

Stoichiometry II

7+8

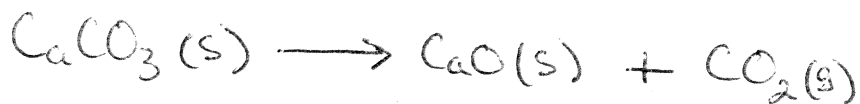
The use of a balanced chemical equation to calculate the quantities of the reactants and products of a reaction.

Abbreviations

1. Reaction = rxn

2. Solution = soln

② A typical chemical equation:



1. Read left to right

2. s, l, g, aq define physical states

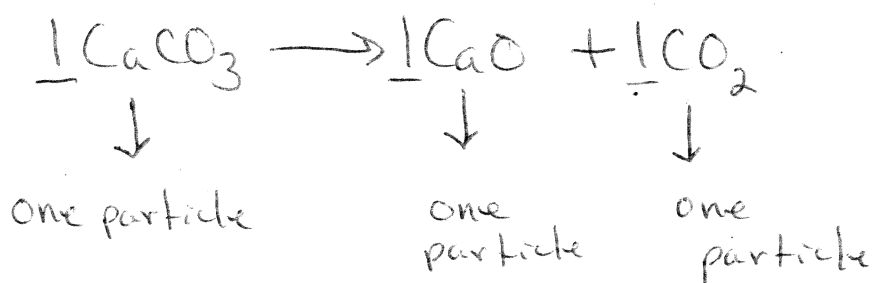
3. proper chemical formulas or elemental symbols

4. Coefficients are the numbers that appear before the formula. There is a 1 in front of CaCO_3 , CaO and CO_2 . Coefficients are in units of moles.

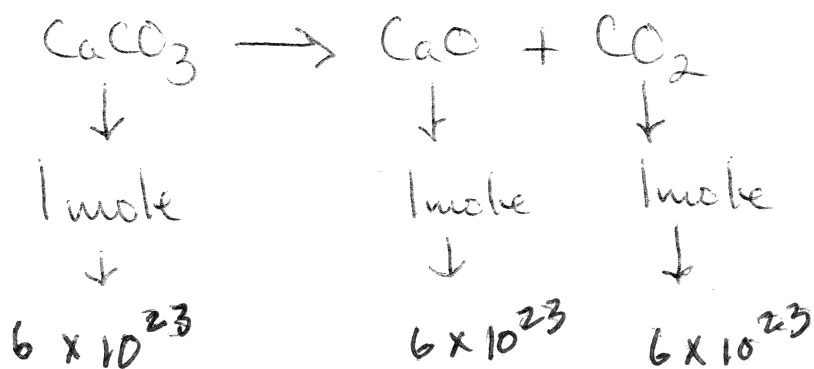
③ Law Conservation of Mass: A balanced equation has the same numbers and types of atoms on both sides of the arrow.

④ Stoichiometry is based on the idea that a balanced equation relates the mole ratio in which reactants combine and the mole ratio in which the products are produced.

