Chapter 3: Chemical Equations and Reaction Stoichiometry

Chemical equations describe chemical reactions - they give a great deal of information:

1. Substances that react \( \text{REACTANTS} \)
2. Substances formed \( \text{PRODUCTS} \)
3. Relative amounts of the substances involved \( \text{(Figure 3-1)} \)

Note: I want to emphasize that reactions are NOT mathematical equations or equalities. Therefore

Therefore, \( \text{Reactants} \neq \text{Products} \)

"are not equal"

We write \( \text{Reactants} \rightarrow \text{Products} \)

"are changing into"

The reactions as written are describing a process. At the start, only reactants are present. At the finish, only products are present (providing the reaction went 100% to completion and stoichiometric amounts of reactants are present - more later). And so, as time progresses, substances (reactants) are changed into other substances (products). This semester, we are only concerned with what is present initially and what the final products are. Next semester, we will discuss why reactions go and how fast they go, among other topics.

Recall, Law of Conservation of Matter: matter is not created or destroyed; there is no detectable change in the quantity of matter in an ordinary chemical reaction.

This law allows us to "balance" equations using stoichiometric coefficients.

For each element, \( \# \text{atoms on reactant side} = \# \text{atoms on product side} \)
Example: Pentane (C$_5$H$_{12}$) burns in oxygen to form carbon dioxide (CO$_2$) and water (H$_2$O).

$$4 \quad C_5H_{12} + 8 O_2 \rightarrow 5 CO_2 + 6 H_2O$$

Hint: start balancing using the most complex compounds first e.g. C$_5$H$_{12}$, then balance the H$_2$O, leaving the single elements e.g. O$_2$ until the end.

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
<th>R</th>
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Example: Ammonia (NH$_3$) burns in oxygen gas to form nitrogen oxide (NO$_x$) and water.

$$2 \quad NH_3 + \frac{5}{2} O_2 \rightarrow 2 \quad NO + 3 \quad H_2O$$

<table>
<thead>
<tr>
<th>Reactants</th>
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<th>R</th>
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However it is more correct to use whole numbers for all coefficients: multiply by 2, giving

$$4 \quad NH_3 + 5 O_2 \rightarrow 4 \quad NO + 6 \quad H_2O$$

Another hint: Look for species that are the same on both sides and treat them as a unit:

Examples:
- NO$_3^-$ nitrate ion
- SO$_4^{2-}$ sulfate ion
- NH$_3$ ammonia/ NH$_4^+$ ammonium ion
- HNO$_3$ → Na NO$_3$
- H$_2$SO$_4$ → Mg SO$_4$
- NH$_3$ → NH$_4$Cl
Example: \[ \text{Fe(NO}_3\text{)}_3 + 3\text{NH}_3 + 3\text{H}_2\text{O} \rightarrow \text{Fe(OH)}_3 + 3\text{NH}_4\text{NO}_3 \]

Step 1: balance Fe - OK
2: balance NO\(_3\)
3: balance NH\(_3\)/NH\(_4^+\)
4: balance O using H\(_2\)O
5: check to see if H is balanced.

Calculations Based on Chemical Equations

Chemical equations represent a very precise versatile language. It is possible to extract a large amount of information from them.

Let us consider:

\[ \text{Fe}_2\text{O}_3 + 3\text{ CO} \rightarrow 2\text{ Fe} + 3\text{ CO}_2 \]

iron(III) oxide carbon monoxide iron carbon dioxide

REACTANTS PRODUCTS

Meaning: for every mole (or formula unit (think: molecule)) of \(\text{Fe}_2\text{O}_3\)
3 moles (or molecules) of \(\text{CO}\) are needed to form 2 moles (or atoms)
of \(\text{Fe}\) and 3 moles (or molecules) of \(\text{CO}_2\).

:: the equation relates moles and molecules - NOT mass!!

via the stoichiometric coefficients

:: (1) if there are 2 formula units of \(\text{Fe}_2\text{O}_3\), how much \(\text{CO}\) is needed? 6 molecules
(2) if there are 9 moles of \(\text{CO}\), how much \(\text{CO}_2\) can be formed? 9 moles
(3) if 8 atoms of \(\text{Fe}\) are formed, how much \(\text{CO}\) was needed? 12 molecules.

