Chapter 4

Chemical Reactions

Chemical Equations

Description of a reaction:

1 mole of ethanol reacts with 3 moles of oxygen to produce 2 moles of carbon dioxide and 3 moles of water.

or …

1 molecule of ethanol reacts with 3 **molecules** of oxygen to produce 2 **molecules** of carbon dioxide and 3 **molecules** of water

Representation of a chemical reaction:

A chemical equation

$$C_2H_5OH + 3O_2 \longrightarrow 2CO_2 + 3H_2O$$

reactants products

Reactants		Products
$C_2H_5OH + 3O_2$	$\longrightarrow$	$2CO_2 + 3H_2O$

C:	(1)(2)=**2**	(2)(1)=**2**
H:	(1)(5)+1=**6**	(3)(2)=**6**
O:	(1)(1)+(3)(2)=**7**	(2)(2)+(3)(1)=**7**

The equation is balanced.

All atoms present in the reactants are accounted for in the products.

Reaction equation can also contain information about the physical state of the reactants and products

State	Symbol
Solid	*(s)*
Liquid	*(l)*
Gas	*(g)*
Dissolved in water	*(aq)*

$$CaCO_3\,(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

- The balanced equation represents an overall **ratio** of reactants and products, **not** what actually happens during a **given** reaction, which can involve **arbitrary** amounts of reactants.

- Use the coefficients in the balanced equation to calculate the amount of each reactant that is used, and the amount of each product that is formed.

Balancing Chemical Equations

1. Determine what reaction is occurring. What are the reactants, the products, and the physical states involved?
2. Write the *unbalanced* equation that summarizes the reaction described in step 1.
3. Balance the equation by inspection. The same number of each type of atom needs to appear on both reactant and product sides. Do NOT change the formulas of any of the reactants or products.
4. If a polyatomic ion (SO_4^{2-}, NO_3^-, etc.), as part of a compound or alone, seems to survive the reaction intact, treating it like an "atom" may simplify balancing significantly.

Practice

Which of the following correctly balances the chemical equation given below? There may be more than one correct balanced equation. If a balanced equation is incorrect, explain why.

$$CaO + C \rightarrow CaC_2 + CO_2$$

I. $CaO_2 + 3C \rightarrow CaC_2 + CO_2$

II. $2CaO + 5C \rightarrow 2CaC_2 + CO_2$

III. $CaO + 2.5C \rightarrow CaC_2 + 0.5CO_2$

IV. $4CaO + 10C \rightarrow 4CaC_2 + 2CO_2$

Balancing chemical equations:

- doesn’t need to be an “art form”
- doesn’t need to involve “trial and error”

- can be done by applying general rules

A truly general method for more challenging reactions is more complicated than what we will see and use here, but it still would not be a trial-and-error process.

That is a widely repeated misconception

A general method for balancing reactions (involving neutral substances only)

1. Identify those elements that occur in **only one <u>compound</u>** (*<u>not</u>* an element like $O_2(g)$ or Cu(s)) on **<u>each side</u>** of the equation.
   — If more than one element qualifies, select the one that occurs in the compound with the most kinds of elements.

2. Balance the equation in the selected element.
   — In balancing an element, you fix the coefficient of one reactant and one product. Write the coefficient even if it is just 1, so you can easily see which compounds are "done" at any point.

3. Find an element that occurs in **only <u>one</u> <u>compound whose coefficient we haven't found yet</u>**. Repeat until all compounds have coefficients.

4. Balance any substances that are elements.

If you obtain fractional coefficients at any stage, multiply all known coefficients by the denominator of the fraction

Balancing reactions, Step 1

$$KMnO_4 + HCl \longrightarrow MnCl_2 + KCl + Cl_2 + H_2O$$

Find element(s) that occur in one compound on each side:

K in compound with with Mn, O, Cl ($KMnO_4$, KCl)

Mn in compound with K, O, Cl ($KMnO_4$, $MnCl_2$)

O in compound with K, Mn, H ($KMnO_4$, H_2O)

H in compound with Cl, O (HCl, H_2O)

K, Mn , and O are the best choices to start balancing since they are associated with 3 elements while H is associated with 2.

As we balance each element, formulas will be assigned coefficients. The coefficient marks the formula of a substance as “done”.

Balancing reactions, Step 2

$$\mathbf{1}\ KMnO_4 + HCl \longrightarrow MnCl_2 + KCl + Cl_2 + \mathbf{4}\ H_2O$$

Let's choose to balance **O** first to demonstrate how we produce the first pair of coefficients:

- To balance the first element, use its subscript in the product as coefficient for the reactant in which it occurs, and
- Use its subscript in the reactant as coefficient for the product in which it occurs
- Note: We explicitly write the coefficient even if it is 1 in order to mark the substance as "done" and distinguish it from those that did not receive a coefficient yet ("coefficientless").

$$\mathbf{1}\ KMnO_4 + \quad HCl \longrightarrow \quad MnCl_2 + \quad KCl + \quad Cl_2 \quad + \mathbf{4}\ H_2O$$

After the first step, we find and balance elements that occur in only one *coefficientless* **compound** (i.e. can't be an elemental substance like Cl_2) in the entire equation.

