Calculations from chemical equations

If you know the amount of any reactant or product involved in the reaction:

- ***you can calculate the amounts of all the other reactants and products that are consumed or produced in the reaction***

$$C_3H_8(g) + 5\ O_2(g) \longrightarrow 3\ CO_2(g) + 4\ H_2O(g)$$

BUT REMEMBER!

The coefficients in a chemical equation provide information ONLY about the proportions of MOLES of reactants and products

- ***given the number of moles of a reactant/product involved in a reaction, you CAN directly calculate the number of moles of other reactants and products consumed or produced in the reaction***
- ***given the mass of a reactant/product involved in a reaction, you can NOT directly calculate the mass of other reactants and products consumed or produced in the reaction***

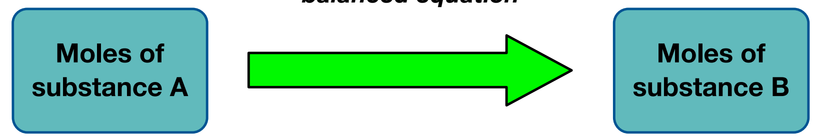
Mole - mole calculations

Given: • A balanced chemical equation

• A known quantity of one of the reactants/product (in moles)

Calculate: The quantity of one of the other reactants/products (in moles)

Use conversion factor based on ratio between coefficients of substances A and B from balanced equation



Mole - mole calculations

Example:

How many moles of ammonia are produced from 8.00 mol of hydrogen reacting with nitrogen?

Equation: $\boxed{3\ H_2} + N_2 \longrightarrow \boxed{2\ NH_3}$

Conversion factor:

Mole ratio between unknown substance (ammonia) and known substance (hydrogen):

$$\frac{2 \text{ moles } NH_3}{3 \text{ moles } H_2}$$

$$8.00 \cancel{\text{ moles } H_2} \left(\frac{2 \text{ moles } NH_3}{3 \cancel{\text{ moles } H_2}} \right) = \mathbf{5.33 \text{ moles } NH_3}$$

Remember the baking analogy?


1 bag flour + 1 carton milk + 6 eggs $\xrightarrow{\Delta}$ 24 pancakes

How many eggs do you need to make 60 pancakes?

Conversion factor between eggs and pancakes: $\dfrac{6 \text{ eggs}}{24 \text{ pancakes}}$

Remember the baking analogy?

$$1 \text{ bag flour} + 1 \text{ carton milk} + 6 \text{ eggs} \xrightarrow{\Delta} 24 \text{ pancakes}$$

How many eggs do you need to make 60 pancakes?

$$60 \cancel{\text{pancakes}} \left[\frac{6 \text{ eggs}}{24 \cancel{\text{pancakes}}} \right] = \mathbf{15 \text{ eggs}}$$

Mole - mole calculations

Given the balanced equation:

$$\boxed{K_2Cr_2O_7} + \boxed{6\ KI} + 7\ H_2SO_4 \longrightarrow Cr_2(SO_4)_3 + 4\ K_2SO_4 + 3\ I_2 + 7\ H_2O$$

a) How many moles of potassium dichromate ($K_2Cr_2O_7$) are required to react with 2.0 mol of potassium iodide (KI)

Conversion Factor:
Mole ratio between the unknown substance (potassium dichromate) and the known substance (potassium iodide):

$$\frac{1 \text{ mol } K_2Cr_2O_7}{6 \text{ mol KI}}$$

$$2.0 \cancel{\text{mol KI}} \left[\frac{1 \text{ mol } K_2Cr_2O_7}{6 \cancel{\text{mol KI}}} \right] = \mathbf{0.33 \text{ mol } K_2Cr_2O_7}$$

Mole - mole calculations

Given the balanced equation:

$$K_2Cr_2O_7 + 6\ KI + \boxed{7\ H_2SO_4} \longrightarrow Cr_2(SO_4)_3 + 4\ K_2SO_4 + \boxed{3\ I_2} + 7\ H_2O$$

b) How many moles of sulfuric acid (H_2SO_4) are required to produce 2.0 moles of iodine (I_2)

Conversion factor:
Mole ratio between the unknown substance (sulfuric acid) and the known substance (iodine):