Example: How many carbon monoxide molecules (CO) are required to react with 130 formula units of \(\text{Fe}_2\text{O}_3\)?

We have learned that the stoichiometric coefficients give the relative number of moles of reactants needed and products produced.
\[
\frac{\text{molecules or moles CO}}{\text{formula units or moles } \text{Fe}_2\text{O}_3} = \frac{3}{1}
\]

\[ \therefore \text{moles of CO} = 3 \times \text{moles } \text{Fe}_2\text{O}_3 \]
\[ \therefore \text{molecules of CO} = 3 \times 130 \text{ formula units} = 390 \text{ molecules}. \]

**Example:** What mass of CO is required to react with 61.4 g of \( \text{Fe}_2\text{O}_3 \)?

\[
\text{Fe}_2\text{O}_3 + 3 \text{ CO} \rightarrow 2 \text{ Fe} + 3 \text{ CO}_2
\]

\[
\begin{align*}
61.4 \text{ g} & \quad ? \text{ g} \\
\end{align*}
\]

**Plan of Action:**

\[ q \text{ Fe}_2\text{O}_3 \rightarrow q \text{ Fe}_2\text{O}_3 \rightarrow 3 \text{ moles CO} \rightarrow q \text{ CO} \]

\[ (\text{moles} = \frac{q}{\text{MW}}) \quad (\text{stoch coeff}) \quad (q = \text{moles} \times \text{MW}) \]

1. moles \( \text{Fe}_2\text{O}_3 \) = \( \frac{q \text{ Fe}_2\text{O}_3}{\text{FW}} \) = \( \frac{61.4 \text{ g}}{159.7 \text{ g/mol}} \) = 0.384 mol

2. moles CO = \( \text{mol } \text{Fe}_2\text{O}_3 \times 3 \) = 1.15 mol

3. mass CO = \( \text{MW} \times \text{mol CO} \) = 28.0 \times 1.15 = 32.2 g

By dimensional analysis:

\[
\begin{align*}
\text{mass CO} & = 61.4 \text{ g } \text{Fe}_2\text{O}_3 \times \left( \frac{1 \text{ mol } \text{Fe}_2\text{O}_3}{159.7 \text{ g } \text{Fe}_2\text{O}_3} \right) \times \left( \frac{3 \text{ mol CO}}{1 \text{ mol } \text{Fe}_2\text{O}_3} \right) \times \left( \frac{28.0 \text{ g CO}}{1 \text{ mol CO}} \right) \\
& = 32.2 \text{ g}
\end{align*}
\]

**Example:** What mass of \( \text{Fe}_2\text{O}_3 \) reacted with excess CO if the \( \text{CO}_2 \) produced by the reaction has a mass of 5.86 g?

\[
\text{Fe}_2\text{O}_3 + 3 \text{ CO} \rightarrow 2 \text{ Fe} + 3 \text{ CO}_2
\]

\[
? \text{ g} \quad \text{excess} \quad 5.86 \text{ g}
\]

\[
\begin{array}{c|c|c}
\text{C} & 12 & \text{Fe} \\
\hline
\text{O} & 16 & \text{CO}_2 \\
\end{array}
\]

**Question:** is the MW of \( \text{CO}_2 \) 44 or 3\times 44 = 132? 44! MW does \underline{not} depend on stochiometric coefficient!!!
Plan of Action:  \[ g \, CO_2 \xrightarrow{1} \text{mole} \, CO_2 \xrightarrow{2} \text{mole} \, Fe_2O_3 \rightarrow g \, Fe_2O_3 \]

\[
\text{(mole} = \frac{g}{MW}) \quad \text{(use ratio of coefficients)} \quad (g = \text{mole} \times MW) \\
\]

1. mole CO\(_2\) = \frac{g \, CO_2}{MW} = \frac{5.86 \, g}{44.0 \, g/mol} = 0.133 \, \text{mole} \\

