Chapter 4: Stoichiometry: Quantitative Information About Chemical Reactions

A balanced reaction shows the quantitative relationships between reactants and products in a chemical reaction.

Consider \( \text{Fe}_2\text{O}_3(s) + 3\text{CO}(g) \rightarrow 2\text{Fe}(s) + 3\text{CO}_2(g) \)

<table>
<thead>
<tr>
<th>Initial</th>
<th>1 mol</th>
<th>3 mol</th>
<th>0 mol</th>
<th>0 mol</th>
</tr>
</thead>
<tbody>
<tr>
<td>Change</td>
<td>-1 mol</td>
<td>-3 mol</td>
<td>+2 mol</td>
<td>+3 mol</td>
</tr>
<tr>
<td>Final</td>
<td>0 mol</td>
<td>0 mol</td>
<td>2 mol</td>
<td>3 mol</td>
</tr>
</tbody>
</table>

Note: 1. The stoichiometric coefficients do not tell us how much we have. 2. They tell us the ratios in which reactants are lost and products are formed (more later).

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

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

\[
61.4\text{g} \quad ?\text{g}
\]

Plan of action:
1. \( \frac{g \text{Fe}_2\text{O}_3}{\text{FW}} \rightarrow \text{mol Fe}_2\text{O}_3 \rightarrow \text{mol CO} \rightarrow \frac{\text{FW} \ g \text{CO}}{\text{mol}} \)

\[
\begin{align*}
\text{1.} \ \?\text{mol Fe}_2\text{O}_3 &= \frac{61.4\text{g}}{159.7\text{g/mol}} = 0.384\text{mol} \\
\text{FW Fe}_2\text{O}_3 &= 2 \times 55.85 = 111.70 \\
\text{FW CO} &= 12.0 + 16.0 = 28.0\text{g/mol} \\
\text{2.} \ ?\text{mol CO} &= 0.384\text{mol Fe}_2\text{O}_3 \times \frac{3\text{mol CO}}{1\text{mol Fe}_2\text{O}_3} = 1.15\text{mol} \\
\text{FW CO} &= 12.0 + 16.0 = 28.0\text{g/mol} \\
\text{3.} \ ?\text{g CO} &= 1.15\text{mol CO} \times 28.0\text{g/mol} = 32.2\text{g CO} \\
\end{align*}
\]

HINT: don’t round until the end!

By dimensional analysis:

\[
\begin{align*}
?\text{g CO} &= 61.4\text{g Fe}_2\text{O}_3 \times \frac{1\text{mol Fe}_2\text{O}_3}{159.7\text{g Fe}_2\text{O}_3} \times \frac{3\text{mol CO}}{1\text{mol Fe}_2\text{O}_3} \times \frac{28.0\text{g CO}}{1\text{mol CO}} \\
&= 32.2\text{g}
\end{align*}
\]
Example: What mass of Fe₂O₃ reacted with excess CO if the CO₂ produced by the reaction has a mass of 5.86 g?

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

\[
\frac{\text{g}}{? \text{ g}} \quad \text{excess} \quad \frac{\text{g}}{5.86 \text{ g}} \quad \text{CO}_2
\]

Question: Is the FW of CO₂ 44.05 g/mol or 3 x 44.05 g/mol? 44 kg
FW does NOT depend on the stoichiometric coefficient.

Plan of Action:

1. \( g \text{ CO}_2 \rightarrow \text{ mol CO}_2 \) (FW)
2. \( \text{ mol Fe}_2\text{O}_3 \rightarrow \text{ mol Fe}_2\text{O}_3 \) (stoch coeff)
3. \( \text{ mol Fe}_2\text{O}_3 \rightarrow g \text{ Fe}_2\text{O}_3 \) (FW)

\[
? \ g \text{ Fe}_2\text{O}_3 = 5.86 \ g \text{ CO}_2 \times \frac{1 \text{ mol CO}_2}{44 \ g \text{ CO}_2} \times \frac{1 \text{ mol Fe}_2\text{O}_3}{3 \text{ mol CO}_2} \times \frac{159.7 \ g}{1 \text{ mol Fe}_2\text{O}_3}
\]

\[
= 7.09 \ g
\]

**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 to give the products.
2. Reactants did react 100% but formed "undesirable" products by "side reactions" as well as the wanted product.

\[
R_1 + R_2 \rightarrow P_1 \quad (\text{desired})
\]

\[
\rightarrow P_2 \quad \text{side reaction - happens often}
\]

**Theoretical yield:** assume \( R_1 + R_2 \rightarrow P_1 \) 100% completion according to stoichiometry

**Percent yield:**

\[
\text{percent yield} = \frac{\text{actual yield of product}}{\text{theoretical yield of product}} \times 100
\]
% yield problem types: (1) calculate % yield
(2) calculate actual yield, given % yield & reactant
more difficult (3) calculate amount of reactant necessary to
give a certain amnt. of product knowing % yield.