- K and Mn that occurred in the same compound as O and qualified as the first element to balance are natural candidates for the second step because now $KMnO_4$ has a coefficient, leaving these elements occurring in only one coefficientless compound.
- H also qualifies.
- Cl doesn't qualify because it occurs in multiple compounds, not to mention as an elemental substance (which should be handled last)

$$\mathbf{1}\,KMnO_4 + \quad HCl \longrightarrow \quad MnCl_2 + \mathbf{1}\,KCl + \quad Cl_2 + \mathbf{4}\,H_2O$$

We can choose to balance K, which is pretty trivial here.

Next:

Mn is the obvious choice to balance next, since it qualified to be the first element to balance and was in $KMnO_4$, which gained a coefficient. So now it occurs in only one coefficientless compound, $MnCl_2$.

$$\mathbf{1}\,KMnO_4 + \quad HCl \longrightarrow \mathbf{1}\,MnCl_2 + \mathbf{1}\,KCl + \quad Cl_2 + \mathbf{4}\,H_2O$$

Balancing Mn is again pretty trivial.

Next:

Now H is the only element that occurs in only one coefficientless compound.

Balancing reactions, Step 3 (cont.)

$$1\,KMnO_4 + 8\,HCl \longrightarrow 1\,MnCl_2 + 1\,KCl + Cl_2 + 4\,H_2O$$

We now balance H: $x\,(1) = (4)(2) \Rightarrow x = 8$

Next:

The only element that hasn't been balanced is Cl

Balancing reactions, Step 5

$$1\,KMnO_4 + 8\,HCl \longrightarrow 1\,MnCl_2 + 1\,KCl + \frac{5}{2}Cl_2 + 4\,H_2O$$

We now balance Cl: $(8)(1) = (1)(2) + (1)(1) + \mathbf{x}(2) \Rightarrow \mathbf{x} = 5/2$

The equation is now balanced, but we normally want integer coefficients

Balancing reactions, Step 6

$$\mathbf{1}\,KMnO_4 + \mathbf{8}\,HCl \longrightarrow \mathbf{1}\,MnCl_2 + \mathbf{1}\,KCl + \frac{\mathbf{5}}{\mathbf{2}}Cl_2 + \mathbf{4}\,H_2O$$

We now multiply all coefficients by 2 (i.e. the denominator of 5/2)

$$\mathbf{2}\,KMnO_4 + \mathbf{16}\,HCl \longrightarrow \mathbf{2}\,MnCl_2 + \mathbf{2}\,KCl + \mathbf{5}\,Cl_2 + \mathbf{8}\,H_2O$$

The equation is now balanced with the smallest integer coefficients

Balancing an equation might need a bit of math in addition to the rules

We start balancing the following example in the usual way

$$Cu + H_2SO_4 + \longrightarrow CuSO_4 + SO_2 + H_2O$$

Find element(s) that occur in one **compound** on each side:

H is the only element that qualifies

Balance **H**

$$Cu + \mathbf{1}H_2SO_4 \longrightarrow CuSO_4 + SO_2 + \mathbf{1}H_2O$$

Again, note: We explicitly write the coefficient even if it is 1 in order to mark the substance as “done” and distinguish it from those that did not receive a coefficient yet (“coefficientless”).

$$Cu + 1H_2SO_4 \longrightarrow xCuSO_4 + ySO_2 + 1H_2O$$

After the first step, we find and balance elements that <u>occur in only one *coefficientless* **compound**</u>

- No elements occur in only one coefficientless compound
- We assign x, y, etc. as coefficients to the remaining compounds
- Then solve for the unknowns x, y, etc. by setting up the balancing equations for the remaining elements occurring only in compounds

Balance S: $(1)(1) = x\,(1) + y\,(1) \Rightarrow x = 1 - y$

Balance O: $(1)(4) = x\,(4) + y\,(2) + (1)(1) \Rightarrow 4 = 4x + 2y + 1$

$\Rightarrow 4 = \boxed{4-4y}+2y+1 \Rightarrow 4 = 5-2y \Rightarrow \boxed{y = ½} \Rightarrow \boxed{x = 1- ½ = ½}$

Substituted for x

$$Cu + 1H_2SO_4 \longrightarrow \tfrac{1}{2}CuSO_4 + \tfrac{1}{2}SO_2 + 1H_2O$$

Get rid of fractional coefficients by multiplying all **known** coefficients by 2 (the denominator of 2)

$$Cu + 2H_2SO_4 \longrightarrow 1CuSO_4 + 1SO_2 + 2H_2O$$

$$Cu + \mathbf{2}H_2SO_4 \rightarrow \mathbf{1}CuSO_4 + \mathbf{1}SO_2 + \mathbf{2}H_2O$$

Now that we are done with balancing elements that occur in compounds only, we balance the elemental substance

$$\mathbf{x}\ Cu + \mathbf{2}H_2SO_4 \rightarrow \mathbf{1}CuSO_4 + \mathbf{1}SO_2 + \mathbf{2}H_2O$$

$$\mathbf{1}\ Cu + \mathbf{2}H_2SO_4 \rightarrow \mathbf{1}CuSO_4 + \mathbf{1}SO_2 + \mathbf{2}H_2O$$

Now the entire equation is balanced.
We normally omit coefficients equal to 1

$$\boxed{Cu + \mathbf{2}H_2SO_4 \rightarrow CuSO_4 + SO_2 + \mathbf{2}H_2O}$$

Balancing combustion reactions

$$C_cH_hO_o + \ x\ O_2 \longrightarrow \ c\ CO_2 + \ \mathbf{h/2}\ H_2O$$

Combustion:

A "fuel" containing C, H, O reacting with O_2 to produce CO_2 and H_2O

- We don't need to use the general rules
- We "follow" and balance the carbon and hydrogen:
  - All the C in CO_2 comes from the C in the fuel
  - All the H in H_2O comes from the H in the fuel
- Oxygen is not so simple, but becomes simple to balance after we balance C and H

$x = [c + (h/2 - o)/2]$ **but there is no need memorize that**

$$1\,C_2H_6O + \quad O_2 \longrightarrow 2\,CO_2 + 3\,H_2O$$

We start with **one** molecule of "fuel" (C_2H_6O here)

Balance carbon and hydrogen first:

- **Number of CO_2 molecules** equals the number of C atoms in the fuel molecule (each CO_2 molecule has one C atom)

$$(1)(2) = x\,(1) \Rightarrow x = 2$$

- **Number of H_2O molecules** is half the number of H atoms in the fuel molecule (each H_2O molecule has 2 H atoms)

$$(1)(6) = x\,(2) \Rightarrow x = 3$$

- Then balance oxygen

We now balance O:

$$C_2H_6O + \; 3\,O_2 \longrightarrow 2\,CO_2 + 3\,H_2O$$

$$(1)(1) + \mathbf{x}\,(2) = (2)(2) + (3)(1) \quad \Rightarrow \quad \mathbf{x} = 3$$

Concept question

Which of the following is/are true concerning balanced chemical equations? There may be more than one true statement.

I. The number of molecules is conserved.

II. The coefficients tell you how much of each substance you have.

III. Atoms are neither created nor destroyed.

IV. The coefficients indicate the mass ratios of the substances used.

V. The sum of the coefficients on the reactant side equals the sum of the coefficients on the product side.

Stoichiometric Calculations:
About amounts of reactants and products

Note:

- The number of atoms of each type of element must be the same on both sides of a balanced equation.
- Subscripts must not be changed to balance an equation.
- A balanced equation tells us the ratio of the number of molecules which react and are produced in a chemical reaction.
- Coefficients can be fractions, although they are usually given as smallest possible integers.

Stoichiometric Calculations:
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Practice

When the equation

$$NH_3 + O_2 \longrightarrow NO + H_2O$$

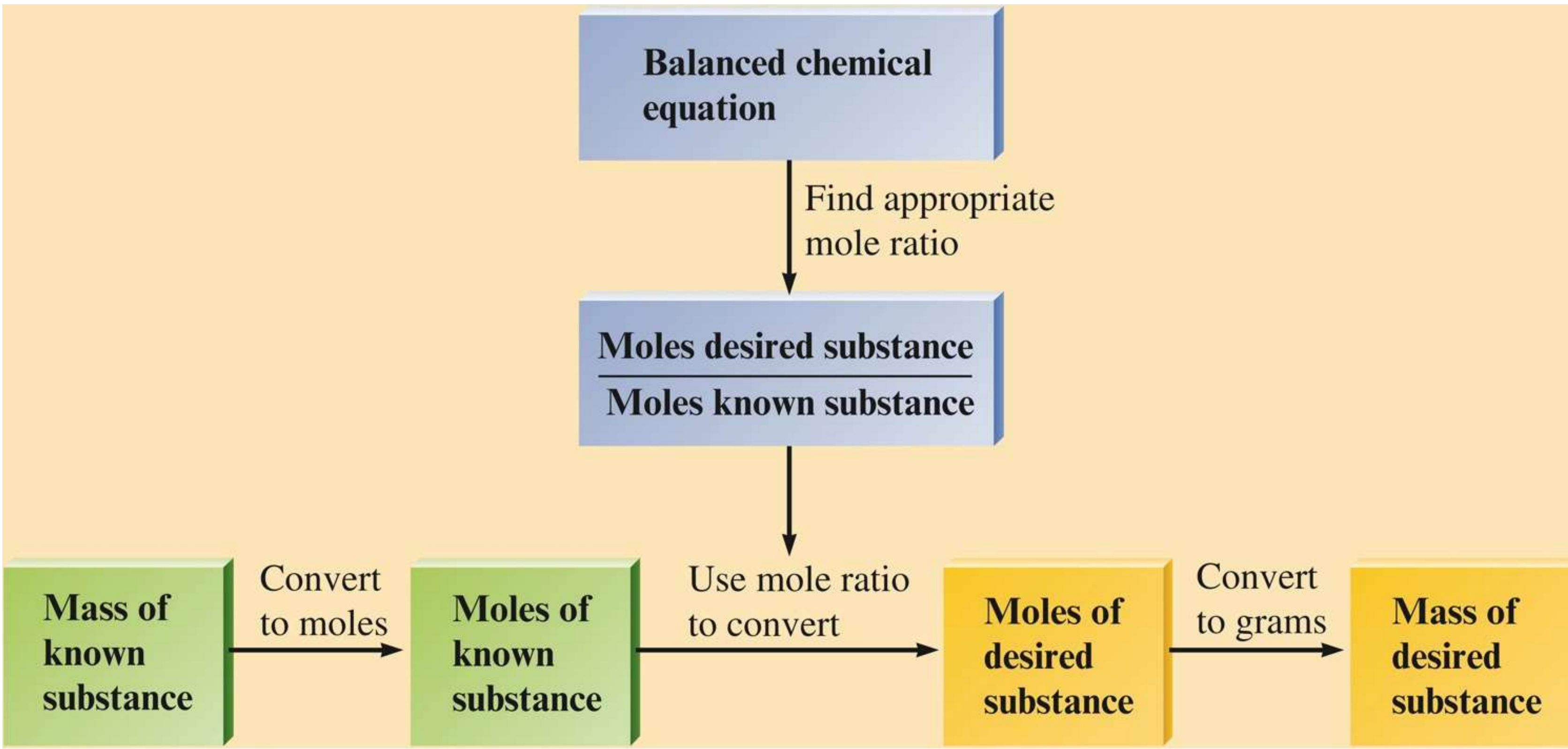
is balanced with the smallest set of integers, the sum of the coefficients is