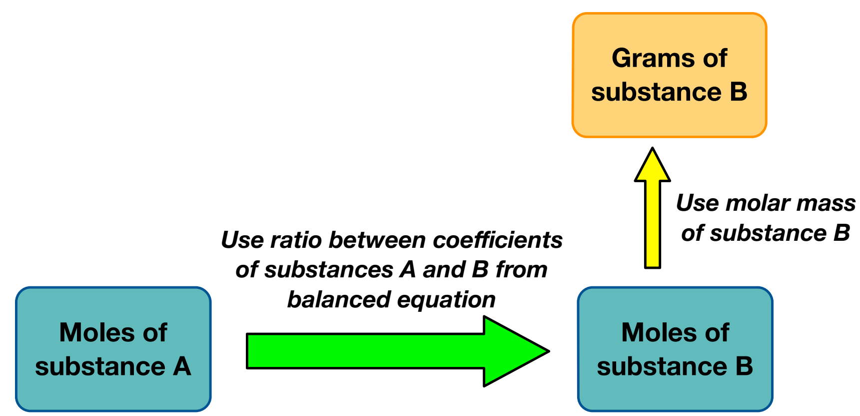
$$\frac{7 \text{ mol } H_2SO_4}{3 \text{ mol } I_2}$$

$$2.0 \cancel{\text{mol } I_2} \left[\frac{7 \text{ mol } H_2SO_4}{3 \cancel{\text{mol } I_2}} \right] = \mathbf{4.7 \text{ mol } H_2SO_4}$$

Mole - mass calculations

Given:
- A balanced chemical equation
- A known quantity of one of the reactants/product (in moles)

Calculate: The mass of one of the other reactants/products (in grams)



Mole - mass calculations

Example:

What mass of hydrogen is produced by reacting 6.0 mol of aluminum with hydrochloric acid?

Equation: $\boxed{2\ Al(s)} + 6\ HCl(aq) \longrightarrow 2\ AlCl_3(aq) + \boxed{3\ H_2(g)}$

Conversion Factor:

Mole ratio between unknown substance (hydrogen) and known substance (aluminum):

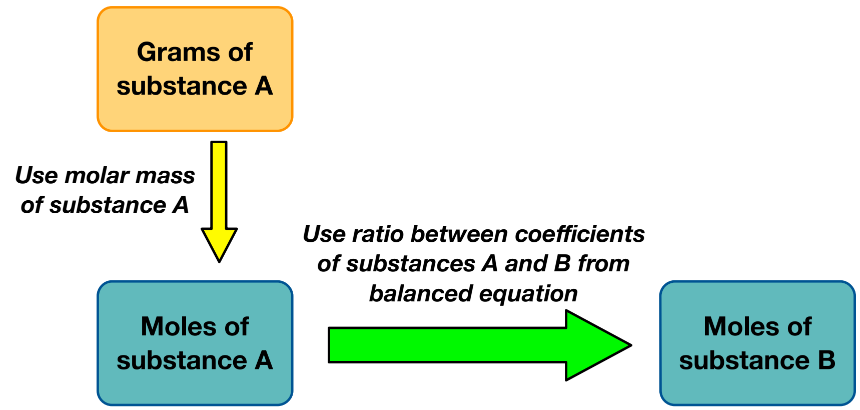
$$\frac{3\ \text{mol}\ H_2}{2\ \text{mol}\ Al}$$

$$6.0\ \cancel{\text{mol Al}} \left[\frac{3\ \text{mol}\ H_2}{2\ \cancel{\text{mol Al}}}\right] = 9.0\ \cancel{\text{mol}\ H_2} \left[\frac{2.016\ \text{g}}{1\ \cancel{\text{mol}\ H_2}}\right] = \mathbf{18\ g\ H_2}$$

Mass - mole calculations

Given:
- A balanced chemical equation
- A known mass of one of the reactants/product (in grams)

Calculate: The quantity of one of the other reactants/products (in moles)



Mass - mole calculations

How many moles of silver nitrate ($AgNO_3$) are required to produce 100.0 g of silver sulfide (Ag_2S)?

$$2\ AgNO_3 + H_2S \longrightarrow \boxed{Ag_2S} + 2\ HNO_3$$

Step 1: *Convert the amount of known substance (Ag_2S) from grams to moles*

$$100.0\ \cancel{\text{g}\ Ag_2S} \left[\frac{1\ \text{mol}\ Ag_2S}{247.87\ \cancel{\text{g}\ Ag_2S}}\right] = \mathbf{0.403\ mol\ Ag_2S}$$

Mass - mole calculations

How many moles of silver nitrate ($AgNO_3$) are required to produce 100.0 g of silver sulfide (Ag_2S)?