2. mole Fe\(_2\)O\(_3\) = \frac{\text{mole} \, CO \times \frac{1}{3}}{\text{mole} \, CO \times \frac{1}{3} \, \text{mole} \, Fe_2O_3} = 0.0444 \quad \text{since} \quad \text{mole} \, CO_2 = \frac{1}{3} \\

3. mass Fe\(_2\)O\(_3\) = \text{FW} \times \text{mole} \, Fe_2O_3 = 159.7 \, \text{g/mol} \times 0.0444 \, \text{mol} = 7.09 \, g \\

\[ \text{since} \quad \text{FW} = \frac{\text{g}}{\text{mol}} \]

Percent Yield from Chemical Reactions

Many reactions in real life do NOT go to completion, i.e., reactants are not completely transformed into the desired products. There are at least 2 reasons for this:

1. reactants did not react entirely 100\% to give the products.
2. reactants did react totally but formed "undesirable" products by "side reactions" as well as the wanted product, \( P_i \).

\[ R_1 + R_2 \rightarrow P_i \quad \text{(desired)} \]

\[ \quad \rightarrow \quad \text{P side reaction} \]

theoretical yield: assume \( R_1 + R_2 \rightarrow P \rightarrow 100\% \)

determined from stoichiometry

percent yield = \[ \frac{\text{actual yield of product}}{\text{theoretical yield of product}} \times 100 \]

To calculate \% yield, you can express yields in units of moles or mass.

\% yield problem types: ① calculate \% yield
② calculate actual yield, given \% yield and reactant
③ calculate amount of reactant necessary to give a certain amount of product knowing the \% yield.
% yield:

when you do the stoichiometry (g → mole → mole → g) you are working with the theoretical yield—what you should have if the reaction had gone 100%.

1. 90% yield
   Reactant: A → 90% B
   Product: 60 g

2. % yield
   Actual product: 85% A → 2B
   Reactant: 60 g
   Actual yield

3. % yield
   Actual product: 85% A → 2B
   Reactant: 60 g
   Actual yield

Plan:

Plan: gA → molA → molB → gB

Theoretical yield

Actual yield

% yield = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100
Example: Calculate the % yield of ethyl acetate if a 10.0 g sample of ethanol, $\text{C}_2\text{H}_5\text{OH}$, was boiled with excess acetic acid, $\text{CH}_3\text{COOH}$, to produce 14.8 g of ethylacetate, $\text{CH}_3\text{COOC}_2\text{H}_5$.

$$\text{C}_2\text{H}_5\text{OH} + \text{CH}_3\text{COOH} \rightarrow \text{CH}_3\text{COOC}_2\text{H}_5 + \text{H}_2\text{O} \quad 10.0\text{ g} \quad \text{excess} \quad 14.8\text{ g}$$

we know \[\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\]

1) find theoretical yield of ethyl acetate: (do the stoichiometric problem)

Plan: $g\quad \text{C}_2\text{H}_5\text{OH} \rightarrow \text{mole}\quad \text{C}_2\text{H}_5\text{OH} \rightarrow \text{mole}\quad \text{CH}_3\text{COOC}_2\text{H}_5 \rightarrow g\quad \text{CH}_3\text{COOH}$

$$\text{mole } \text{C}_2\text{H}_5\text{OH} = \frac{g}{\text{MW}} = \frac{10.0\text{ g}}{46.07\text{ g/mol}} = 0.217\text{ mole}$$

$$\text{mole } \text{CH}_3\text{COOC}_2\text{H}_5 = \text{mole } \text{C}_2\text{H}_5\text{OH} = 0.217\text{ mole}$$

$$\text{mass } \text{CH}_3\text{COOC}_2\text{H}_5 = \text{mole } \times \text{MW} = 0.217\text{ mole } \times 88.11\text{ g/mol} = 19.1\text{ g}$$

but only 14.8 g of product, ethyl acetate, were formed

2) \[\% \text{ yield} = \frac{14.8\text{ g}}{19.1\text{ g}} \times 100 = 77.5\%\]

In this reaction, we were expecting to make 19.1 g of ethylacetate. But we only produced 77.5% of that amount or 14.8 g.

