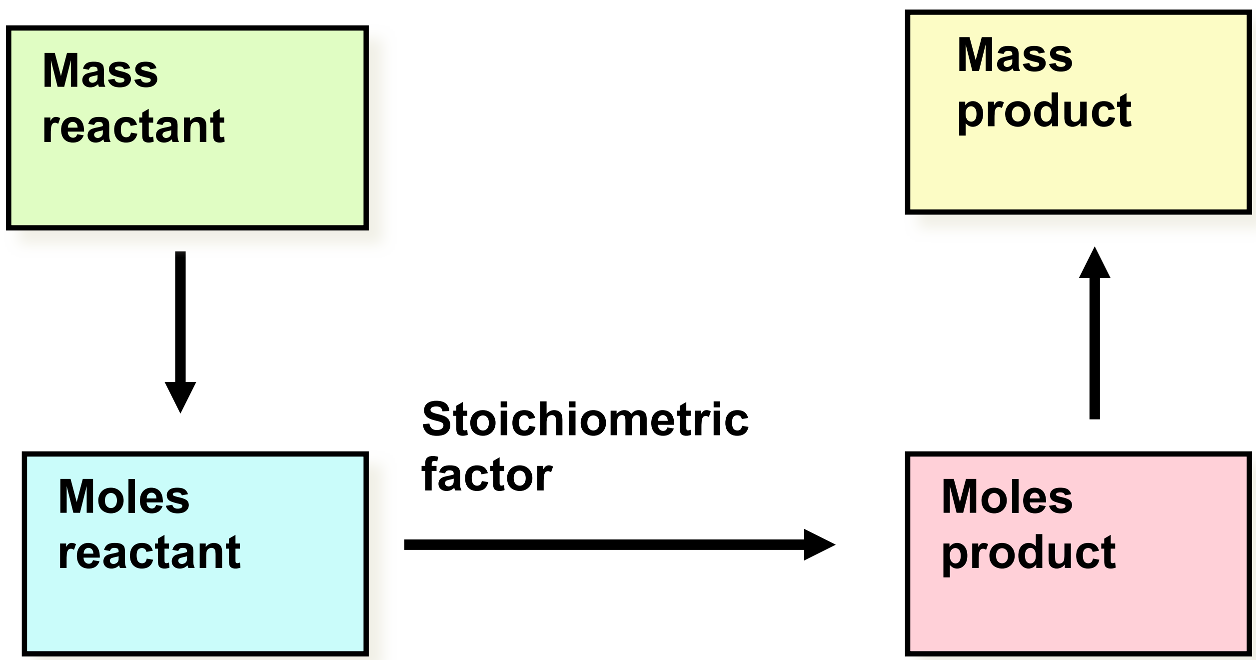
STOICHIOMETRY

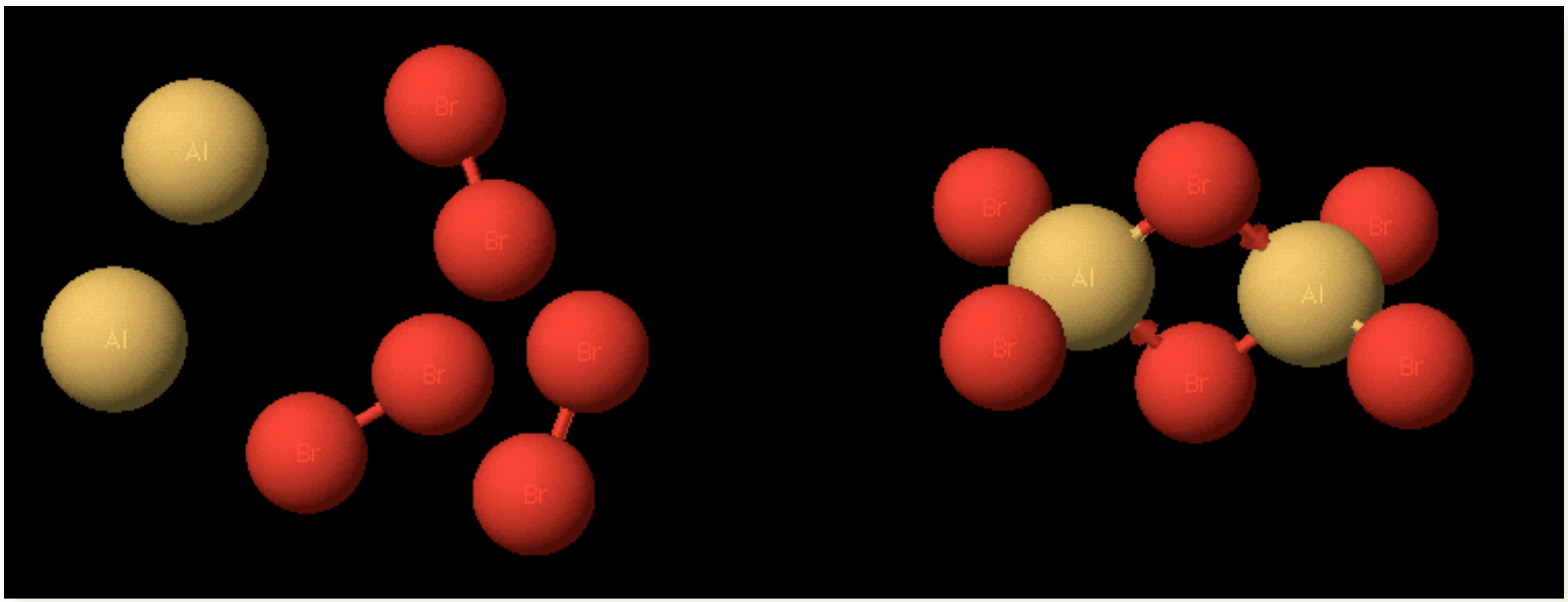
– the study of the quantitative aspects of chemical