$$\boxed{2\ AgNO_3} + H_2S \longrightarrow \boxed{Ag_2S} + 2\ HNO_3$$

Step 2: *Determine the number of moles of the unknown substance ($AgNO_3$) required to produce the number of moles of the known substance (0.403 mol Ag_2S)*

Conversion Factor:

Mole ratio between the unknown substance (silver nitrate) and the known substance (silver sulfide):

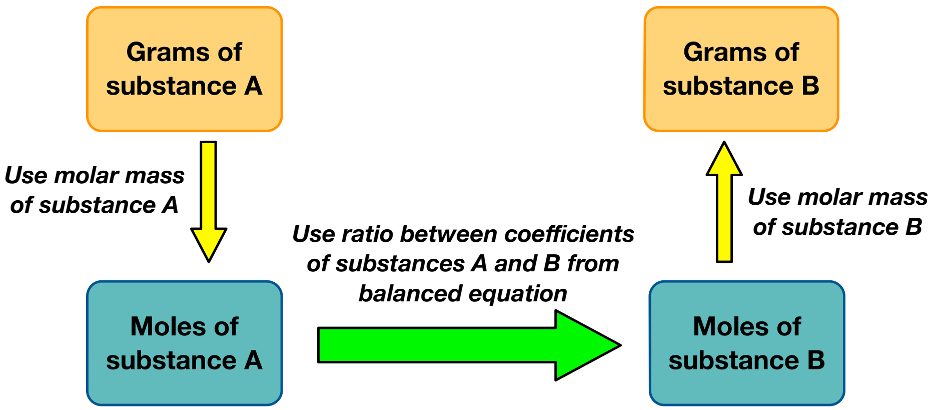
$$\frac{2\ \text{mol}\ AgNO_3}{1\ \text{mol}\ Ag_2S}$$

$$0.403\ \cancel{\text{mol}\ Ag_2S} \left[\frac{2\ \text{mol}\ AgNO_3}{1\ \cancel{\text{mol}\ Ag_2S}}\right] = \mathbf{0.806\ mol\ AgNO_3}$$

Mass - mass calculations

Given: • A balanced chemical equation
• A known mass of one of the reactants/product (in grams)

Calculate: The mass of one of the other reactants/products (in grams)



Mass - mass calculations

How many grams of nitric acid are required to produce 8.75 g of dinitrogen monoxide (N_2O)?

The balanced equation is:

$$4\,Zn(s) + 10\,HNO_3(aq) \longrightarrow 4\,Zn(NO_3)_2(aq) + N_2O(g) + 5\,H_2O(l)$$

Step 1: Convert the amount of known substance (N_2O) from grams to moles

Molar mass N_2O: (2 x 14.01 g/mol) + 16.00 g/mol = **44.02 g/mol**

$$8.75\ \cancel{g\ N_2O}\left[\frac{1\ mol\ N_2O}{44.02\ \cancel{g\ N_2O}}\right] = \mathbf{0.199\ mol\ N_2O}$$

Mass - mass calculations

How many grams of nitric acid are required to produce 8.75 g of dinitrogen monoxide (N_2O)?

The balanced equation is:

$$4\,Zn(s) + \boxed{10\,HNO_3(aq)} \longrightarrow 4\,Zn(NO_3)_2(aq) + \boxed{N_2O(g)} + 5\,H_2O(l)$$

Step 2: Determine the number of moles of the unknown substance (HNO_3) required to produce the number of moles of the known substance (0.199 mol N_2O)

Conversion Factor:
Mole ratio between the unknown substance (nitric acid) and the known substance (dinitrogen monoxide):

$$\frac{10\ mol\ HNO_3}{1\ mol\ N_2O}$$

$$0.199\ \cancel{mol\ N_2O}\left[\frac{10\ mol\ HNO_3}{1\ \cancel{mol\ N_2O}}\right] = \mathbf{1.99\ mol\ HNO_3}$$

Mass - mass calculations

How many grams of nitric acid are required to produce 8.75 g of dinitrogen monoxide (N_2O)?

The balanced equation is:

$$4\,Zn(s) + 10\,HNO_3(aq) \longrightarrow 4\,Zn(NO_3)_2(aq) + N_2O(g) + 5\,H_2O(l)$$

Step 3: Convert the amount of unknown substance (1.99 moles HNO_3) from moles to grams

Molar mass HNO_3: 1.008 g/mol + 14.01 g/mol + (3 x 16.00 g/mol)
= **63.02 g/mol**

$$1.99\ \cancel{mol\ HNO_3}\left[\frac{63.02\ g\ HNO_3}{1\ \cancel{mol\ HNO_3}}\right] = \mathbf{125\ g\ HNO_3}$$

Mass - mass calculation: Another example

How many grams of carbon dioxide are produced by the complete combustion of 100. g of pentane (C_5H_{12})?