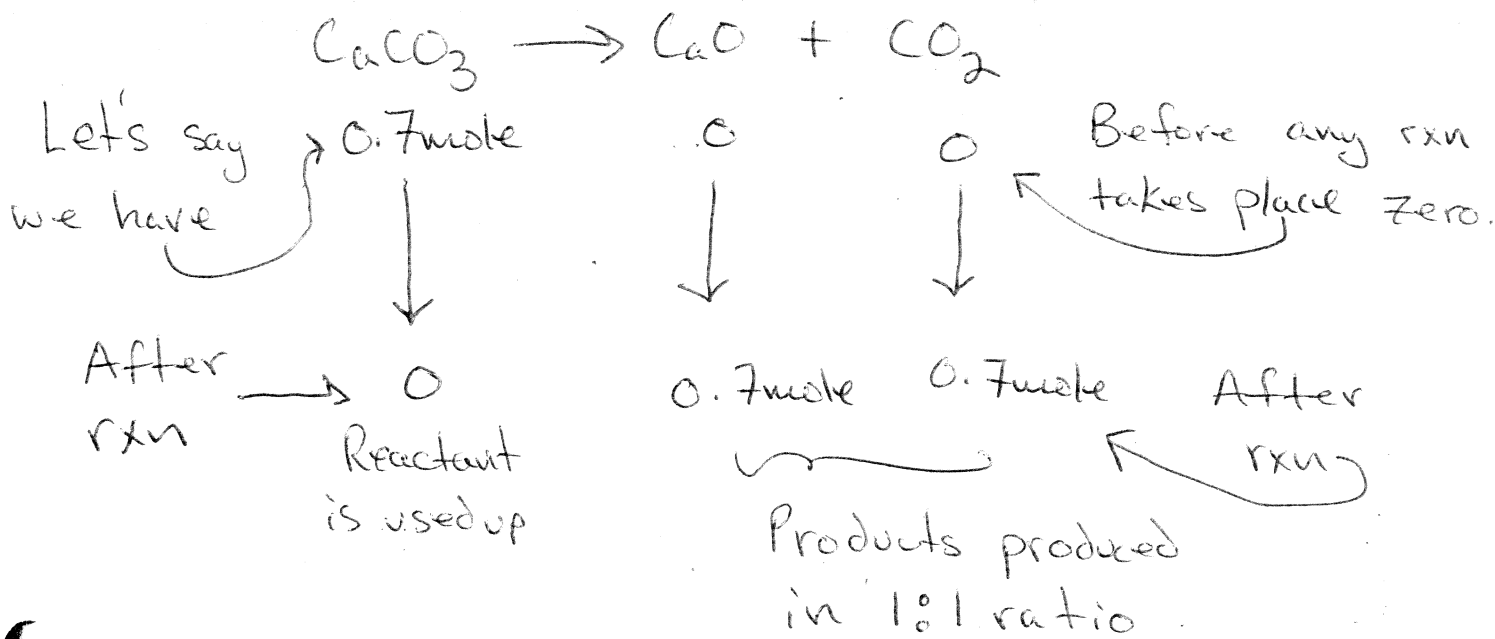
Example



So it takes one particle of CaCO_3 to produce one particle each of CaO and CO_2 .



5) Use the equation to do a basic calculation:



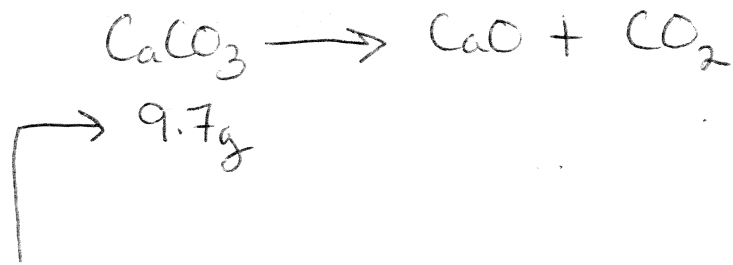
Use unit multiplier method for the same calculation:

$$\frac{0.7 \text{ mole CaCO}_3}{1} \times \frac{1 \text{ mole CaO}}{1 \text{ mole CaCO}_3} = 0.7 \text{ mole CaO}$$

Coefficient from balanced equation

$$\frac{0.7 \text{ mole CaCO}_3}{1} \times \frac{1 \text{ mole CO}_2}{1 \text{ mole CaCO}_3} = 0.7 \text{ mole CO}_2$$

6) Use the equation for a more complicated calculation:



This time we're given mass. Step one is to convert mass to moles.

$$\frac{9.7\text{g CaCO}_3}{1} \times \frac{1\text{mole}}{100\text{g CaCO}_3} = 0.097\text{mole CaCO}_3$$

↖
molecular weight
CaCO₃

Now : ratio is 1:1:1



Before rxn	0.097mole	0	0
	↓	↓	↓
After rxn	0	0.097 mole	0.097 mole

unit multiplier :

↙ coefficient

$$a. 0.097\text{mole CaCO}_3 \times \frac{1\text{mole CaO}}{1\text{mole CaCO}_3} = 0.097\text{mole CaO}$$