GENERAL PLAN FOR STOICHIOMETRY



STOICHIOMETRY

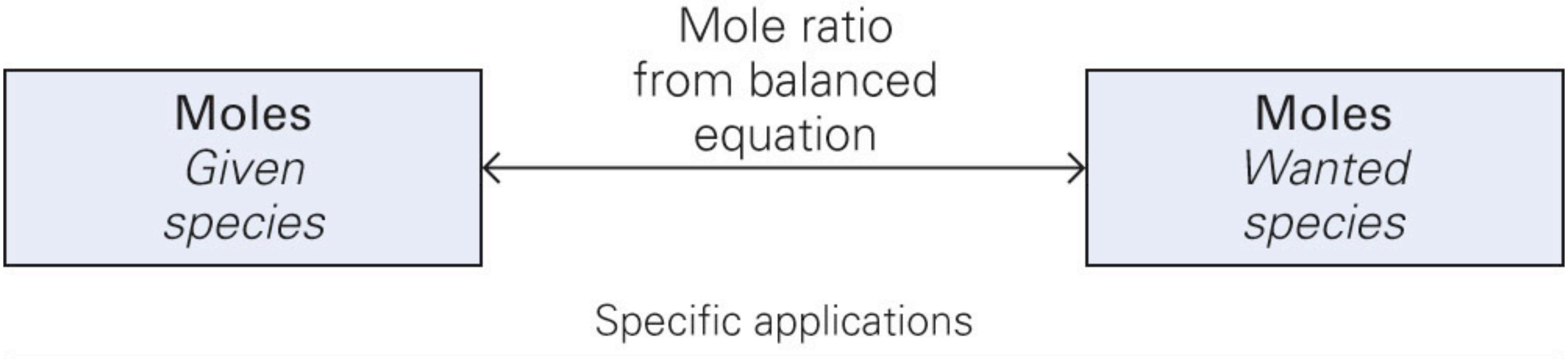
It rests on the principle of the conservation



