams.

> (0.09251447 mol) (150.158 g/mol) = 13.891943 g

4) However, we are not done. We were asked for the amount remaining and the answer just above is the amount which was used up, so the final step is:

> 15.00 g − 13.891943 g = 1.108 g

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**Example #3:** If there is 35.0 grams of $C_6H_{10}$ and 45.0 grams of $O_2$, how many grams of the excess reagent will remain after the reaction ceases?

> $2C_6H_{10} + 17O_2 \text{ ---> } 12CO_2 + 10H_2O$

**Solution:**

1) Convert each substance to moles:

> $C_6H_{10}$ ---> 35.0 g / 82.145 g/mol = 0.426 mol
> $O_2$ ---> 45.0 g / 31.998 g/mol = 1.406 mol

2) Determine the limiting reagent:

> $C_6H_{10}$ ---> 0.426 mol / 2 = 0.213
> $O_2$ ---> 1.406 mol / 17 = 0.083
> $O_2$ is the limiting reagent.
>
> Comment: the units don't matter in this step. What we are looking for is the smallest number after carrying out the divisions. The value of 0.083 is the important thing. Not if it has a unit attached to it or not.

3) Determine how many moles of the excess reagent is used up when the limiting reagent is fully consumed:

> the mole ratio we desire is 2/17 ($C_6H_{10}$ to $O_2$)
>
> $$\frac{2}{17} = \frac{x}{1.406 \text{ mol}}$$
>
> x = 0.1654 mol of $C_6H_{10}$ consumed

4) Determine grams of $C_6H_{10}$ remaining:

> 0.426 mol − 0.1654 mol = 0.2606 mol of $C_6H_{10}$ remaining
> (0.2606 mol) (82.145 g/mol) = 21.4 g remaining (to three sig figs)

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**Example #4:** (a) What mass of $Al_2O_3$ can be produced from the reaction of 10.0 g of Al and 19.0 g of $O_3$? (b) How much of the excess reagent remains unreacted?

**Solution to a:**

1) Write balanced chemical equation:

> $2Al + O_3 \text{ ---> } Al_2O_3$

2) Convert grams to moles:

> Al ---> 10.0 g / 26.982 g/mol = 0.37062 mol
> $O_3$ ---> 19.0 g / 47.997 g/mol = 0.39586 mol

3) Determine limiting reagent:

Al ---> 0.37062 / 2 = 0.18531
$O_3$ ---> 0.39586 / 1 = 0.39586
Al is the limiting reagent

4) Determine moles of product formed:

Al to $Al_2O_3$ molar ratio is 2 to 1.

$$\frac{2}{1} = \frac{0.37062 \text{ mol}}{x}$$

x = 0.18531 mol

5) Determine grams of product:

(0.18531 mol) (101.961 g/mol) = 18.8944 g
To three sig figs, 18.9 g

**Solution to b:**

1) Determine moles of ozone that reacted:

Al to $O_3$ molar ratio is 2 to 1

$$\frac{2}{1} = \frac{0.37062 \text{ mol}}{x}$$

x = 0.18531 mol

2) Determine moles of ozone remaining:

0.39586 mol − 0.18531 mol = 0.21055 mol

3) Determine grams of ozone remaining:

(0.21055 mol) (47.997 g/mol) = 10.1 g (to three sig figs)

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**Example #5:** Based on the balanced equation:

$C_4H_8 + 6O_2 \text{ ---> } 4CO_2 + 4H_2O$

Calculate the number of excess reagent units remaining when 28 $C_4H_8$ molecules and 228 $O_2$ molecules react?

**Solution:**

Remember, numbers of molecules are just like moles, so treating the 28 and 228 as moles is perfectly acceptable. This is because I could divide the 28 and the 228 by Avogadro's Number to obtain the moles. Those mole amounts could be used in the calculation below and the final answer could then be multiplied by Avogadro's Number to obtain the answer of 60.

1) Determine the limiting reagent:

> butane ---> 28 / 1 = 28
> oxygen ---> 228 / 6 = 38
> Butane is the limiting reagent.

2) Determine how much oxygen reacts with 28 $C_4H_8$ molecules:

> the butane to oxygen molar ratio is 1:6
> 28 x 6 = 168 oxygen molecules react

3) Determine excess oxygen:

> 228 − 168 = 60

Here's aother way to consider this:

> The 38 above means that there are 38 "groupings" of six oxygen molecules.
> 38 − 28 = 10 oxygen "groupings" remain after the butane is used up
>
> 10 x 6 = 60

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**Example #6:** Determine the maximum mass of $TiCl_4$ that can be obtained from 35.0 g of $TiO_2$, 45.0 g $Cl_2$ and 11.0 g of C. (See comment below problem.)