↖ coefficient

$$b. 0.097\text{mole CaCO}_3 \times \frac{1\text{mole CO}_2}{1\text{mole CaCO}_3} = 0.097\text{mole CO}_2$$

To find the mass of CaO + CO₂ produced:

$$0.097 \text{ mole CaO} \times \frac{\overset{\text{molecular weight}}{56 \text{ g CaO}}}{1 \text{ mole}} = 5.43 \text{ grams CaO}$$

$$0.097 \text{ moles CO}_2 \times \frac{44 \text{ g CO}_2}{1 \text{ mole}} = 4.27 \text{ g CO}_2$$

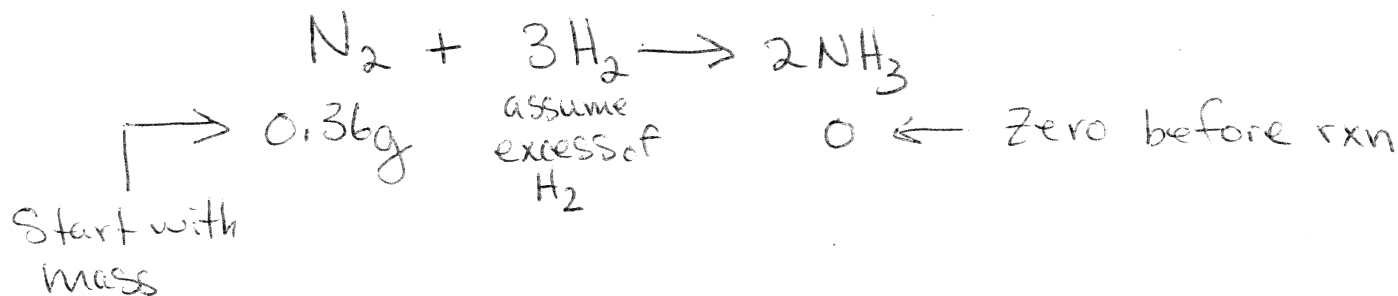
9.7g total mass

Remember this is the mass of CaCO₃ we started with.

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Now lets work with an equation that is not 1:1:1.

1:3:2



Calculate the mass of NH₃ produced:

Step 1: convert mass N₂ to moles N₂

$$0.36 \text{ g N}_2 \times \frac{1 \text{ mole N}_2}{28 \text{ g N}_2} = 0.013 \text{ mole N}_2$$

Step 2: convert moles N₂ to moles NH₃ produced.

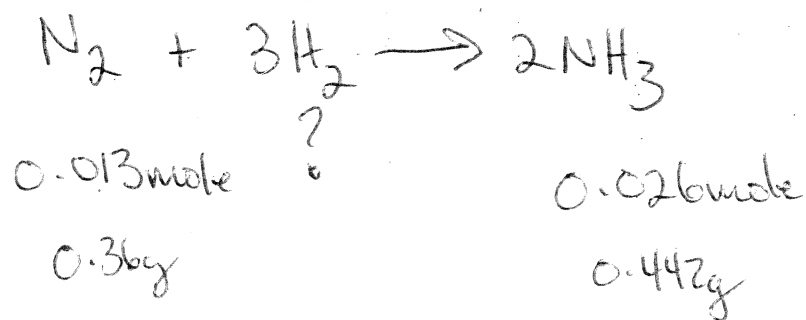
$$0.013 \text{ mole N}_2 \times \frac{2 \text{ mole NH}_3}{1 \text{ mole N}_2} = 0.026 \text{ mole NH}_3$$

Step 3:

Now find mass of NH_3 produced.

$$0.026 \text{ mole } \text{NH}_3 \times \frac{17 \text{ g } \text{NH}_3}{1 \text{ mole}} = 0.442 \text{ g } \text{NH}_3$$