The balanced equation is:

$$C_5H_{12}(g) \;+\; 8\,O_2(g) \longrightarrow 5\,CO_2(g) \;+\; 6\,H_2O(g)$$

Step 1: Convert the amount of known substance (C_5H_{12}) from grams to moles

Molar mass C_5H_{12}: (5 x 12.01 g/mol) + (12 x 1.008 g/mol)

= **72.15 g/mol**

$$100.\ \cancel{g\ C_5H_{12}} \left[\frac{1\ \text{mol}\ C_5H_{12}}{72.15\ \cancel{g\ C_5H_{12}}}\right] = \mathbf{1.39\ mol\ C_5H_{12}}$$

Mass - mass calculation: Another example

How many grams of carbon dioxide are produced by the complete combustion of 100. g of pentane (C_5H_{12})?

The balanced equation is:

$$\boxed{C_5H_{12}(g)} \;+\; 8\,O_2(g) \longrightarrow \boxed{5\,CO_2(g)} \;+\; 6\,H_2O(g)$$

Step 2: Determine the number of moles of the unknown substance (CO_2) required to produce the number of moles of the known substance (1.39 mol C_5H_{12})

Conversion Factor:

Mole ratio between the unknown substance (carbon dioxide) and the known substance (pentane):

$$\frac{5\ \text{mol}\ CO_2}{1\ \text{mol}\ C_5H_{12}}$$

$$1.39\ \cancel{\text{mol}\ C_5H_{12}} \left[\frac{5\ \text{mol}\ CO_2}{1\ \cancel{\text{mol}\ C_5H_{12}}}\right] = \mathbf{6.95\ mol\ CO_2}$$

Mass - mass calculation: Another example

How many grams of carbon dioxide are produced by the complete combustion of 100. g of pentane (C_5H_{12})?

The balanced equation is:

$$C_5H_{12}(g) \;+\; 8\,O_2(g) \longrightarrow 5\,CO_2(g) \;+\; 6\,H_2O(g)$$

Step 3: Convert the amount of unknown substance (6.95 moles CO_2) from moles to grams

Molar mass CO_2: 12.01 g/mol + (2 x 16.00 g/mol) = **44.01 g/mol**

$$6.95\ \cancel{\text{mol}\ CO_2} \left[\frac{44.01\ \text{g}\ CO_2}{1\ \cancel{\text{mol}\ CO_2}}\right] = \mathbf{306\ g\ CO_2}$$

Yields

Theoretical yield -- the calculated amount (mass) of product that can be obtained from a given amount of reactant based on the balanced chemical equation for a reaction

Actual yield -- the amount of product actually obtained from a reaction

The actual yield observed for a reaction is almost always less than the theoretical yield due to:

- side reactions that form other products
- incomplete / reversible reactions
- loss of material during handling and transfer from one vessel to another

The actual yield should never be greater than the theoretical yield

— if it is, it is an indicator of experimental error

$$\textbf{Percent yield} = 100 \times \frac{\text{actual yield}}{\text{theoretical yield}}$$

Yields

Silver bromide was prepared by reacting 200.0 g of magnesium bromide and an excess amount of silver nitrate.

Equation: $MgBr_2 + 2\ AgNO_3 \longrightarrow Mg(NO_3)_3 + 2\ AgBr$

a) What is the theoretical yield of silver bromide?

Step 1: Convert the amount of $MgBr_2$ from grams to moles

$200.0\ \cancel{g\ MgBr_2}\ (1\ \text{mol}\ MgBr_2 / 184.1\ \cancel{g\ MgBr_2}) = 1.086\ \text{mol}\ MgBr_2$

Yields

Silver bromide was prepared by reacting 200.0 g of magnesium bromide and an excess amount of silver nitrate.

Equation: $MgBr_2 + 2\ AgNO_3 \longrightarrow Mg(NO_3)_3 + 2\ AgBr$

a) What is the theoretical yield of silver bromide?