Example: Calculate the % yield of Fe, if a 10.0 g sample of
Fe₂O₃ reacted with excess CO to produce 5.61 g of Fe.

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

\[
10.0 \text{ g} \quad \text{excess} \quad 5.61 \text{ g}
\]

We know \( \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 \)

(1) Find theoretical yield of Fe (do straightforward stoich. problem)

Plan: \( g \text{ Fe}_2\text{O}_3 \rightarrow \text{mol Fe}_2\text{O}_3 \rightarrow \text{mol Fe} \rightarrow g \text{ Fe} \)

\[
? \text{ g Fe} = 10.0 \text{ g Fe}_2\text{O}_3 \times \frac{1 \text{ mol Fe}_2\text{O}_3}{159.7 \text{ g Fe}_2\text{O}_3} \times \frac{2 \text{ mol Fe}}{1 \text{ mol Fe}_2\text{O}_3} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}}
\]

\[
= 6.99 \text{ g}
\]

But only 5.61 g of Fe were produced!!

(2) % yield = \( \frac{5.61 \text{ g}}{6.99 \text{ g}} \times 100 \)

= 80.3%

So, in this reaction, we expected to make 6.99 g of Fe, but we only
produced 80.3% of that amount or 5.61 g of Fe.

Example: Using the same reaction, how many grams of Fe₂O₃
are required to produce 10.0 g of Fe assuming a percent yield of
85.0%?

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

\[
? \text{ g} \quad \text{excess} \quad 10.0 \text{ g}
\]
Plan: (1) Calculate theoretical yield from mass of Fe and % yield
(2) Using theoretical yield, calculate mass of Fe₂O₃.
\[ g \to \text{mol} \to \text{mol} \to g \]

(1) We know
\[ \% \text{yield} = \frac{\text{actual}}{\text{theoretical}} \times 100 \]
So, theoretical yield = \[ \frac{100.0 \text{g}}{85.0\%} \times 100 = 11.8 \text{g Fe} \]

(2) \[ ? \text{g Fe₂O₃} = 11.8 \text{g Fe} \times \frac{1 \text{mol Fe}}{55.85 \text{g Fe}} \times \frac{1 \text{mol Fe₂O₃}}{2 \text{mol Fe}} \times \frac{159.7 \text{g}}{1 \text{mol Fe₂O₃}} \]
\[ = 116.9 \text{g} \]

So with this reaction with a yield of 85.0%, you need more Fe₂O₃ than you would think to make 10.0 g of Fe.

The Limiting Reactant Concept

Up until now, we have had either stoichiometric amounts of reactants or we were told which reactant was in excess. But how do we decide, which, if any, of the reactants is in excess and which reactant is limiting, i.e. controlling how much product we can make?

Example: NON-CHEMICAL

Suppose we have 78 bolts, 112 washers, 99 nuts. How many sets, each consisting of 1 bolt, 2 washers and 1 nut, can we construct?

Equation: 1 bolt + 2 washers + 1 nut \[ \rightarrow \text{1 set} \]
\[ 78 \quad 112 \quad 99 \]
\[ \downarrow \]
56 pairs of washers

\[ \therefore \text{we can only put together 56 total sets} \]
the number of washers limited the number of sets we can make.

\[ \therefore \text{Washers is the limiting reactant!} \]
Example: CHEMICAL

Suppose we have 6.0 mol of CS₂ and 6.0 moles of O₂ in our container. How much SO₂ and CO₂ can we make?

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

initial \hspace{1cm} 6.0\text{mol} \hspace{1cm} 6.0\text{mol} \hspace{1cm} \text{0mol} \hspace{1cm} \text{0mol}

Now you have to make a choice. Is CS₂ going to be used up first or is O₂? Let's make an arbitrary decision — CS₂ will be used up first and its number of moles will go to 0.