a) 4
b) 12
c) 14
d) 19
e) 24

Calculating Masses of Reactants and Products in Reactions

1. Balance the equation for the reaction.
2. Convert the known mass of the reactant or product to moles of that substance.
3. Use the balanced equation to set up the appropriate mole ratios.
4. Use the appropriate mole ratios to calculate the number of moles of the desired reactant or product.
   - They are the “conversion factors” between moles of different substances
5. Convert from moles back to grams if required by the problem.

Calculating Masses of Reactants and Products in Reactions



Example

Consider the following reaction:

$$P_4(s) + 5\ O_2(g) \rightarrow 2\ P_2O_5(s)$$

If 6.25 g of phosphorus is burned, what mass of oxygen does it combine with?

$$6.25\ \cancel{g\ P_4} \times \frac{1\ \cancel{mol\ P_4}}{123.88\ \cancel{g\ P_4}} \times \frac{5\ \cancel{mol\ O_2}}{1\ \cancel{mol\ P_4}} \times \frac{32.00\ g\ O_2}{1\ \cancel{mol\ O_2}} = \mathbf{8.07\ g\ O_2}$$

Practice (Part I)

Methane (CH_4) reacts with the oxygen in the air to produce carbon dioxide and water.

Ammonia (NH_3) reacts with the oxygen in the air to produce nitrogen monoxide and water.

Write balanced equations for each of these reactions.

Practice (Part II)

$CH_4 + 2\ O_2 \longrightarrow CO_2 + 2\ H_2O$

$4NH_3 + 5\ O_2 \longrightarrow 4\ NO + 6\ H_2O$

What mass of ammonia would produce the **same amount of water** as 1.00 g of methane reacting with excess oxygen?

1. Find amount of water produced from 1.00 g of methane

mass of CH_4 $\xrightarrow{\text{molar mass}}$ moles of CH_4 $\xrightarrow{\text{rxn. coeff.}}$ moles of water

2. Find amount of ammonia needed to produce the amount of water calculated above.

moles of water $\xrightarrow{\text{rxn. coeff.}}$ moles of NH_3 $\xrightarrow{\text{molar mass}}$ mass of NH_3

Practice (Part II)

$CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$

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

What mass of ammonia would produce the **same amount** of water as 1.00 g of methane reacting with excess oxygen?

$$1.00\ \cancel{g\ CH_4} \times \frac{1\ \cancel{mol\ CH_4}}{16.04\ \cancel{g\ CH_4}} \times \frac{2\ mol\ H_2O}{1\ \cancel{mol\ CH_4}} = 0.1247\ mol\ H_2O$$

2 mol H_2O from 1 mol CH_4

We now calculate the amount of NH_3 needed to produce the same amount of water.

Starting the calculation from the amount of water.

$$0.1247\ \cancel{mol\ H_2O} \times \frac{4\ \cancel{mol\ NH_3}}{6\ \cancel{mol\ H_2O}} \times \frac{17.03\ g\ NH_3}{1\ \cancel{mol\ NH_3}} = \mathbf{1.42\ g\ NH_3}$$

4 mol NH_3 gives 6 mol H_2O

To solve problems:

- You need to connect what's given to what's asked
- Often there is a chain of connections
- You just need to construct that chain
    - You must figure out what piece of information you need for each link in the chain

You need to practice solving many problems

There is no way around it!

