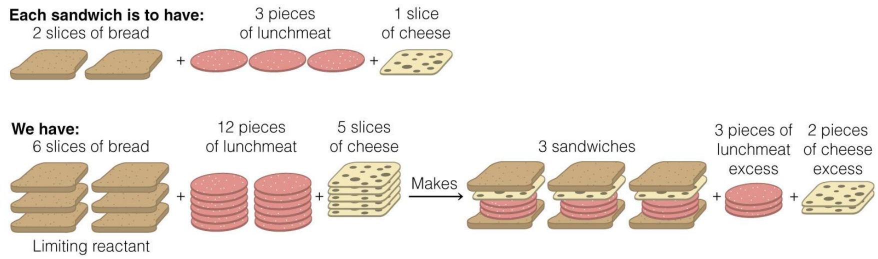
Limiting Reactants

If one reactant is completely used up before any other reactants, it is called the **limiting reactant** because it prevents reaction from going any further; thus, it limits the amount of product made. Any reactant that is not used up is called an **excess reactant.**

Often, a related term, reagent, appears at this point in association with stoichiometry questions. This is an older term which referred to a chemical added to a system in order to bring about a specific chemical reaction, that is, to see if a reaction occurs, thereby proving another specific chemical must be present. Since this was a qualitative test, the purity of a reagent could be lower meaning the active ingredient would run out during a reaction and so it would usually be limiting. This term is now typically used to mean a limiting reactant.



Consider the following example: Based on the recipe of what is required to make a complete sandwich, you can see that you will run out of bread before you run out of lunchmeat or cheese. Thus, the bread prevents you from using up the other ingredients to make more sandwiches. That is, the bread limits production of more sandwiches because it is used up first. The bread is the limiting ingredient or reactant in this situation.

The same thing can happen with chemical equations. For example, silver reacts with nitric acid according to the following reaction: $3\,Ag + 4\,HNO_3 \rightarrow 3\,AgNO_3 + NO + 2\,H_2O$. What mass of nitrogen monoxide will be produced when 100 g of silver reacts with 75 g of nitric acid. This is a limiting reagent question because the amounts of all reactants were given, so it is likely one will be used up. Here is how to handle this type of question:

i) identify substances involved in this question and indicate amounts

$$\underset{100\text{ g}}{\underline{3\,Ag}} + \underset{75\text{ g}}{\underline{4\,HNO_3}} \rightarrow 3\,AgNO_3 + \underset{x\text{ grams}}{\underline{NO}} + 2\,H_2O$$

ii) convert to moles

$$Ag:\ \text{moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{100\text{ g}}{107.9\text{ g/mol}} = 0.927\text{ moles of Ag}$$

$$HNO_3:\ \text{moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{75\text{ g}}{63.0\text{ g/mol}} = 1.19\text{ moles of }HNO_3$$

iii) convert mole ratio of reactants to 1:1 by dividing by their respective numerical coefficients

mole ratio of Ag to HNO_3 is 0.927 mol : 1.19 mol so $\frac{0.927}{3} : \frac{1.19}{4} = 0.309\text{ mol} : 0.298\text{ mol}$

Thus nitric acid (0.298 moles) limits as there is not enough of it to completely react with (0.309 moles) silver. You <u>**must use the limiting reactant**</u> to calculate the product.

iv) write ratio statements using limiting reactant and product

4 moles of HNO_3 produces 1 mole of NO — known statement
1.19 moles of HNO_3 produces x moles of NO — unknown statement

Notice that you must use the real moles of HNO_3 , not the reduced moles (1.19)

v) cross multiply and solve for x

4 moles of HNO_3 produces 1 mole of NO
1.19 moles of HNO_3 produces x moles of NO

(4 moles)(x moles) = (1.19 moles)(1 mole)

Therefore, $x\text{ moles} = \frac{(1.19\text{ moles})(1\text{ mole})}{4\text{ moles}} = 0.298$ moles of nitrogen monoxide (NO)

vi) convert to grams $\quad$ moles = $\frac{\text{mass}}{\text{molar mass}}$ $\quad$ mass = (moles)(molar mass) = (0.298 mol)(30.0 g/mol)
= 8.925 g of nitrogen monoxide

Therefore, 8.925 g of nitrogen monoxide was produced when 75 g of nitric acid reacted with 100 g of silver. Notice that there is a shortcut. You can just use the reduced moles of the limiting reactant and multiply by the coefficient of the desired chemical species. In this case you would just take 0.298 moles of HNO_3 X the coefficient of 1 for NO to get 0.298 moles of NO and convert that to grams. This shortcut works every time, but the danger lies in mixing up the shortcut with the long way, so be careful.

Often, once the limiting reactant has been identified, chemists need to know how much of the excess reactant remains as it can be used again. Consider the reaction, $2\ Al + 3\ I_2 \rightarrow 2\ AlI_3$ as an example. A student mixed 1.2 g of aluminum metal with 2.4 g of solid iodine and 2.25 g of aluminum iodide was collected. Determine the amount of excess reactant left over.

- you need to determine the limiting reactant by finding the limiting moles; first find moles of each reactant

  Al: 1.2 g/(27.0 g/mol) = 0.0444 moles
  I_2: 2.4 g/(253.8 g/mol) = 0.009456 moles

- the next step is to determine the amount of reduced moles by dividing the moles present by the numerical coefficient of the balanced reaction

  Al: 0.0444 moles/2 = 0.0222 moles
  I_2: 0.09456 moles/3 = 0.003152 moles

- since the lower number is the iodine, it is the limiting reactant
- now to determine the amount of excess reactant left over, determine the amount of excess reactant used by using stoich

  2 moles Al reacts with 3 moles of I_2
  x moles of Al reacts with 0.09456 moles of I_2

  x = (2)(0.009456/3 = 0.006304 moles of Al used

- subtract the moles of excess reagent used from the moles originally present

  0.0444 moles of Al originally present – 0.006304 moles of Al used = 0.03814 moles of Al left over

- convert the left over moles to mass

  (0.03814 moles)(27.0 g/mol) = 1.03 g of Al left over

- sometimes you are asked to determine the % excess left over

  (0.03814 moles)/(0.0444 moles) x 100% = 85.82 % of the Al is left over

- or you could also calculate the % excess left over using grams

  1.03/1.2 x 100% = 85.82 % of the Al is left over