Step 2: Determine how many moles of AgBr can be formed from this amount of $MgBr_2$ (i.e., 1.086 moles)

$$1.086\ \cancel{\text{mol}\ MgBr_2} \left(\frac{2\ \text{mol AgBr}}{1\ \cancel{\text{mol}\ MgBr_2}} \right) = \mathbf{2.172\ mol\ AgBr}$$

Step 3: Convert from moles to grams

This is the *theoretical yield*

$2.172\ \cancel{\text{mol AgBr}}\ (187.8\ \text{g AgBr} / 1\ \cancel{\text{mol AgBr}}) = \mathbf{407.9\ g\ AgBr}$

Yields

Silver bromide was prepared by reacting 200.0 g of magnesium bromide and an excess amount of silver nitrate.

Equation: $MgBr_2 + 2\ AgNO_3 \longrightarrow Mg(NO_3)_3 + 2\ AgBr$

b) Calculate the percent yield if 375.0 g of silver bromide was obtained from the reaction

theoretical yield = 407.9 g AgBr

$$\textbf{percent yield} = 100 \times \frac{\text{actual yield}}{\text{theoretical yield}}$$

$$\textbf{percent yield} = 100 \times \frac{375.0\ \text{g}}{407.9\ \text{g}} = \mathbf{91.93\ \%}$$

The concept of *limiting reactants*

In some chemical reactions, all reagents are present in the exact amounts required to completely react with one another.

Example: In a lab experiment, ammonia is produced by reacting 6.05 g of hydrogen gas (3 moles) with 28.02 g of nitrogen gas (1 mole)

$$3\ H_2 + N_2 \longrightarrow 2\ NH_3$$

In this case, hydrogen and nitrogen are said to react in ***stoichiometric amounts***

The concept of *limiting reactants*

But in many cases, a chemical reaction will take place under conditions where one (or more) of the reactants is present *in excess*

-- *i.e.,* there is more than enough of that reactant available for the reaction to proceed

Example: Combustion of 85.0 g of propane in air

$$C_3H_8(g) \ + \ 5\ O_2(g) \longrightarrow 3\ CO_2(g) \ + \ 4\ H_2O(g)$$

There is more than enough oxygen available in the air to react with all of the propane

-- the reaction will proceed until all of the 85.0 g of propane has been consumed

The concept of *limiting reactants*

But in many cases, a chemical reaction will take place under conditions where one (or more) of the reactants is present in excess

-- *i.e.,* there is more than enough of that reactant available for the reaction to proceed

The **limiting reactant** is the reactant that is ***not*** present in excess

-- the limiting reactant will be used up first (the reaction will stop when the limiting reactant is depleted)

-- the limiting reactant therefore limits the amount of product that can be formed by the reaction

In the previous example, ***propane*** *was the* ***limiting reactant*** *(oxygen was present in excess)*

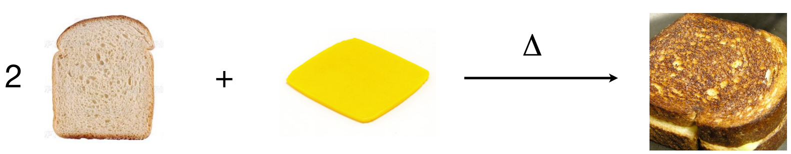
$$\boxed{C_3H_8(g)} \ + \ 5\ O_2(g) \longrightarrow 3\ CO_2(g) \ + \ 4\ H_2O(g)$$

Another food analogy…

Recipe for a grilled cheese sandwich:

Two slices bread and one slice of cheese gives one sandwich

Balanced equation:



If you have 10 slices of bread and 4 slices of cheese, how many sandwiches can you make?

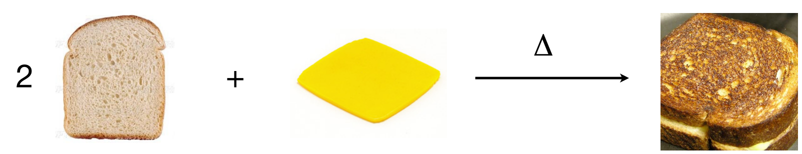
- enough bread for (10 / 2) = 5 sandwiches
- enough cheese for (4 / 1) = 4 sandwiches
- you can only make 4 sandwiches before the cheese is used up
- **cheese** is the limiting reactant

Another food analogy…

Recipe for a grilled cheese sandwich:

Two slices bread and one slice of cheese gives one sandwich

Balanced equation:



If you have 8 slices of bread and 6 slices of cheese, how many sandwiches can you make?

- enough bread for (8 / 2) = 4 sandwiches
- enough cheese for (6 / 1) = 6 sandwiches
- you can only make 4 sandwiches before the bread is used up
- **bread** is the limiting reactant

Limiting reactants

Chemistry example:

Hydrogen and chlorine gas combine to form hydrogen chloride:

$$H_2 + Cl_2 \longrightarrow 2\ HCl$$

If 4 moles of hydrogen reacts with 3 moles of chlorine, how many moles of HCl will be formed?