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

change \hspace{1cm} -6.0\text{mol} \hspace{1cm} -18.0\text{mol}

final \hspace{1cm} 0\text{mol} \hspace{1cm} [-12.0\text{mol}] \hspace{1cm} \text{impossible}

So, CS₂ cannot be the limiting reactant; O₂ must be it.

From the beginning,

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

initial \hspace{1cm} 6.0\text{mol} \hspace{1cm} 6.0\text{mol} \hspace{1cm} \text{0mol} \hspace{1cm} \text{0mol}

change \hspace{1cm} -2.0\text{mol} \hspace{1cm} -6.0\text{mol} \hspace{1cm} +4.0\text{mol} \hspace{1cm} +2.0\text{mol}

\[
\begin{array}{c}
\frac{2}{3}(6.0)\\
\frac{2}{3}(6.0\text{mol})
\end{array}
\]

final \hspace{1cm} 4.0\text{mol} \hspace{1cm} 0\text{mol} \hspace{1cm} 4.0\text{mol} \hspace{1cm} 2.0\text{mol}

excess \hspace{1cm} limiting reactant

Now let's do another problem!
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>95.6 g</th>
<th>111 g</th>
</tr>
</thead>
<tbody>
<tr>
<td>MW:</td>
<td>76.13 g/mol</td>
<td>32.0 g/mol</td>
</tr>
</tbody>
</table>

Note: the MW of \( \text{O}_2 \) is 32.0 g/mol. Not \( 3 \times 32 = 96 \cdot \) 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

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

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

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 \( \text{O}_2 \) \( \therefore \text{O}_2 \) 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

\[
\text{CS}_2 : \quad 1.26 \div 1 = 1.26
\]

\[
\text{O}_2 : \quad 3.47 \div 3 = 1.16
\]

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{Step 2: do the problem like any other stoichiometric problem}

Plan: \text{mole O}_2 \rightarrow \text{mole SO}_2 \rightarrow \text{g SO}_2

? \text{mole SO}_2 = 3.47 \text{mole O}_2 \times \frac{2 \text{ mol SO}_2}{3 \text{ mol O}_2} = 2.31 \text{ mole}

*Hint: use dimensional analysis to decide if the ratio is \(\frac{2}{3}\) or \(\frac{3}{2}\)

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

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

How much of the excess reactant remains? in moles? in grams?

? \text{mol CS}_2 \text{ used up} = 3.47 \text{ mol O}_2 \times \frac{1 \text{ mol CS}_2}{3 \text{ mol O}_2} = 1.16 \text{ mol}

? \text{mol CS}_2 \text{ remains} = 1.26 \text{ mol} - 1.16 \text{ mol} = 0.10 \text{ mol}

\[? \text{ g CS}_2 \text{ remains} = 0.10 \text{ mol} \times 76.13 \text{ g/mol} = 7.6 \text{ g CS}_2\]

**Analysis of Mixtures:**

A sample of limestone and other soil materials is heated, and the limestone decomposes to give calcium oxide and carbon dioxide.

\[
\text{CaCO}_3(s) \xrightarrow{\Delta} \text{CaO}(s) + \text{CO}_2(g)
\]

A 1.500 g sample of limestone containing material gives 0.558 g of \text{CO}_2 in addition to \text{CaO} after being heated to a high temperature.

What is the mass percent of \text{CaCO}_3 in the original sample?

Plan: \text{g CO}_2 \rightarrow \text{mol CO}_2 \rightarrow \text{mol CaCO}_3 \rightarrow \text{g CaCO}_3 \rightarrow \% \text{CaCO}_3 \text{ in sample}
\[ \text{? g CaCO}_3 \text{ in sample} = 0.558 \text{ g CO}_2 \times \frac{1 \text{ mol CO}_2}{44.0 \text{ g CO}_2} \times \frac{1 \text{ mol CaCO}_3}{1 \text{ mol CO}_2} \times \frac{100.1 \text{ g}}{1 \text{ mol CaCO}_3} \]
\[ = 1.27 \text{ g} \]

\[ \therefore \text{ % CaCO}_3 \text{ in sample by mass} = \frac{\text{g CaCO}_3}{\text{g sample}} \times 100 \]
\[ = \frac{1.27 \text{ g CaCO}_3}{1.506 \text{ g sample}} \times 100 \]
\[ = 84.3 \% \text{ by mass} \]

**Concentrations of Solutions**

A solution is composed of two parts:

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

Example: NaCl (aq) solution of NaCl in water, H\textsubscript{2}O

\[ \text{Concentration} = \frac{\text{amount of solute}}{\text{mass or volume of solution}} \]

Relative terms:
- **Dilute solution**: small amount of solute in large amount of solvent
- **Concentrated solution**: large amount of solute in solvent

We will discuss 2 concentration units

1. **Percent by mass**
   - (do not confuse with % by mass of element in compound)
2. **Molarity**
3. **ppm (parts per million)** – see lab
   - lots of concentration units
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, \( \text{H}_2\text{O} \).