> $3TiO_2 + 4C + 6Cl_2 \longrightarrow 3TiCl_4 + 2CO_2 + 2CO$

**Solution:**

1) Assume each reactant is the limiting reagent. Determine the moles of product produced by each assumption:

Note: the first factor in each case converts grams of each reactant to moles. The second factor uses a molar ratio from the chemical equation to convert from moles of the reactant to moles of product. There is no need to convert to grams because all three calculations yield moles of the same compound (the $TiCl_4$).

$$45.0\text{ g } Cl_2 \times \frac{1\text{ mole } Cl_2}{70.9064\text{ g } Cl_2} \times \frac{3\text{ mole } TiCl_4}{6\text{ mol } Cl_2} = 0.31732\text{ mol } TiCl_4$$

$$11.0\text{ g C} \times \frac{1\text{ mole C}}{12.01078\text{ g C}} \times \frac{3\text{ mole } TiCl_4}{4\text{ mol C}} = 0.68688\text{ mol } TiCl_4$$

$$35.0\text{ g } TiO_2 \times \frac{1\text{ mole } TiO_2}{79.8658\text{ g } TiO_2} \times \frac{3\text{ mole } TiCl_4}{3\text{ mol } TiO_2} = 0.438235\text{ mol } TiCl_4$$

> $Cl_2$ makes the least amount of $TiCl_4$, so $Cl_2$ is the limiting reactant.

2) The mass of $TiCl_4$ produced is:

> ($0.31732$ mol $TiCl_4$) ($189.679$ g $TiCl_4$/mol) = 60.2 g $TiCl_4$ (to three sig figs)

Note that the "divide moles by coefficient" was not used to determine the limiting reagent. Instead, a full calculation was done and the least amount of product identified the limiting reagent. Here is what the "divide moles by coefficient" set up looks like:

> $Cl_2$ ---> 0.63464 / 6 = 0.10577 <--- there's our limiting reagent
> C ---> 0.915844 / 4 = 0.228961
> $TiO_2$ ---> 0.438235 / 3 = 0.14608

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**Example #7:** Determine the starting mass of each reactant if 46.3 of $K_3PO_4$ is produced and 92.8 of $H_3PO_4$ remains unreacted.

> $3KOH(aq) + H_3PO_4(aq) \text{ ---> } K_3PO_4(aq) + 3H_2O(\ell)$

**Solution:**

1) The fact that some phosphoric acid remains tells us it is the excess reagent. Let us determine the amount of KOH (the limiting reagent) required to produce the 46.3 g of $K_3PO_4$.

> 46.3 g / 212.264 g/mol = 0.2181246 mol of $K_3PO_4$
> Three moles of KOH are required to produce one mole of $K_3PO_4$
>
> (3) (0.2181246 mol) = 0.6543738 mol of KOH required
>
> (0.6543738 mol) (56.1049 g/mol) = 36.7 g (to thee sig figs)

2) Determine the starting mass of $H_3PO_4$

> 0.2181246 mol of $K_3PO_4$ requires 0.2181246 mol of $H_3PO_4$ based on the 1:1 molar ratio from the balanced equation.
> (0.2181246 mol) (97.9937 g/mol) = 21.4 g (to three sig figs)
>
> 21.4 + 92.8 = 114.2 g

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**Example #8:** Determine the limiting reagent of this reaction:

> $Na_2B_4O_7 + H_2SO_4 + 5H_2O \text{ ---> } 4H_3BO_3 + Na_2SO_4$
> There are 5.00 g of each reactant.

**Solution:**

1) Convert everything into moles, by dividing each 5.00 g by their respective molar masses:

> $Na_2B_4O_7$ ---> 0.02485 mol
> $H_2SO_4$ ---> 0.05097 mol
> $H_2O$ ---> 0.2775 mol

2) Note that there are three reactants. How is the limiting reagent determined when there are three reactants? Answer: determine the limiting reagent between the first two:

> $Na_2B_4O_7$ ---> 0.02485 / 1 = 0.02485
> $H_2SO_4$ ---> 0.05097 / 1 = 0.05097
> $Na_2B_4O_7$ is the limiting reagent when compared to $H_2SO_4$

3) Now, compare the "winner" to the third reagent:

> $Na_2B_4O_7$ ---> 0.02485 / 1 = 0.02485
> $H_2O$ ---> 0.2775 / 5 = 0.0555
> $Na_2B_4O_7$ is the limiting reagent between itself and $H_2O$.
>
> $Na_2B_4O_7$ is the overall limiting reagent in this problem.

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**Example #9:** How much $O_2$ could be produced from 2.45 g of $KO_2$ and 4.44 g of $CO_2$?

> $4KO_2 + 2CO_2 \text{ ---> } 2K_2CO_3 + 3O_2$

**Solution:**

I will do a solution assuming $KO_2$ is the limiting reagent, then I will do a solution assuming $CO_2$ is the limiting reagent. The reactant that produces the lesser amount of oxygen is the limiting reagent and that lesser amount will be the answer to the question.

1) Solution using $KO_2$:

> $2.45 \text{ g} / 71.096 \text{ g/mol} = 0.03446045 \text{ mol}$
>
> $$\frac{4}{3} = \frac{0.03446045 \text{ mol}}{x}$$
>
> $x = 0.02584534 \text{ mol}$
>
> $(0.02584534 \text{ mol}) (31.998 \text{ g/mol}) = 0.827 \text{ g of } O_2$

2) Solution using $CO_2$:

> $4.44 \text{ g} / 44.009 \text{ g/mol} = 0.10088845 \text{ mol}$
>
> $$\frac{2}{3} = \frac{0.10088845 \text{ mol}}{x}$$
>
> $x = 0.151332 \text{ mol}$
>
> $(0.151332 \text{ mol}) (31.998 \text{ g/mol}) = 4.84 \text{ g of } O_2$

3) 0.827 g is the answer.