How much HCl can be formed from 4 mol H_2?

$$4\ \cancel{\text{mol } H_2}\left[\frac{2\text{ mol HCl}}{1\ \cancel{\text{mol } H_2}}\right] = \textbf{8 mol HCl}$$

How much HCl can be formed from 3 mol Cl_2?

$$3\ \cancel{\text{mol } Cl_2}\left[\frac{2\text{ mol HCl}}{1\ \cancel{\text{mol } Cl_2}}\right] = \textbf{6 mol HCl}$$

Limiting reactants

Chemistry example:

Hydrogen and chlorine gas combine to form hydrogen chloride:

$$H_2 + Cl_2 \longrightarrow 2\ HCl$$

If 4 moles of hydrogen reacts with 3 moles of chlorine, how many moles of HCl will be formed?

8 mol HCl can be formed from 4 mol of H_2

6 mol HCl can be formed from 3 mol of Cl_2

3 moles of H_2 will react with 3 moles of Cl_2

- At this point, the Cl_2 will have been completely consumed and the reaction stops (chlorine is the ***limiting reactant***)
- 1 mole of H_2 will remain unreacted (hydrogen is present in excess)
- 6 moles of HCl will have been formed

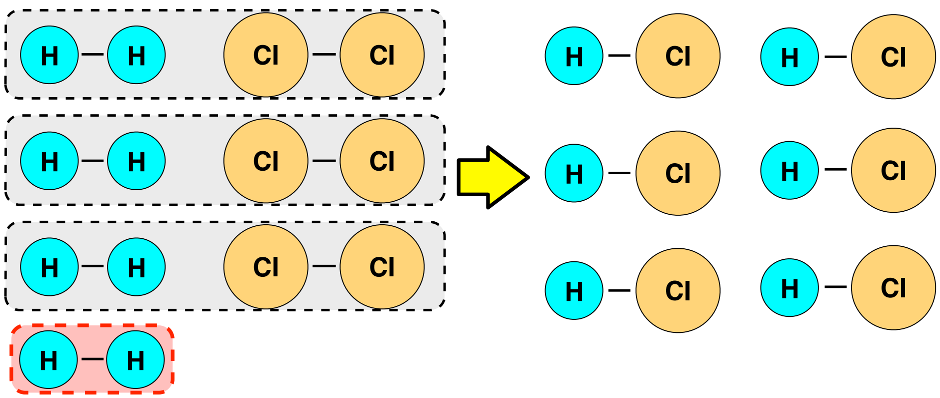
Limiting reactants

Chemistry example:

Hydrogen and chlorine gas combine to form hydrogen chloride:

$$H_2 + Cl_2 \longrightarrow 2\ HCl$$

If 4 moles of hydrogen reacts with 3 moles of chlorine, how many moles of HCl will be formed?



Procedure for identifying the limiting reactant

1. Calculate the amounts of ***product*** that can be formed from each of the reactants
2. Determine which reactant gives the least amount of product **-- *this is the limiting reactant***

Do *not* just compare the numbers of moles of reactants -- you must also account for the ratios in which the reactants combine

3. To find the amount of the non-limiting reactant remaining after the reaction:

   -- *calculate the amount of the non-limiting reactant required to react completely with the limiting reactant*

   -- *subtract this amount from the starting quantity of the non-limiting reactant*

Limiting reactant

How many grams of silver bromide (AgBr) can be formed when solutions containing 55.0 g of $MgBr_2$ and 95.0 g of $AgNO_3$ are mixed together?

Equation: $MgBr_2 + 2\,AgNO_3 \longrightarrow 2\,AgBr + Mg(NO_3)_2$

Convert the amounts of reactants from grams to moles

$$55.0\ \cancel{g\ MgBr_2}\left(\frac{1\ mol\ MgBr_2}{184.11\ \cancel{g\ MgBr_2}}\right) = \mathbf{0.299\ mol\ MgBr_2}$$

$$95.0\ \cancel{g\ AgNO_3}\left(\frac{1\ mol\ AgNO_3}{169.91\ \cancel{g\ AgNO_3}}\right) = \mathbf{0.559\ mol\ AgNO_3}$$

Now determine how much product can be formed from each reactant

Limiting reactant

How many grams of silver bromide (AgBr) can be formed when solutions containing 55.0 g of $MgBr_2$ and 95.0 g of $AgNO_3$ are mixed together?