Example: Calculate the mass of potassium nitrate, \( \text{KNO}_3 \) required to prepare 250.0 g of solution that is 20.0 \% \( \text{KNO}_3 \) by mass. What is the mass of water in the solution?

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

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

\[ g \text{ KNO}_3 = \frac{200 \ \% \times 250.0 \ g}{100} \]

\[ = 50.0 \ g \]

(b) \( \text{mass of solution} = \text{mass of KNO}_3 + \text{mass H}_2\text{O} \)

\[ \text{mass H}_2\text{O} = \text{mass of solution} - \text{mass of KNO}_3 \]

\[ = 250.0 - 50.0 \ g \]

\[ = 200.0 \ g \]

Example: What volume of 24.0 \% lead(II) nitrate, \( \text{Pb(NO}_3)_2 \) solution contains 40.0 g of \( \text{Pb(NO}_3)_2 \)? The density of the solution is 1.25 g/mL.

Step 1 \( \% \text{ by mass} = \frac{9 \text{ g Pb(NO}_3)_2}{9 \text{ g solution}} \times 100 \)

24.0 \% = \( \frac{40.0 \ g}{9 \text{ g solution}} \times 100 \)

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

\[ = 16.7 \ g \]

Step 2: \( \frac{g}{\text{mL}} = \frac{\# \text{g solute}}{\# \text{mL}} \)

\[ \frac{g}{\text{mL}} = \frac{167 \ g}{125 \ g/mL} \]

\[ = 134 \ mL \]
2. Molarity:

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

(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.

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

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

Step 1: \(\#\text{ moles} = M \times \#L\)
\[
= 0.300 \frac{\text{mol}}{\text{L}} \times 0.100 \text{ L}
= 0.0300 \text{ mol}
\]

Step 2: \(\#g = \#\text{ moles} \times \text{MW}\)
\[
= 0.0300 \text{ mol} \times 132.13 \frac{g}{\text{mol}}
= 3.96 g
\]

Example: The specific gravity of commercial nitric acid is 1.44 and it is 70.4% \(\text{HNO}_3\) by mass. Calculate the molarity of \(\text{HNO}_3\).

Assume we have 100 g of solution; therefore, the solution contains 70.4 g \(\text{HNO}_3\).

\[
\begin{align*}
\text{concentration} & = \frac{70.4 \text{ g HNO}_3}{100 \text{ g solution}} \\
\Rightarrow \frac{\# \text{ moles HNO}_3}{\# \text{ L solution}} & = \frac{70.4}{100} \Rightarrow \frac{0.704}{\#L}
\end{align*}
\]

Plan: change numerator \((g \rightarrow \#\text{ moles})\); change denominator \((g \rightarrow \#L)\); then divide, since \(M = \frac{\#\text{ moles}}{\#L}\)
numerator: moles $\text{HNO}_3 = \frac{g \text{ HNO}_3}{\text{MW}} = \frac{70.4 \, g}{63.01 \, \text{g/mol}} = 1.12 \, \text{moles}$

denominator: mL solution = $\frac{g \text{ solution}}{\text{sp.gr. or density}} = \frac{100 \, g}{1.14 \, g/mL} = 87.7 \, \text{mL}$

$\# \, \text{L solution} = 87.7 \, \text{mL} \times \frac{1 \, \text{L}}{1000 \, \text{mL}} = 0.0877 \, \text{L}$

$\therefore M = \frac{\# \text{ mol}}{\# \text{ L}} = \frac{1.12 \, \text{ mol}}{0.0877 \, \text{ L}} = 12.7 \, \text{ M}$

Example: What volume of 18.0 M $\text{H}_2\text{SO}_4$, sulfuric acid, is required to prepare 2.00 L of 1.40 M $\text{H}_2\text{SO}_4$?

This is a DILUTION problem, in which a solution is made less concentrated by adding more solvent (in this case, $\text{H}_2\text{O}$).