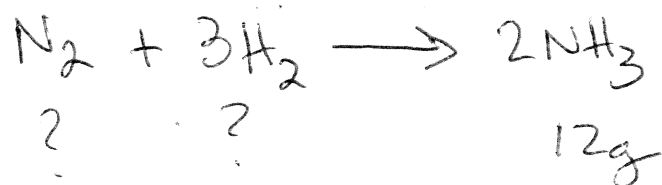
⑧ Now we'll use the same equation to find the mass of H_2 required to react with N_2 .



$$0.013 \text{ mole } \text{N}_2 \times \frac{3 \text{ mole } \text{H}_2}{1 \text{ mole } \text{N}_2} = 0.039 \text{ mole } \text{H}_2$$

$$0.039 \text{ mole } \text{H}_2 \times \frac{2 \text{ g}}{1 \text{ mole}} = 0.078 \text{ g } \text{H}_2$$

9) Similar calculations can be used going in reverse. Let's say 12g NH_3 is produced in a rxn. Calculate mass N_2 and H_2 reacted.



Step 1: convert 12g NH_3 to moles NH_3

$$12\text{g NH}_3 \times \frac{1\text{mole}}{17\text{g NH}_3} = 0.71\text{moles NH}_3$$

Step 2: use equation to find N_2 + H_2

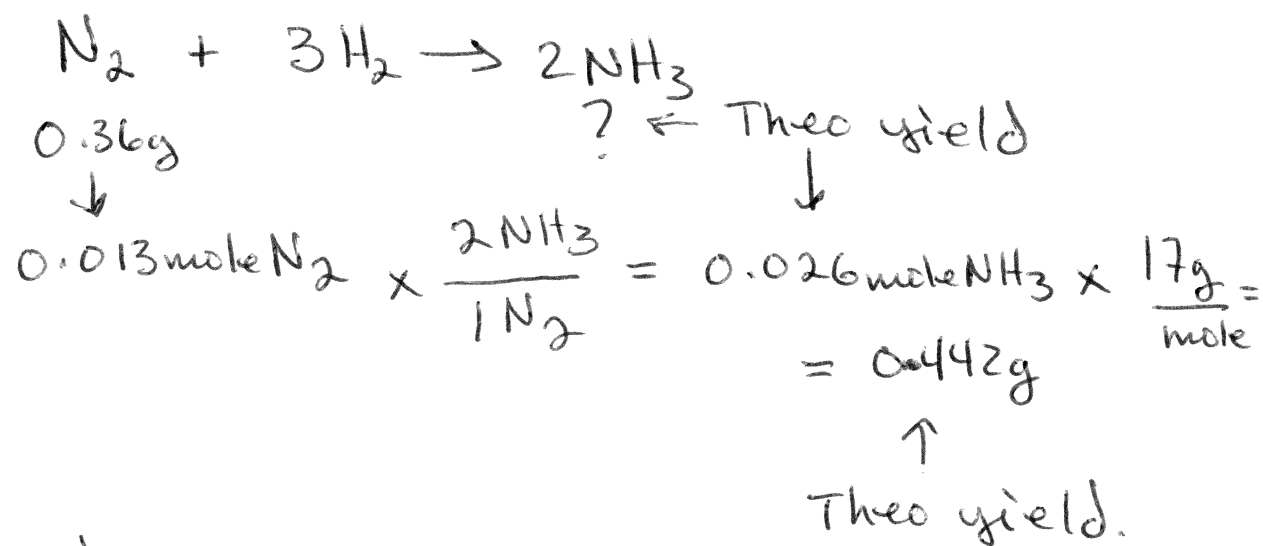
$$0.71\text{moles NH}_3 \times \frac{1\text{mole N}_2}{2\text{mole NH}_3} \times \frac{28\text{g N}_2}{1\text{mole}} = 9.94\text{g N}_2$$

$$0.71\text{moles NH}_3 \times \frac{3\text{mole H}_2}{2\text{mole NH}_3} \times \frac{2\text{g H}_2}{1\text{mole}} = 2.13\text{g H}_2$$

⑩ Theoretical yield

Theoretical yield represents the maximum amount of product in mass or moles that can be formed in a chemical rxn assuming that the rxn went to 100% completion.