Equation: $MgBr_2 + 2\,AgNO_3 \longrightarrow 2\,AgBr + Mg(NO_3)_2$

Amount of AgBr that can be produced by 0.299 mol of $MgBr_2$:

$$0.299\ \cancel{mol\ MgBr_2}\left(\frac{2\ mol\ AgBr}{1\ \cancel{mol\ MgBr_2}}\right) = \mathbf{0.597\ mol\ AgBr}$$

Amount of AgBr that can be produced by 0.559 mol of $AgNO_3$:

$$0.559\ \cancel{mol\ AgNO_3}\left(\frac{2\ mol\ AgBr}{2\ \cancel{mol\ AgNO_3}}\right) = \mathbf{0.559\ mol\ AgBr}$$

$AgNO_3$ is the limiting reactant

Limiting reactant

How many grams of silver bromide (AgBr) can be formed when solutions containing 55.0 g of $MgBr_2$ and 95.0 g of $AgNO_3$ are mixed together?

Equation: $MgBr_2 + 2\,AgNO_3 \longrightarrow 2\,AgBr + Mg(NO_3)_2$

Amount of AgBr that can be produced by 0.559 mol of $AgNO_3$:

$$0.559\ \cancel{mol\ AgNO_3}\left(\frac{2\ mol\ AgBr}{2\ \cancel{mol\ AgNO_3}}\right) = \mathbf{0.559\ mol\ AgBr}$$

$AgNO_3$ is the limiting reactant

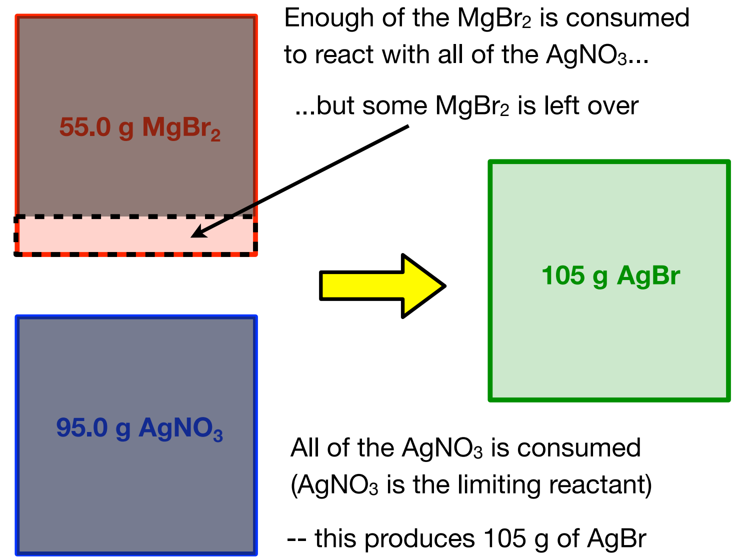
Convert moles of AgBr to grams of AgBr:

$$0.559\ \cancel{mol\ AgBr}\left(\frac{187.8\ g\ AgBr}{1\ \cancel{mol\ AgBr}}\right) = \mathbf{105\ g\ \ AgBr}$$

Limiting reactant

Enough of the $MgBr_2$ is consumed to react with all of the $AgNO_3$...

...but some $MgBr_2$ is left over



All of the $AgNO_3$ is consumed ($AgNO_3$ is the limiting reactant)

-- this produces 105 g of AgBr

Limiting reactant

How many grams of silver bromide (AgBr) can be formed when solutions containing 55.0 g of $MgBr_2$ and 95.0 g of $AgNO_3$ are mixed together?

Equation: $MgBr_2 + 2\ AgNO_3 \longrightarrow 2\ AgBr + Mg(NO_3)_2$

How many grams of the excess reactant remain unreacted?

Since $AgNO_3$ is the limiting reactant, all 95.0 g of $AgNO_3$ will be used up.

Calculate how much $MgBr_2$ (the excess reactant) is required to react with 95.0 g $AgNO_3$ and then determine how much $MgBr_2$ is left over.

Remember that in the previous steps, we calculated that 95.0 g of $AgNO_3$ is equal to 0.559 moles of $AgNO_3$.

So we first need to determine how many moles of $MgBr_2$ are required to react with 0.559 moles of $AgNO_3$

-- then convert to grams and subtract from the original amount of $MgBr_2$

Limiting reactant

How many grams of silver bromide (AgBr) can be formed when solutions containing 55.0 g of $MgBr_2$ and 95.0 g of $AgNO_3$ are mixed together?