Limiting Reactant

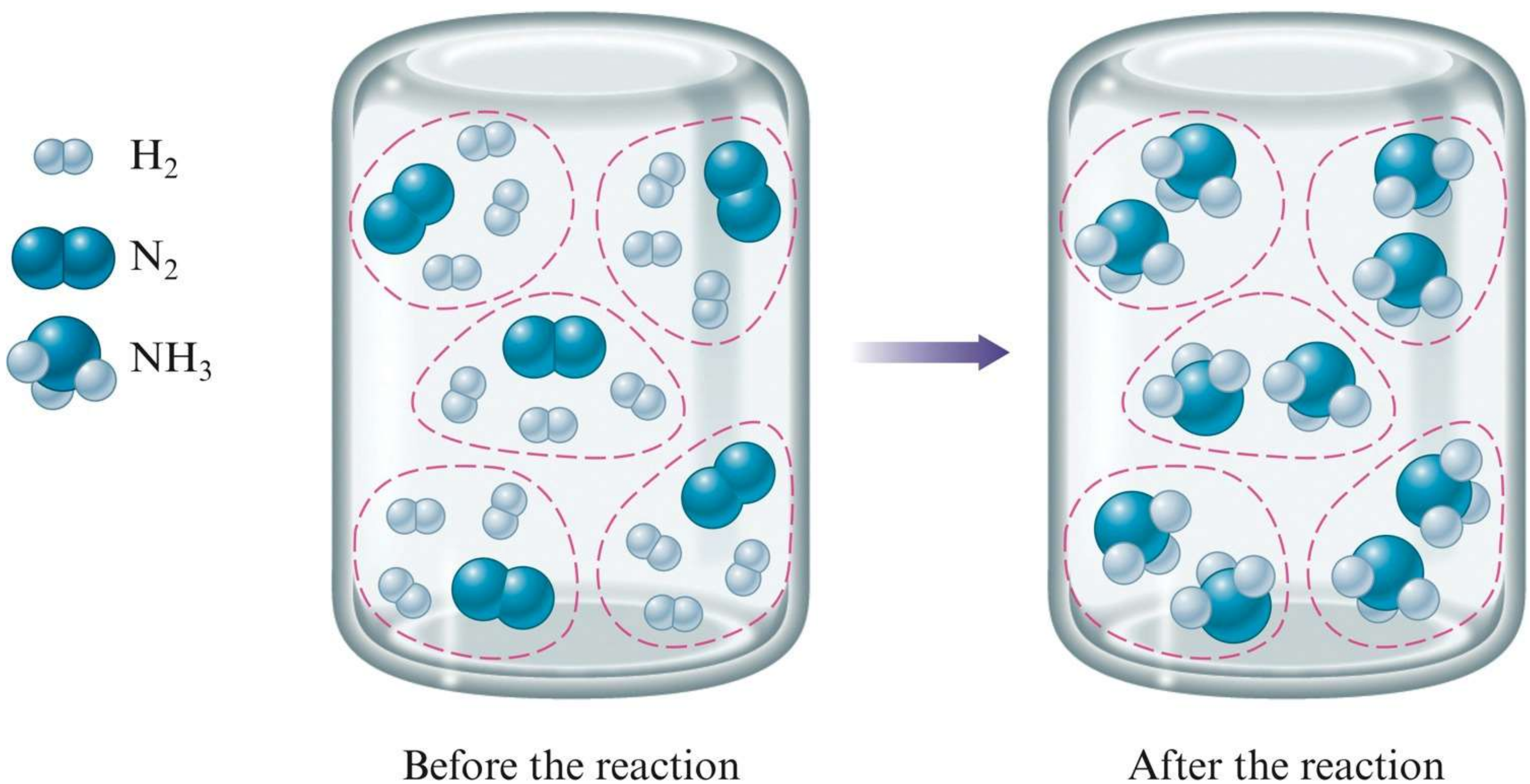
If the reactants are not present in ratios exactly in line with the chemical equation, one of them will be limiting, and the others will automatically be in excess.

Limiting Reactant

- The reactant that runs out first and thus limits the amounts of products that can be formed.
- Other reactants would lead to more product, but they can't because the limiting reactant runs out
- If we calculate the amount of product each reactant would produce if it was completely consumed, limiting reactant would correspond to the least product.
  - **Maximum amount of product the reaction can produce is the amount that can be produced by the limiting reactant**

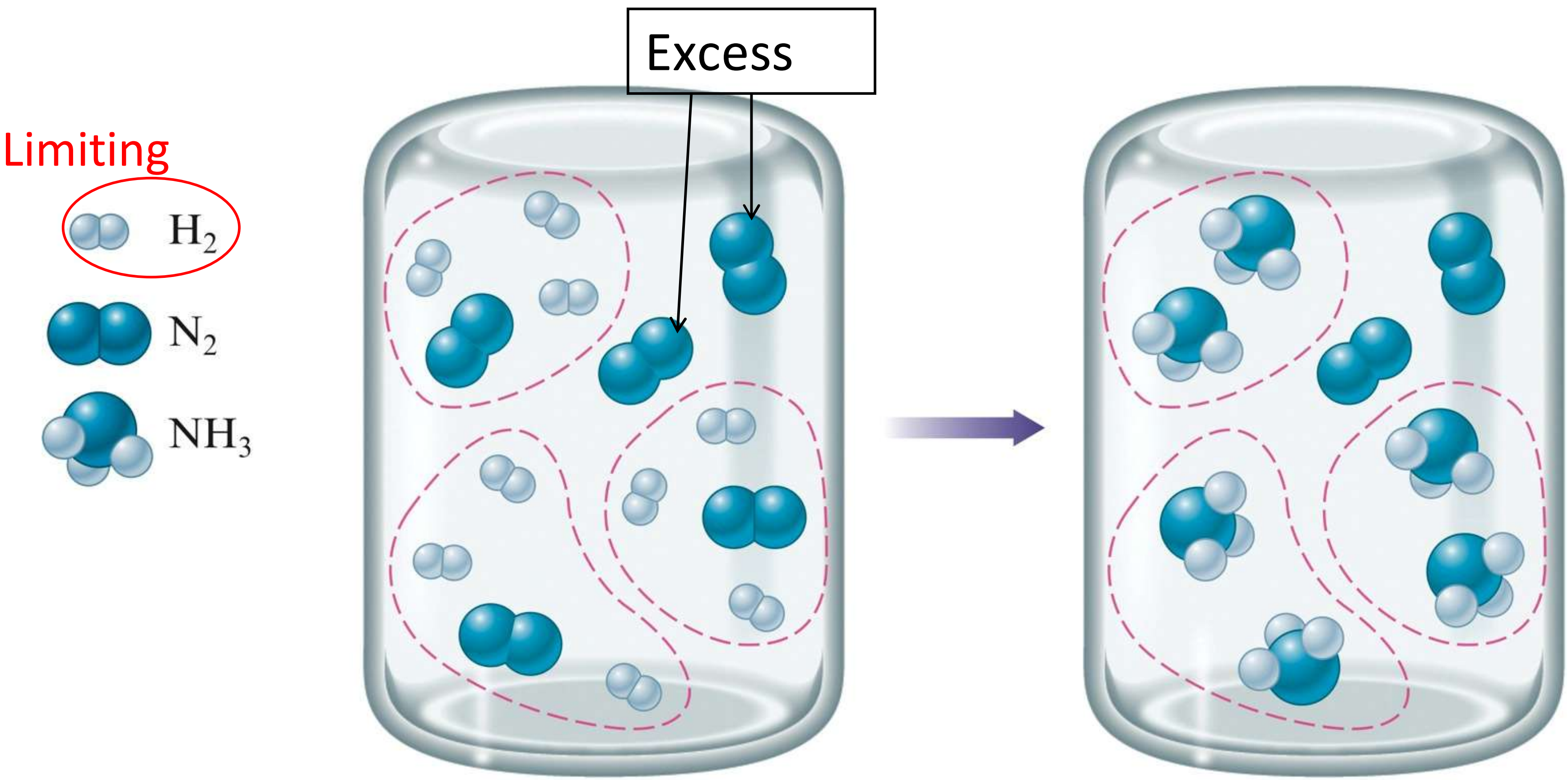
$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