Example: Using the same reaction as above, how much $\text{C}_2\text{H}_5\text{OH}$ is necessary to give a yield of 40.0 g of $\text{CH}_3\text{COOC}_2\text{H}_5$?

$$\text{C}_2\text{H}_5\text{OH} + \text{CH}_3\text{COOH} \rightarrow \text{CH}_3\text{COOC}_2\text{H}_5 + \text{H}_2\text{O} \quad \text{40.0 g}$$

Plan: 1) calculate theoretical yield

\[g\quad \text{CH}_3\text{COOC}_2\text{H}_5, \quad \% \text{ yield} \rightarrow \text{theoretical yield}\]

2) theoretical yield (g) $\rightarrow$ mole $\text{CH}_3\text{COOC}_2\text{H}_5$ $\rightarrow$ mole $\text{C}_2\text{H}_5\text{OH}$ $\rightarrow$ g $\text{C}_2\text{H}_5\text{OH}$
(1) \[ \% \text{ yield} = \frac{\text{actual}}{\text{theoretical}} \times 100 \]

\[ \therefore \text{theoretical} = \frac{\text{actual}}{\% \text{ yield}} \times 100 = \frac{40.0 g}{77.5 \%} \times 100 = 51.6 g \text{ CH}_3\text{COOC}_2\text{H}_5 \]

(2) \[ \text{moles CH}_3\text{COOC}_2\text{H}_5 = \frac{\# g}{\text{MW}} = \frac{51.6 g}{88.11 \text{ g/mol}} = 0.586 \text{ moles} \]

\[ \text{moles C}_2\text{H}_5\text{OH} = \text{moles CH}_3\text{COOC}_2\text{H}_5 = 0.586 \text{ moles} \]

\[ \text{mass C}_2\text{H}_5\text{OH} = \text{moles} \times \text{MW} = 0.586 \text{ moles} \times 46.07 \text{ g/mol} = 27.0 g \]

\[ \therefore \text{with this reaction having a yield of 77.5\% you need more C}_2\text{H}_5\text{OH than you would think to end up with 40.0 g of CH}_3\text{COOC}_2\text{H}_5. \]

The Limiting Reagent (Reactant) Concept

Until now, in problems, we have had either stoichiometric amounts of the reactants or we were told which reagent is in excess. But how do we decide which, if any, of the reactants is in excess and which reactant is the limiting - i.e. controlling the reaction?

Example: NON-CHEMICAL

Suppose we have 78 bolts, 113 washers, 99 nuts. How many sets, each consisting of 1 bolt, 2 washers and 1 nut, can we construct?

Equation: \[ 1 \text{ bolt} + 2 \text{ washers} + 1 \text{ nut} \rightarrow 1 \text{ set} \]

\[ \frac{78}{113} \text{ 99} \]

\[ 56 \frac{1}{2} \text{ pairs of washers} \]

\[ \therefore \text{we can only put together 56 total sets} \]

the number of washers limited the number of sets we could make.
Example: What is the maximum mass of sulfur dioxide, \(\text{SO}_2\), that can be produced by the reaction of 95.6 g of carbon disulfide, \(\text{CS}_2\), with 111 g of oxygen, \(\text{O}_2\)?

\[
\text{CS}_2 + 3 \text{ O}_2 \rightarrow 2 \text{SO}_2 + \text{CO}_2
\]

<table>
<thead>
<tr>
<th>Given</th>
<th>Value</th>
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<tbody>
<tr>
<td>(\text{CS}_2)</td>
<td>95.6 g</td>
</tr>
<tr>
<td>(\text{O}_2)</td>
<td>111 g</td>
</tr>
<tr>
<td>MW (\text{CS}_2)</td>
<td>76.13 \text{ g/mol}</td>
</tr>
<tr>
<td>MW (\text{O}_2)</td>
<td>32.0 \text{ g/mol}</td>
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</tbody>
</table>

Note: the MW of \(\text{O}_2\) is 32.0 \text{ g/mol}
Not \(3 \times 32 = 96\)
Do NOT include stoichiometric coefficients when determining MW.

Step 1: decide which reactant is limiting -- must work in moles!!