In the previous example we determined that when 0.36g of N_2 reacted with an excess of H_2 that 0.442g NH_3 was produced. 0.442g NH_3 is the theoretical yield.



% yield

In order to evaluate a chemical rxn in practice we calculate $\% \text{ yield} = \frac{\text{actual yield}}{\text{Theo yield}} \times 100$

In the last rxn, if the actual yield was 0.39g

$$\text{then the \% yield is: } \% = \frac{0.39}{0.442} \times 100 = 88.2\%$$

The actual yield is the maximum product in mass or moles obtained when the rxn is carried out in the laboratory. The % yield is always smaller than 100% because no rxn is ever perfectly carried out by the chemist.

Limiting Reactant

In any chemical rxn where two or more reactants are combined by providing mass or mole amounts of all reactants at once, one of the reactants will run out bringing a halt to the rxn. The reactant that runs out is the limiting reactant.

Consider the following example: If we combine 100 cars with no tires with 350 tires we cannot produce 100 functioning cars. This is because for 100 cars we need 400 tires.

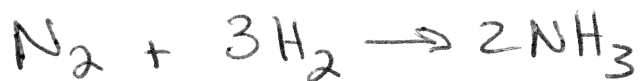
$$\text{For 100 cars: } 100 \text{ cars} \times \frac{4 \text{ tires}}{1 \text{ car}} = 400 \text{ tires needed}$$

$$350 \text{ tires} \times \frac{1 \text{ car}}{4 \text{ tires}} = 87.5 \text{ cars or } 87 \text{ complete cars.}$$

The bottom line is that 13 cars will be left uncompleted. The tires are the limiting reactant.

Now let's consider a chemical example:

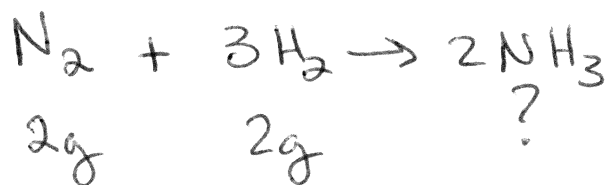
2g of N_2 is combined with 2g H_2 to produce ammonia (NH_3):



The goal:

1. calculate the limiting reactant
2. calculate the yield of NH_3
3. calculate mass of the remaining reactant.

(A) First re-write information



(B) Calculate the yield of NH_3 using BOTH reactants

$$2g N_2 \times \frac{1 \text{ mole}}{28g} = 0.071428 \text{ mole } N_2 \times \frac{2NH_3}{1N_2} = 0.142857 \text{ mole } NH_3$$

$\times 17g/\text{mole} = 2.43g NH_3$

$$2g H_2 \times \frac{1 \text{ mole}}{2.016g} = 0.992063 \text{ mole } H_2 \times \frac{2NH_3}{3H_2} = 0.661375 \text{ mole } NH_3$$

(C) Evaluate which reactant produced less NH_3 . This is the limiting reactant. Thus, the Limiting Reactant (LR) is N_2 . Now multiply by $M_m NH_3$ to get the

$$\text{The yield: } 0.142857 \text{ mole } NH_3 \times \frac{17g}{\text{mole}} = 2.43g NH_3$$

① Now find mass remaining reactant. The main thing here to know is that you must use the moles of the limiting reactant to first find the moles of the other reactant that were used in the rxn. Then this mass value is subtracted from the mass given.

we pull down the moles N_2 :

$$0.071428 \text{ moles } N_2 \times \frac{3H_2}{N_2} = 0.21428 \text{ moles } H_2 \times \frac{2.016g}{\text{mole}} = 0.432g H_2$$

↓
This is the mass of H_2 used up in the rxn when the 2g of N_2 reacts completely.

To find H_2 that remains behind:

$$\begin{array}{ccccc} 2g H_2 & - & 0.432g H_2 & = & 1.568g H_2 \\ \uparrow & & \uparrow & & \uparrow \\ \text{given} & & \text{used} & & \text{Remains} \\ & & & & \text{unreacted} \end{array}$$