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 ($\text{H}_2\text{O}$)

$\therefore \# \text{ moles before} = \# \text{ moles after}$

$M_{\text{before}} \times \# \text{L before} = M_{\text{after}} \times \# \text{L after}$

$\therefore \text{ mole} = M \times \# \text{L}$

or more generally

$M_{\text{before}} V_{\text{before}} = M_{\text{after}} V_{\text{after}}$ where $Y = \text{mL, L}$ any volume

$18.0 \, \text{M} \times Y = 1.40 \, \text{M} \times 2.00 \, \text{L}$

$V = \frac{0.156 \, \text{L}}{18.0 \, \text{M}}$, or 156 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 (e.g., M) and amount of solution (e.g., mL), we can calculate amount of solute (e.g., moles = M x #L) and use this in an equation.

**Example:** What volume of 0.100 M NaOH solution is required to react with 50.0 mL of 0.200 M Al(NO₃)₃?

![Note: (aq) means aqueous, i.e., dissolved in H₂O (s): solid]

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

<table>
<thead>
<tr>
<th>Component</th>
<th>Concentration</th>
<th>Volume (L)</th>
<th>M x V</th>
<th>Moles</th>
</tr>
</thead>
<tbody>
<tr>
<td>Al(NO₃)₃</td>
<td>0.200 M</td>
<td>0.050 L</td>
<td>0.0100</td>
<td>\text{(1)}</td>
</tr>
<tr>
<td>NaOH</td>
<td>0.100 M</td>
<td></td>
<td>\text{#L}</td>
<td>\text{#mol NaOH}</td>
</tr>
</tbody>
</table>

1. moles Al(NO₃)₃ = M x #L since M = \# mol/L
   \[= 0.200 \text{ M} \times 0.050 \text{ L} \]
   \[= 0.0100 \text{ mol} \]

2. moles NaOH = 3 x moles Al(NO₃)₃
   \[= 0.0300 \text{ mol} \]

3. #L NaOH = \frac{\# \text{ mol 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.

\[\therefore \# \text{ mol Al(OH)}_3 = \# \text{ mol Al(NO}_3\text{)}_3 = 0.0100 \text{ mol} \]

\[\therefore \text{mass Al(OH)}_3 = \text{mol} \times \text{MW} = 0.0100 \text{ mol} \times 78.0 \frac{\text{g}}{\text{mol}} = 0.780 \text{ g} \]
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.

\[ \text{mole of reactant} = \frac{M \times V}{\text{mole of titrant}} \]

the number of moles of one reactant is 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.
Redox titration

To analyze an iron-containing compound, you convert all the iron to Fe$^{2+}$ in solution then titrate the solution with standardized KMnO$_4$. The balanced net ionic equation is

$$\text{MnO}_4^- (aq) + 5 \text{Fe}^{2+} (aq) + 8 \text{H}_2\text{O}(aq) \rightarrow \text{Mn}^{2+} (aq) + 5 \text{Fe}^{3+} (aq) + 12 \text{H}_2\text{O}(l)$$

A 0.598 g sample of the iron-containing compound requires 22.25 mL of 0.0123 M KMnO$_4$ for titration to the equivalence point. What is the mass % of iron in the sample? (What happened to the K in KMnO$_4$?)

Plan: $M, V \text{KMnO}_4 \rightarrow \text{mol KMnO}_4 \rightarrow \text{mol Fe}^{2+} = \text{mol Fe} \rightarrow \text{g Fe (pure)}$

% Fe in sample

<table>
<thead>
<tr>
<th>Oxidation</th>
<th>Reduction reaction</th>
</tr>
</thead>
<tbody>
<tr>
<td>+5 2</td>
<td>+2 1-2 2 3 1-2</td>
</tr>
<tr>
<td>MnO$_4^-$</td>
<td>5 Fe$^{2+}$ + 8 H$_2$O$^+$</td>
</tr>
</tbody>
</table>

Mn is element being reduced $\text{Mn} \rightarrow \text{Mn}^{2+}$

$\text{MnO}_4^-$ is the oxidizing agent $\text{Fe} \rightarrow \text{Fe}^{2+}$

Fe is element being oxidized $\text{Fe} \rightarrow \text{Fe}$

$\text{Fe}^{2+}$ is the reducing agent

? g Fe (pure) = 0.02235 L MnO$_4^-$ $\times$ $\frac{0.0123 \text{ mol MnO}_4^-}{1 \text{ L MnO}_4^-}$ $\times$ $\frac{5 \text{ mol Fe}^{2+}}{1 \text{ mol MnO}_4^-}$ $\times$ $\frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}^{2+}}$

= 0.0764 g Fe

% Fe in ore = $\frac{0.0764 \text{ g Fe}}{0.598 \text{ g ore}} \times 100 = 12.8 \% \text{ Fe}$

* It's a spectator ion!
**pH - measure of the molar concentration of H⁺ ion**

In aqueous solution, the concentration of H⁺ can vary widely: from > 1 M to < 1 x 10⁻¹⁴ M. To manage these numbers, it is easier to work in logarithms.