- determine number of moles of each reactant

\[
\text{moles } \text{CS}_2 = \frac{95.6 \text{ g}}{76.13 \text{ g/mol}} = 1.26 \text{ mol}
\]
\[
\text{moles } \text{O}_2 = \frac{111 \text{ g}}{32.0 \text{ g/mol}} = 3.47 \text{ mol}
\]

Let us assume that \(\text{CS}_2\) is limiting (50/50 chance)
- this means that \(\text{CS}_2\) is all used up in the reaction and \(\text{O}_2\) is excess
- to use up all \(\text{CS}_2\) we need at least \(3 \times 1.26 \text{ mol} = 3.78 \text{ mol}\)

But we only have 3.47 moles of \(\text{O}_2\) -- \(\text{O}_2\) must be limiting
\(\text{CS}_2\) is in excess

Mathematically, therefore, to determine the limiting reagent:

1. change mass \(\rightarrow\) moles
2. divide the number of moles by stoichiometric coefficient

\(\text{CS}_2\): \(1.26 \div 1 = 1.26\) \(\{\text{this is a normalization procedure like nuts + bolts example earlier}\}
\(\text{O}_2\): \(3.47 \div 3 = 1.16\)

3. the reactant with the smallest number is the limiting reagent and controls how much product is formed eg. \(\text{O}_2\) in this example.

All the other reactants are in excess and can be ignored.

All calculations are based on the amount of the limiting reagent!
\[ \text{CS}_2 + 3 \text{O}_2 \rightarrow 2 \text{SO}_2 + \text{CO}_2 \]

\[ \text{excess} \quad 3.47 \text{ mol} \]

**Step(2)** do the problem like any other stoichiometric problem

**Plan:** moles \( \text{O}_2 \rightarrow \text{moles SO}_2 \rightarrow q \text{ SO}_2 \)

\[
\text{mole SO}_2 = \frac{2}{3} \times \text{moles O}_2 = \frac{2 \times \text{mole SO}_2}{3 \times \text{mole O}_2} \times 3.47 \text{ mole O}_2 \]

\[ = 2.31 \text{ mole} \]

**Hint:** use dimensional analysis to check whether the ratio is \( \frac{2}{3} \) or \( \frac{3}{2} \).

\[
\text{mass SO}_2 = 2.31 \text{ mole} \times 64.1 \frac{\text{g}}{\text{mol}} = 148 \text{ g}
\]

**Note:** What is in the container when the reaction is done? \( \text{SO}_2, \text{CO}_2 + \text{CS}_2 \)

Exam question could be - how much \( \text{CS}_2 \) remains? \( 0.10 \text{ mole} \) or 7.4 g

**Concentrations of Solutions**

**Definition of solution:** a homogeneous mixture of 2 or more substances.

A solution is composed of 2 parts

1. **Solute:** dissolved substance (or substance in lesser amount)
2. **Solvent:** dissolving substance (or substance in greater amount)

**Example:** solution of \( \text{NaCl} \) in water, \( \text{H}_2\text{O} \): \( \text{NaCl} \): solute, \( \text{H}_2\text{O} \): solvent

**Concentration** = \( \frac{\text{amount of solute}}{\text{mass or volume of solution}} \)

Relative terms \{ dilute solution: small amount of solute in large amount of solvent, concentrated solution: large amount of solute in smaller amount of solvent \}

At this point, we will discuss 2 concentration units:

1. Percent by mass (do not confuse with % by mass of element in compound)
2. Molarity
1. Percent by mass:

\[
\% \text{ solute} = \frac{\text{mass solute}}{\text{mass solution}} \times 100
\]

\[
= \frac{\text{mass solute}}{\text{mass of solute} + \text{mass of solvent}} \times 100
\]

Note: if the question says the solution is aqueous or does not specify the solvent, the solvent is water, H\text{2}O.

Example: Calculate the mass of potassium nitrate, KNO\text{3}, required to prepare 250.0 g of solution that is 20.0\% KNO\text{3} by mass. What is the mass of water in the solution?