Equation: $MgBr_2 + 2\ AgNO_3 \longrightarrow 2\ AgBr + Mg(NO_3)_2$

How many grams of the excess reactant remain unreacted?

$$0.559\ \cancel{\text{mol } AgNO_3} \left[\frac{1\ \text{mol } MgBr_2}{2\ \cancel{\text{mol } AgNO_3}} \right] = 0.280\ \text{mol } MgBr_2$$

$$0.280\ \cancel{\text{mol } MgBr_2}\ (184.11\ \text{g } MgBr_2 / 1\ \cancel{\text{mol } MgBr_2}) = \mathbf{51.5\ g\ MgBr_2}$$

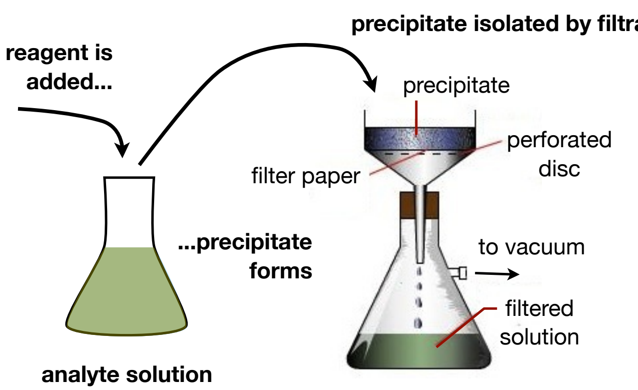
When all 95.0 g of $AgNO_3$ has reacted, 51.5 g of $MgBr_2$ will have been consumed. The amount of unreacted $MgBr_2$ left over is given by:

$$55.0\ \text{g} - 51.5\ \text{g} = \mathbf{3.5\ g\ MgBr_2}$$

Gravimetric analysis

Gravimetric analysis is a chemical analytical method based on the measurement of *masses*

- it can be used in combination with precipitation reactions to determine the amount of dissolved substances present in a solution



Using stoichiometry relationships, the amounts of dissolved substances in the solution can be determined from the mass of precipitate that is formed

Example: A 1-liter sample of industrial wastewater is analyzed for lead (in its Pb^{2+} ionic form) by gravimetric analysis.

This is done by adding excess sodium sulfate to the water sample to precipitate lead (II) sulfate.

The mass of $PbSO_4$ produced is 300.0 mg. **Calculate the mass of lead in the water sample.**

Solution: The dissolved Pb^{2+} ions in the water sample will react with the SO_4^{2-} ions added as sodium sulfate to form insoluble $PbSO_4$

Net ionic equation: $Pb^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow PbSO_4(s)$

This is a mass–mass stoichiometry problem

It also involves a limiting reactant, but the problem statement tells you what it is beforehand — i.e., $Pb^{2+}(aq)$

Net ionic equation: $Pb^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow PbSO_4(s)$

Step 1: *Convert mass of $PbSO_4$ produced to moles*

$$300.0\ \cancel{mg}\ PbSO_4 \times \frac{1\ g}{1000\ \cancel{mg}} = 0.3000\ g\ PbSO_4$$

$$0.3000\ \cancel{g\ PbSO_4} \times \frac{1\ mol\ PbSO_4}{303.3\ \cancel{g\ PbSO_4}} = 9.891 \times 10^{-4}\ mol\ PbSO_4$$

Step 2: *Calculate moles of Pb^{2+} required to produce the observed amount of $PbSO_4$*

$$9.891 \times 10^{-4}\ \cancel{mol\ PbSO_4} \left(\frac{1\ mol\ Pb^{2+}}{1\ \cancel{mol\ PbSO_4}} \right) = \mathbf{9.891 \times 10^{-4}\ mol\ Pb^{2+}}$$

Net ionic equation: $Pb^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow PbSO_4(s)$

Step 3: *Convert moles Pb^{2+} to mass*

$$9.891 \times 10^{-4}\ \cancel{mol\ Pb^{2+}} \times \frac{207.2\ g\ Pb^{2+}}{1\ \cancel{mol\ Pb^{2+}}} = 0.2049\ g\ Pb^{2+}$$

Note: atomic mass of Pb^{2+} = atomic mass of Pb

(electron mass is ignored)

$$0.2049\ \cancel{g}\ Pb^{2+} \times \frac{1000\ mg}{1\ \cancel{g}} = \mathbf{204.9\ mg\ Pb^{2+}}$$