Stoichiometric mixture: No excess or limiting reactants



$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

Limiting reactant mixture



Before the reaction

After the reaction

Concept question

Which of the following reaction mixtures could produce the greatest amount of product? Each involves the reaction symbolized by the equation:

$$2H_2 + O_2 \longrightarrow 2H_2O$$

a) 4 moles of H_2 and 4 moles of O_2

b) 4 moles of H_2 and 6 moles of O_2

c) 4 moles of H_2 and 2 mole of O_2

d) 6 moles of H_2 and 2 moles of O_2

e) They all produce the same amount of product.

Note:

We cannot simply add the total moles of all the reactants to decide which reactant mixture makes the most product.

Amount of product depends on the amount of limiting reactant only

Finding the limiting reactant doesn't actually require calculating the amount of product!

$$a\,\mathrm{A} + b\,\mathrm{B} \longrightarrow c\,\mathrm{C}$$

moles of **C** from **A** = (moles of **A**) $\dfrac{c}{a}$ ← *"a moles of A gives c moles of C"*

Remember your basic algebra $x\dfrac{y}{z} = \dfrac{x}{z}y$

$$\text{moles of } \mathbf{C} \text{ from } \mathbf{A} = \frac{\textbf{(moles of A)}}{a}\, c$$

$$\text{moles of } \mathbf{C} \text{ from } \mathbf{B} = \frac{\textbf{(moles of B)}}{b}\, c$$

The coefficient of the product affects the product amount equally for all reactants

It's the ratio (moles of reactant)/(coefficient of reactant) that determines which reactant would produce the least product

The reactant with the smallest value for the ratio

$$\frac{\text{no. of moles of reactant}}{\text{coefficient of reactant in reaction equation}}$$

is the limiting reactant

When "Reactants are present in *stoichiometric* amounts"

- The mole amounts of all reactants are in the ratios dictated by the coefficients in the balanced reaction equation

- All reactants have the same ratio

$$\frac{\text{moles available}}{\text{coefficient in equation}}$$

- No limiting reactant

- Another way to look at it: all reactants are limiting

If the reaction equation is

$$2A + 3B \longrightarrow 2C$$

and if we have

	A		B	
2 moles of A, 3 moles of B	$\frac{2}{2} = \mathbf{1}$ (*moles* over *coefficient*)	**=**	$\frac{3}{3} = \mathbf{1}$	***A and B in stoichiometric amounts***
3 moles of A, 4.5 moles of B	$\frac{3}{2} = \mathbf{1.5}$	**=**	$\frac{4.5}{3} = \mathbf{1.5}$	
20 moles of A, 30 moles of B	$\frac{20}{2} = \mathbf{10}$	**=**	$\frac{30}{3} = \mathbf{10}$	

But if we have

	A		B	
2 moles of A, 2 moles of B	$\frac{2}{2} = \mathbf{1}$	**>**	$\frac{2}{3} = \mathbf{0.67}$	***B is limiting***
2 moles of A, 3.5 moles of B	$\frac{2}{2} = \mathbf{1}$	**<**	$\frac{3.5}{3} = \mathbf{1.17}$	***A is limiting***

Concept question

The limiting reactant in a reaction

a) has the smallest coefficient in a balanced equation.

b) is the reactant for which you have the fewest number of moles.

c) has the lowest ratio of $\dfrac{\text{moles available}}{\text{coefficient in balanced equation}}$

d) none of these

e) all of these

Practice

You react 10.0 g of **A** with 10.0 g of **B** according to the reaction equation

$$A + 3B \longrightarrow 2C + D$$

What mass of product **C** will be produced given that the molar masses of A, B, and C are 10.0 g/mol, 20.0 g/mol, and 30.0 g/mol, respectively?