(a) \% by mass = \frac{g \text{ KNO}_3}{g \text{ solution}} \times 100

20.0\% = \frac{g \text{ KNO}_3}{250.0 \ g} \times 100

\[g \text{ KNO}_3 = \frac{20.0 \times 250.0 \ g}{100} = 50.0 \ g\]

(b) mass of solution = mass of KNO\text{3} + mass H\text{2}O

mass H\text{2}O = mass of solution - mass of KNO\text{3}

= 250.0 - 50.0 \ g

= 200.0 \ g

Example: What volume of 24.0\% lead(II) nitrate, Pb(NO\text{3})\text{2}, solution contains 400.0 g of Pb(NO\text{3})\text{2}? The density of the solution is 1.25 g/ml.

\[\text{Step 1: } \% \text{ by mass} = \frac{g \text{ Pb(NO}_3)_2}{g \text{ solution}} \times 100\]

24.0\% = \frac{40.0 \ g}{g \text{ solution}} \times 100

\[g \text{ solution} = \frac{40.0 \ g}{167 \ g} \times 100 = 24.0 \ mL\]

\[\text{Step 2: } D(\text{mL}) = \frac{\# g \text{ solu}}{\# mL} \]

\[\# mL = \frac{\# g}{D} = \frac{167 \ g}{1.25 \ g/mL} = 134 \ mL\]
2. Molarity:

\[
molarity \ (M) = \frac{\text{moles solute}}{\text{liters of solution}} \quad \text{write: (mol/L) or (mmol/mL)}
\]

(Show Fig. 3-2: preparation of a solution)

Example: Calculate the mass of ammonium sulfate, \((NH_4)_2SO_4\), required to prepare 100 mL of 0.300 M \((NH_4)_2SO_4\) solution

\[
M \ (\text{mol/L}) = \frac{\# \text{mol}}{\# \text{L}}
\]

Plan: \(M, \# \text{L} \rightarrow \# \text{mol} \rightarrow g\)

Step 1 \(\# \text{mol} = M \times \# \text{L} = 0.300 \text{ mol/L} \times 0.100 \text{ L} = 0.0300 \text{ mol}\)

Step 2 \(\# g = \# \text{mol} \times MW\) since \(MW = \frac{\# g}{\# \text{mol}}\)

\[
\begin{align*}
\# g &= 0.0300 \text{ mol} \times 132.13 \frac{g}{\text{mol}} \\
&= 3.96 \text{ g}
\end{align*}
\]

Example: The specific gravity of commercial nitric acid is 1.44 and it is 70.4% \(HNO_3\) by mass. Calculate the molarity of \(HNO_3\).

Assume we have 100 g of solution, therefore the solution contains 70.4 g \(HNO_3\)

\[
\text{concentration} = \frac{70.4 \ g \ HNO_3}{100 \ g \ \text{solution}} \rightarrow \frac{\# \text{mol} \ HNO_3}{\# \text{L} \ \text{solution}}
\]

\[
\text{mol} = \frac{g}{MD}
\]

Plan: change numerator (g → mol), change denominator (g → #L), then divide, since \(M = \frac{\# \text{mol}}{\# \text{L}}\)
numerator: moles HNO₃ = \( \frac{9 \text{ g HNO}_3}{63.01 \text{ g/mol}} = 0.144 \text{ mol} \)

denominator: mL solution = \( \frac{100.0 \text{ g}}{1.44 \text{ g/mL}} = 69.4 \text{ mL} \)

\# L solution = \( 69.4 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.0694 \text{ L} \)

\[ \therefore M = \frac{\# \text{ mol}}{\# \text{ L}} = \frac{0.144 \text{ mol}}{0.0694 \text{ L}} = 16.1 \text{ M} \]

**Example:** What volume of 18.0 M H₂SO₄, sulfuric acid, is required to prepare 2.00 L of 1.40 M H₂SO₄?

This is a **DILUTION** problem, in which a solution is made less concentrated by adding more solvent (in this case, H₂O).