$2\ Al(s)\ +\ 3\ Br_2(liq)$ f $Al_2Br_6(s)$

Conversion Factors from Eqns

The mole ratio between two species, as given in the balanced chemical

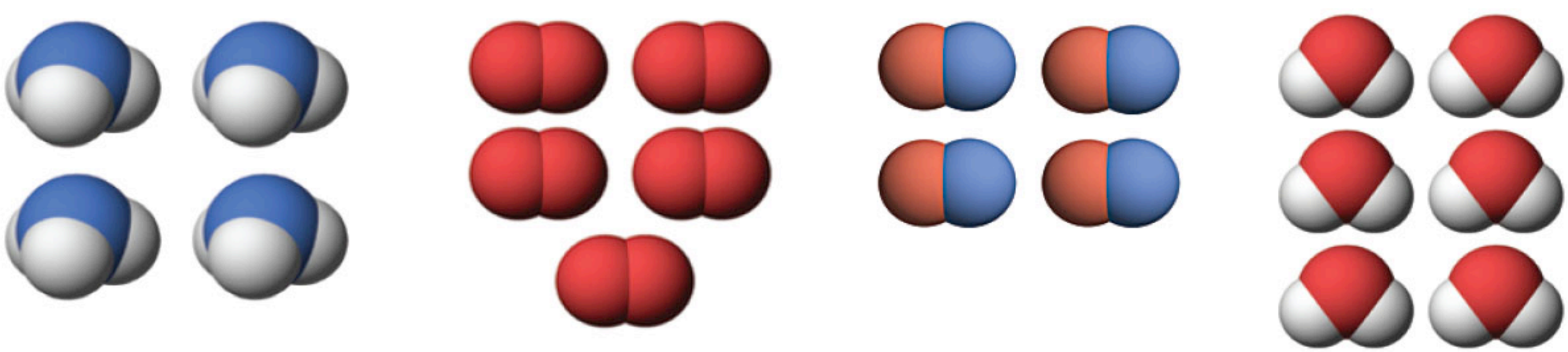


Conversion Factors from Eqns

Example:

How many moles of oxygen are needed to completely react with 2.34 moles of ammonia in a reaction that yields nitrogen monoxide and water?

$$4\,NH_3(g) + 5\,O_2(g) \rightarrow 4\,NO(g) + 6\,H_2O(g)$$



Conversion Factors from Eqns

How many moles of oxygen are needed to completely react with 2.34 moles of ammonia in a reaction that yields nitrogen monoxide and water?

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

$$\frac{5 \text{ mol } O_2}{4 \text{ mol } NH_3}$$

The PER relationship comes from the balanced equation:

Conversion Factors from Eqns

GIVEN: 2.34 mol NH_3, $4\ NH_3 + 5\ O_2 \rightarrow 4\ NO + 6\ H_2O$

WANTED: mol O_2 $\xrightarrow{\frac{5 \text{ mol } O_2}{4 \text{ mol } NH_3}}$

PER:

$$\frac{5 \text{ mol } O_2}{4 \text{ mol } NH_3}$$

Mass–Mass Stoichiometry

Goal 2

Given a chemical equation, or a reaction for which the equation can be written, and the number of grams or moles of one species in the reaction, find the number of grams or moles of any other species.

Mass–Mass Stoichiometry

Stoichiometry
The quantitative relationships among substances
involved in a chemical reaction.

Established by the balanced equation for the reaction.

A stoichiometry problem asks,
"How much or how many?"

Mass-Mass Stoichiometry

Prerequisite Skills for Stoichiometry

Write chemical formulas. Ch. 6

Calculate molar masses from chemical formulas. Sect. 7.4

Use molar masses to convert
between moles and mass. Sect. 7.5

Write and balance chemical equations. Ch. 8

Use an equation to convert from moles

Mass-Mass Stoichiometry

Mass-to-Mass Stoichiometry Path

Mass of → Moles of → Moles → Mass of
of Given → Given → Wanted → Wanted

Molar Mass | Equation | Molar Mass
g PER mol | mol PER mol | g PER mol

$$\text{Mass Given} \times \frac{\text{mol Given}}{\text{mass Given}} \times \frac{\text{mol Wanted}}{\text{mol Given}} \times \frac{\text{mass Wanted}}{\text{mol Wanted}}$$

Mass-Mass Stoichiometry

How to Solve a Stoichiometry Problem: The Stoichiometry Path

Step 1: Change the mass of the given species to moles.

Step 2: Change the moles of the given species to the moles of the wanted species.

Step 3: Change the moles of the wanted species

Mass-Mass Stoichiometry

Example:

How many grams of carbon dioxide are produced when 10.0 g of methane, CH_4, is burned?

Solution:

GIVEN: 10.0 g CH_4 WANTED: g CO_2

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

PER: $\xrightarrow{16.04 \text{ g } CH_4/\text{mol } CH_4}$

PATH: g CH_4 $\xrightarrow{16.04 \text{ g } CH_4/\text{mol } CH_4}$ mol CH_4 $\xrightarrow{1 \text{ mol } CO_2/1 \text{ mol } CH_4}$ $\xrightarrow{44.01 \text{ g } CO_2/\text{mol } CO_2}$

Mass-Mass Stoichiometry

$$10.0\ \cancel{\text{g } CH_4} \times \frac{1\ \cancel{\text{mol } CH_4}}{16.04\ \cancel{\text{g } CH_4}} \times \frac{1\ \cancel{\text{mol } CO_2}}{1\ \cancel{\text{mol } CH_4}} \times \frac{44.01\ \text{g } CO_2}{\cancel{\text{mol } CO_2}}$$

$= 27.4$ g CO_2

Percent Yield

Goal 3

Given two of the following, or information from which two of the following may be determined, calculate the third: theoretical yield, actual yield, percent yield.