Steps:

1. Convert reactant masses to moles
2. Find limiting reactant
3. Find moles of product formed by the limiting reactant
4. Convert moles of product to mass of product

Practice (cont.)

$$\mathbf{A} + \mathbf{3B} \longrightarrow \mathbf{2C} + \mathbf{D}$$

10.0 g 10.0 g ? g

m.m.(A) = 10.0 g
m.m.(B) = 20.0 g

First, convert reactant masses to moles
Then find the limiting reactant

Divide moles by coefficient

A: $10.0\ \cancel{\text{g A}} \times \dfrac{1\ \cancel{\text{mol A}}}{10.0\ \cancel{\text{g A}}} = 1.00\ \text{mol A}$ ⟶ $\dfrac{1.00}{1} = 1.00$

B: $10.0\ \cancel{\text{g B}} \times \dfrac{1\ \cancel{\text{mol B}}}{20.0\ \cancel{\text{g B}}} = 0.500\ \text{mol B}$ ⟶ $\dfrac{0.500}{3} = 0.167$ **Smaller**

⟶ B is limiting

Practice (cont.)

$$\underset{10.0\ g}{A} \ + \ \underset{10.0\ g}{3B} \ \longrightarrow \ \underset{?\ g}{2C} \ + \ D$$

m.m.(A) = 10.0 g
m.m.(B) = 20.0 g
m.m.(C) = 30.0 g

- We can calculate the moles of C produced, using the coefficients in the reaction equation
- We can continue to calculate the mass of C if we know its molar mass

Moles of the limiting reactant

$$0.500 \text{ mol B} \times \frac{2 \text{ mol C}}{3 \text{ mol B}} \times \frac{30.0 \text{ g C}}{1 \text{ mol C}} = 10.0 \text{ g C produced}$$

according to the reaction equation

Practice (cont.)

What if we wanted to know how much is left of each reactant after the reaction is complete?

$$\mathbf{A} \;+\; \mathbf{3B} \;\longrightarrow\; \mathbf{2C} \;+\; \mathbf{D}$$

	A	3B	2C
Before reaction:	10.0 g	10.0 g	0 g
After reaction:	? g	0 g	10.0 g

limiting reactant is consumed completely

To find the amount of remaining excess reactant, we first find the amount consumed, again based on the amount of limiting reactant

$$0.500 \text{ mol B} \times \frac{1 \text{ mol A}}{3 \text{ mol B}} \times \frac{10.0 \text{ g A}}{1 \text{ mol A}} = 1.67 \text{ g of A is consumed}$$

$$\text{remaining A} = (\text{Initial A}) - (\text{A consumed})$$

$$= 10.0 \text{ g} - 1.67 \text{ g} = 8.33 \text{ g of A remaining}$$

Percent Yield

An important indicator of the efficiency of a particular laboratory or industrial reaction.

$$\text{percent yield} = \frac{\text{Actual amount of product}}{\text{Ideal amount of product}} \times 100\%$$

Ideal (theoretical) yield:

The amount of product calculated according to the desired chemical equation and the limiting reactant amount.

For calculations, we typically work with the fractional (out of 1) form of percent yield (e.g. 0.93 instead of 93%).

Algebraic rearrangement gives:

$$\text{Fractional yield} = \frac{\text{Actual amount of product}}{\text{Ideal amount of product}}$$

$$\text{Ideal amount of product} = \frac{\text{Actual amount of product}}{\text{Fractional yield}}$$

$$\text{Actual amount of product} = (\text{Fractional yield})\ (\text{Ideal amount of product})$$

Why would we get less than 100% yield?

Some possible reasons:

- Reaction may not go to completion in the time allowed
- Reactants may undergo different reactions, producing other, unwanted products
- Some product may be lost while being recovered from the reaction mixture

Remember:

- A reaction equation can only be used to relate “ideal” quantities of products and reactants
- The product amount calculated from the reaction equation alone is the “ideal” amount. We then calculate the actual yield using the % yield
- An “actual” yield needs to be converted to the “ideal” yield (using the % yield) before we can work backward to calculate the required reactant amount

- The coefficients (*n, m, i*, etc.) in a reaction equation are "ideal"
- They relate moles of reactants and products in an "ideal" reaction

We assign the loss here for computational purposes

Ideal reaction

$$\boxed{nA + mB \longrightarrow iC} \xrightarrow{\text{multiply by \% yield}} \boxed{rC}$$

$$\boxed{nA + mB \longrightarrow iC} \xleftarrow{\text{divide by \% yield}} \boxed{rC}$$

Real amounts of reactants

ideal yield
(maximum)

Losses occurring in reality

Real yield
(smaller than ideal)

To find reactant amount(s) needed:

- go from <u>real to ideal</u> yield first
- use the ratios in the ideal reaction as usual

Practice

For the following reaction:

$$P_4(s) + 6F_2(g) \longrightarrow 4PF_3(g)$$

What mass of P_4 is needed to produce 85.0 g of PF_3 if the reaction has a 64.9% yield?

Steps:

1. $P_4(s)$ is implied to be the limiting reactant
2. 85.0 g is the "actual amount of product" (actual yield)
3. We know the % yield, so we can find the "ideal" yield
4. Ideal yield comes from the limiting reactant amount and reaction equation
5. Work back to the amount of $P_4(s)$

Practice (cont.)

$$P_4(s) \ + \ 6\,F_2(g) \longrightarrow 4\,PF_3(g)$$

Real yield
85.0 g
@ 64.9% yield

$\div 0.649$

$13\underline{0}.97$ g

Ideal yield
(Can be used to calculate the reactant amounts)

Ideal yield
Calculated from actual yield

$$13\underline{0}.97 \text{ g PF}_3 \times \frac{1 \text{ mol PF}_3}{87.97 \text{ g PF}_3} \times \frac{1 \text{ mol P}_4}{4 \text{ mol PF}_3} \times \frac{123.88 \text{ g P}_4}{1 \text{ mol P}_4} = 46.1 \text{ g of P}_4 \text{ needed}$$

For the actual yield

Applying stoichiometry to reactions in solutions

A huge amount of chemistry, probably the vast majority of it, actually happens in solutions.