Note that I could have calculated the mole amounts, used the "divide moles by coefficient" to determine the limiting reagent, and then done just one complete calculation.

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**Example #10:** (a) What mass of hydrogen peroxide should result when 1.45 g of barium peroxide is treated with 25.5 mL of hydrochloric acid solution containing 0.0277 g of HCl per mL? (b) How much of the excess reactant is left?

> $BaO_2(s) + 2HCl(aq) \text{ ---> } H_2O_2(aq) + BaCl_2(aq)$

**Solution:**

Calculate the amount of product using each reactant. The reactant that produces the lesser of the two amounts will tell you the limiting reactant. This solution will use dimensional analysis (also called the unit-factor, or unit-label, method) for the proposed solution.

1) First, determine the mass of HCl that reacts:

> $(0.0277 \text{ g/mL}) (25.5 \text{ mL}) = 0.70635 \text{ g}$

2) The barium peroxide solution:

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| (1.45 g $BaO_2$) | (1 mol $BaO_2$ / 169.3 g $BaO_2$) | (1 mol $H_2O_2$ / 1 mol $BaO_2$) | (34.0 g $H_2O_2$ / 1 mol $H_2O_2$) | = 0.291 g $H_2O_2$ |
| | ↑ convert grams to moles ↑ | ↑ molar ratio ↑ from equation | ↑ convert moles to grams ↑ | |

3) The hydrochloric acid solution:

> (0.70635 g) (1 mol HCl / 36.46 g HCl) (1 mol $H_2O_2$ / 2 mol HCl) (34.0 g $H_2O_2$ / 1 mol $H_2O_2$) = 0.332 g $H_2O_2$

4) Since 0.291 g is less than 0.332 g, the $BaO_2$ is the limiting reactant.

5) The other method to determine the limiting reagent is to divide the moles of each reactant by their respective coefficient in the balanced equation:

> $BaO_2$ ---> 1.45 g / 169.3 g/mol = 0.008565 mol
> HCl ---> 0.70635 g / 36.46 g/mol = 0.01937 mol
> 0.008565 / 1 = 0.008565
> 0.01937 / 2 = 0.009685
>
> $BaO_2$ (the 0.008565) is the lesser amount, so it is the limiting reagent.

6) To solve part (b), we observe that 0.008565 mol of $BaO_2$ was used. Using a 1:2 molar ratio, we can determine the amount of HCl that was used:

$$\frac{1}{2} = \frac{0.008565 \text{ mol}}{x}$$

> x = 0.01713 mol of HCl used up in the reaction

7) Next, we subtract the amount used up from the total amount that was present:

> 0.01937 mol − 0.01713 mol = 0.00224 mol of HCl remains after reaction stops

8) Convert moles to grams:

> (0.00224 mol) (36.46 g/mol) = 0.0817 g (to three sig figs)

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**Bonus Example:** Consider the following reaction at 1.10 atm and 19.0 °C:

> $2NaCl(s) + 2NH_3(g) + CO_2(g) + H_2O(\ell) \text{ ---> } 2NH_4Cl(aq) + Na_2CO_3(s)$

0.218 mol of sodium chloride, 2.55 L of ammonia, 2.00 L of carbon dioxide, and an unlimited amount of water react to form aqueous ammonium chloride and solid sodium bicarbonate. How many moles of ammonium chloride are formed in the reaction?

Comment: this question was asked and answered on a now-defunct "answers" website and the one answer given (besides mine) totally missed the point of the question. The answerer focused on the non-realistic nature of the above chemical equation. However, the point of the question is to determine the limiting reagent and the non-realistic nature of the chemical equation is completely beside the point.

**Solution:**

1) Use PV = nRT to determine moles of ammonia and carbon dioxide:

> ammonia:
>
> > (1.10 atm) (2.55 L) = (n) (0.08206 L atm / mol K) (292 K)
> > n = 0.11706 mol
>
> carbon dioxide:
>
> > (1.10 atm) (2.00 L) = (n) (0.08206 L atm / mol K) (292 K)
> > n = 0.091814 mol

2) Determine the limiting reagent:

> 0.218 / 2 = 0.109
> 0.11706 / 2 = 0.05853
> 0.091814 / 1 = 0.091814
> Ammonia is the limiting reagent.

3) Now, the problem becomes this: 0.11706 moles of ammonia produces how many moles of ammonium chloride?

> The molar ratio between ammonia and ammonium chloride is 1:1.
> 0.11706 moles of ammonia produces 0.117 moles of ammonium chloride (rounded off to three significant figures).
>
> And we are done.

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| | |
|---|---|
| Limiting Reagent Problems #1-10 | Problems solved using dimensional analysis only |
| Limiting Reagent Problems #11-20 | Stoichiometry Menu |