Review: logs are exponents!

\[
\log_{10} 1000 = 3 \quad \quad \quad \log_{10} (1 \times 10^{-6}) = -6
\]

\[
10^3 = 1000 \quad \quad \quad 10^{-6} = 1 \times 10^{-6}
\]

**Definition of pH:**

\[
pH = -\log_{10} \left( \text{H⁺ concentration in molarity} \right)
\]

where [ ] means M

<table>
<thead>
<tr>
<th>[H⁺]</th>
<th>log[H⁺]</th>
<th>pH = -log[H⁺]</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.00 M</td>
<td>0</td>
<td>acidic</td>
</tr>
<tr>
<td>0.10 M</td>
<td>-1</td>
<td>+1 acidic</td>
</tr>
<tr>
<td>1 x 10⁻⁵ M</td>
<td>-5</td>
<td>+5 acidic</td>
</tr>
<tr>
<td>1 x 10⁻⁷ M</td>
<td>-7</td>
<td>7 neutral</td>
</tr>
<tr>
<td>1 x 10⁻⁹ M</td>
<td>-9</td>
<td>9 basic</td>
</tr>
<tr>
<td>1 x 10⁻¹⁴ M</td>
<td>-14</td>
<td>14 basic</td>
</tr>
</tbody>
</table>

What is the pH of a 2.5 x 10⁻⁴ HNO₃ solution?

HNO₃ is a strong acid so \([\text{HNO}_3] = [\text{H⁺}] \quad \text{HNO}_3 \rightarrow \text{H⁺} + \text{NO}_3^-\)

\[
pH = -\log_{10} (2.5 \times 10^{-4}) = 3.60
\]

What is the concentration of an HCl solution with pH = 1.28?

Another strong acid with \([\text{H⁺}] = [\text{HCl}]\)

\[
pH = -\log [\text{H⁺}] = 1.28
\]

\[
\log [\text{H⁺}] = -1.28
\]

Take inverse log (also called 10⁻¹) of both sides

\[
[\text{H⁺}] = 10^{-1.28} = 0.053
\]
Acid Base Titration

A solution of hydrochloric acid has a volume of 250. mL and pH of 1.92. Exactly 250. mL of 0.0105 M NaOH is added. What is the pH of the resulting solution?

\[\text{HCl(aq)} + \text{NaOH(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)}\]

\[
\begin{align*}
\text{pH} &= 1.92 \\
\text{0.250L} &= 0.250 L
\end{align*}
\]

\[\text{HCl : } \quad \text{pH} = 1.92 \quad \therefore -\log_{10}[H^+] = 1.92 \]
\[\log_{10}[H^+] = -1.92 \]
\[\frac{[H^+]}{10} = 10^{-1.92} = 0.0120 \text{ M} \]

\[? \text{ mol HCl} = 0.0120 \text{ M} \times 0.250 \text{ L} = 0.00300 \text{ mol HCl} \]

\[\text{NaOH : } \quad ? \text{ mol NaOH} = 0.0105 \text{ M} \times 0.250 \text{ L} = 0.002625 \text{ mol NaOH} \]

What kind of a problem is this? Limiting reagent!! But easy once!

What is limiting reactant? What gets used up first? NaOH

When the reaction is over, all NaOH is gone; been converted to NaCl \& H_2O

How much HCl remains?

Since stoichiometric ratio is 1:1, we can just subtract!

\[? \text{ mol HCl in excess} = 0.00300 \text{ mol HCl} - 0.002625 \text{ mol HCl used up} \]
\[= 0.000375 \text{ mol HCl} \]

To find the pH of the solution, we first need to find $[H^+]$

\[? \text{ M HCl} = \frac{0.000375 \text{ mol HCl}}{(0.250L + 0.250L)} = 0.00075 \text{ M} = [H^+] \text{ since HCl is strong acid} \]

\[\text{pH} = -\log_{10}[H^+] = -\log_{10}(0.00075) = 3.12 \]