In a dilution problem, the number of moles of solute before dilution equals the number of moles of solute after dilution, since all we are adding is solvent (H₂O).

\[ \therefore \# \text{ moles before} = \# \text{ moles after} \]

\[ M_{\text{before}} \times \# L_{\text{before}} = M_{\text{after}} \times \# L_{\text{after}} \quad \therefore \text{mole} = M \times \# L \]

or more generally

\[ M_{\text{before}} V_{\text{before}} = M_{\text{after}} V_{\text{after}} \quad \text{where} \ Y = \text{mL, L} \]

\[ 18.0 \text{ M} \times V = 1.40 \text{ M} \times 2.00 \text{ L} \]

\[ V = 0.156 \text{ L} \quad \text{or} \quad 156 \text{ mL} \]

Stoichiometry and Solutions:

In stoichiometric calculations, we must work in moles of substance.

Up until now, our substances could be weighed out on a balance to get mass, then we converted mass (g) to moles.
We now have another way to handle substances in the laboratory - in solutions. Therefore if we know concentration (eg. M) and amount of solution (eg. #L), we can calculate amount of solute (eg. moles = M × #L) and use this in an equation.

Example: What volume of 0.100 M NaOH solution is required to react with 500 mL of 0.200 M Al(NO₃)₃?

\[
\text{Al(NO}_3\text{)}_3 (aq) + 3 \text{NaOH}(aq) \rightarrow \text{Al(OH)}_3 (s) + 3 \text{NaNO}_3 (aq)
\]

<table>
<thead>
<tr>
<th>(aq)</th>
<th>(s)</th>
</tr>
</thead>
<tbody>
<tr>
<td>means aqueous</td>
<td>solid</td>
</tr>
<tr>
<td>i.e. dissolved in H₂O</td>
<td></td>
</tr>
</tbody>
</table>

\[
\begin{align*}
0.200 \text{ M} & \quad 0.100 \text{ M} \\
500 \text{ mL} & \quad 0.050 \text{ L} \\
? \text{ L} & \\
\end{align*}
\]

Plan: 

\[
\begin{align*}
\text{M, #L} & \quad \text{mole} & \quad \text{mole} & \quad \text{#L} \\
\text{Al(NO}_3\text{)}_3 & \quad \text{Al(NO}_3\text{)}_3 & \quad \text{NaOH} & \quad \text{NaOH} \\
\end{align*}
\]

1. moles Al(NO₃)₃ = M × #L
   = 0.200 M × 0.050 L
   = 0.0100 moles

2. moles NaOH = 3 × moles Al(NO₃)₃
   = 0.0300 moles

3. #L NaOH = \frac{\text{# moles NaOH}}{M} = \frac{0.0300 \text{ mol}}{0.100 \text{ M}} = 0.300 \text{ L}

Example (continued): What mass of aluminum hydroxide, Al(OH)₃, precipitates (becomes a solid)?

From the equation, we know that for every mole of Al(NO₃)₃ used up, one mole of Al(OH)₃ (s) is formed.

\[
\begin{align*}
\text{# moles Al(OH)}_3 & = \text{# moles Al(NO}_3\text{)}_3 = 0.0100 \text{ moles} \\
\text{mass Al(OH)}_3 & = \text{moles} \times \text{MW} = 0.0100 \text{ mol} \times 78.0 \text{ g/mol} = 0.780 \text{ g}
\end{align*}
\]
In the laboratory, acid-base reactions are carried out as a titration:

**Titration**: process by which a solution of one reactant (the titrant) is added to the other reactant, and the volume of titrant required for complete reaction is measured.

The number of moles of one reactant is \( \text{moles} = M \times V \) known either by knowing \( M \) and \( V \) or by weighing out a known solid.

The volume of the other is known and the unknown concentration is calculated when the reaction is just completed. It is known when the reaction is finished by adding an **indicator** to the flask - it changes color when reaction is over.

**Standard Solution**: a solution of known concentration.

**Standardization**: process by which the concentration of a solution is determined. By carefully measuring the volume of solution required to react with an exactly known amount of a PRIMARY STANDARD. The standardized solution is known as the SECONDARY STANDARD and is used in the analysis of an unknown.

**Equivalence Point**: the point in a titration at which chemically equivalent amounts of reactants have reacted.

**Endpoint**: the point in a titration at which the indicator changes color and the titration is stopped.