The most practical way to measure quantities of solutions is measuring volumes.

The concept of concentration relates the volume of a solution to the quantity of a substance dissolved in it.

A very common way to represent concentration is “molarity”, which is number of moles of the substance per liter of solution volume.

Concentration of Solutions: **Molarity**

Molarity (*c*) = moles of solute per 1 liter of solution

$$\boldsymbol{c = \textbf{Molarity} = \frac{\textbf{moles of solute}}{\textbf{volume of solution (in L)}} = \frac{\textbf{n}}{\textbf{V}}}$$

Unit of Molarity: "Molar" (*M*)

If 2.0 liters of a solution contains 6.0 moles of solute, its concentration is:

$$\frac{6.0\ \text{mol}}{2.0\ \text{L}} = 3.0\ \text{molar} \quad \text{or} \quad 3.0\ M$$

Example

A 500.0-g sample of potassium phosphate is dissolved in enough water to make 1.50 L of solution. What is the molarity of the solution?

Steps:

1. Convert mass to moles, using molar mass
2. Use formula for molarity

Molar mass of K_3PO_4 = (3)(39.10) + (1)(30.97) + (4)(16.00) = 212.27 g/mol
(m.m.)

$$\text{Number of moles} = n = 500.0\text{ g} \times \frac{1\text{ mol}}{212.27\text{ g}} = 2.36\text{ mol}$$

$$\text{Molarity} = c = \frac{n}{V} = \frac{2.36\text{ mol}}{1.50\text{ L}} = 1.57\text{ M (molar)}$$

Concentration of Ions
-- given the concentration of an ionic compound

- Multiply the concentration of the compound by the number of ions in the formula.
- Applies to cations, anions, i.e. all ions

For a 0.25 *M* $CaCl_2$ solution:

$$CaCl_{2\,(aq)} \longrightarrow Ca^{2+}{}_{(aq)} + 2Cl^{-}{}_{(aq)}$$

Molarities of Ca^{2+} and Cl^{-}:

$$c_{Ca^{2+}} = 1 \times 0.25\,M = 0.25\,M$$

$$c_{Cl^{-}} = 2 \times 0.25\,M = 0.50\,M$$

$$c_{ions} = (1+2) \times 0.25\,M = 0.75\,M$$

Example

Which of the following solutions has the greatest number of ions?

a) 400.0 mL of 0.10 M NaCl

$$c_{\text{ions}} = (1+1)(0.10\ M) = 0.20\ M$$

$$n_{\text{ions}} = c_{\text{ions}} V = (0.20)\ (400.0\text{x}10^{-3}) = 0.080 \text{ mol}$$

b) 300.0 mL of 0.10 M $CaCl_2$

$$c_{\text{ions}} = (1+2)(0.10\ M) = 0.30\ M$$

$$n_{\text{ions}} = c_{\text{ions}} V = (0.30)\ (300.0\text{x}10^{-3}) = \boxed{0.090 \text{ mol}}$$

c) 200.0 mL of 0.10 M $FeCl_3$

$$c_{\text{ions}} = (1+3)(0.10\ M) = 0.40\ M$$

$$n_{\text{ions}} = c_{\text{ions}} V = (0.40)\ (200.0\text{x}10^{-3}) = 0.080 \text{ mol}$$

d) 800.0 mL of 0.10 M sucrose

$c_{\text{ions}} = 0$ mol Not an electrolyte!

Note that finding the greatest number of **moles** is sufficient.

Multiplying each mol by 6.022×10^{23} wouldn't affect the ranking

Number of moles of solute is the product of molarity (*c*) and volume (V)

$$n = c \cdot V$$

Moles of solute

Dilution

The process of adding solvent to a concentrated solution to achieve the molarity desired for a particular solution.

- Dilution with water does not alter the numbers of moles of solute present.

Moles of solute **before** dilution = $n_1 = c_1 \cdot V_1$

Moles of solute **after** dilution = $n_2 = c_2 \cdot V_2$

$$n_1 = n_2 \Longrightarrow \boxed{c_1 \cdot V_1 = c_2 \cdot V_2}$$

Concept Practice

A 0.50 *M* solution of sodium chloride in an open beaker sits on a lab bench. Which of the following would decrease the concentration of the salt solution?

a) Pour some of the solution down the sink drain.

b) Add more sodium chloride to the solution.

c) Let the solution sit in open air for a couple of days.

d) Add water to the solution.

e) At least two of the above would decrease the concentration of the salt solution.

Practice

V_1 c_1

What is the volume of a 2.00 *M* NaOH solution needed to make 150.0 mL of a 0.800 *M* NaOH solution?

V_2 c_2

$$c_1 \; V_1 = c_2 \; V_2$$

If in mL units, then V_1 will also be in mL

$$(2.00\ M)\ V_1 = (0.800\ M)\ (150.0\ \text{mL})$$

$$V_1 = \frac{(0.800\ M)(150.0\ \text{mL})}{(2.00\ M)} = 60.0\ \text{mL}$$