Theoretical Yield = Maximum amount of product possible based on amount of reactant(s).

Actual Yield = How much product was actually obtained.

Percent Yield

The **actual yield** of a chemical reaction is usually less than the **theoretical yield** predicted by a stoichiometry calculation because:

- reactants may be impure
- the reaction may not go to completion
- other reactions may occur

Actual yield is experimentally determined.

Percent Yield

Percent yield expresses the ratio of actual yield to theoretical yield:

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

Percent Yield

Example:

Determine the percent yield if 6.97 grams of ammonia is produced from the reaction of 6.22 grams of nitrogen with excess hydrogen.

Solution:

Notice that two quantities are given in the problem statement.

6.97 g NH_3 is the actual yield

6.22 g N_2 is the amount of reactant. This can be used to

Percent Yield

GIVEN: 6.22 g N_2 (theo) WANTED: g NH_3

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

PER: 28.02 g N_2/mol N_2 2 mol NH_3/1 mol N_2

PATH: ~~g N_2~~ mol N_2

Percent Yield

$$6.22\ g\ N_2 \times \frac{1\ mol\ N_2}{28.02\ g\ N_2} \times \frac{2\ mol\ NH_3}{1\ mol\ N_2} \times \frac{17.03\ g\ NH_3}{mol\ NH_3} = 7.56\ g\ NH_3$$

(theo)

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

$$= \frac{6.97\ g\ NH_3\ (act)}{7.56\ g\ NH_3\ (theo)} \times 100 =$$

Percent Yield

Once a percent yield has been determined for a reaction,

it can be used in stoichiometry calculations.

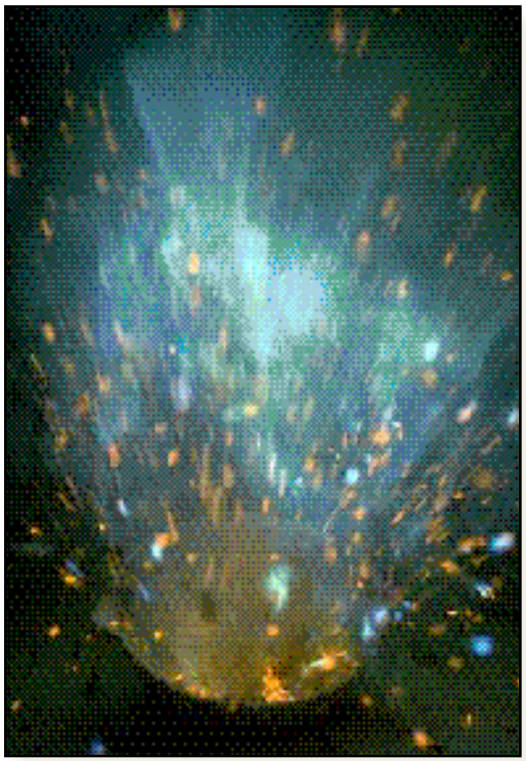
For example, the 92.2% yield from the prior example means:

$$\frac{92.2\ \text{g (act)}}{100\ \text{g (theo)}}$$

and

$$\frac{100\ \text{g (theo)}}{92.2\ \text{g (act)}}$$

PROBLEM: If 454 g of NH_4NO_3 decomposes, how much N_2O and H_2O are formed? What is the theoretical yield of products?



STEP 1

Write the balanced chemical equation

NH_4NO_3(s) f

N_2O(g) + 2 H_2O(g)

454 g of NH_4NO_3 f $N_2O + 2\ H_2O$

STEP 2 Convert mass of reactant (454 g) to amount (mol)

$$454\ \text{g} \left(\frac{1\ \text{mol}}{80.04\ \text{g}}\right) = 5.67\ \text{mol}\ NH_4NO_3$$

STEP 3 Convert amount of reactant (5.67 mol) to amount (mol) of product.

454 g of NH_4NO_3 f $N_2O + 2\ H_2O$

STEP 3 Convert moles reactant f moles product

Relate moles NH_4NO_3 to moles product expected.

1 mol NH_4NO_3 f 2 mol H_2O

Express this relation as the STOICHIOMETRIC FACTOR

$$\frac{2 \text{ mol } H_2O \text{ produced}}{1 \text{ mol } NH_4NO_3 \text{ used}}$$

$454 \text{ g of } NH_4NO_3 \rightarrow N_2O + 2\ H_2O$

STEP 3 Convert moles reactant (5.67 mol) to amount (mol) of product

$$5.67 \text{ mol } NH_4NO_3 \left(\frac{2 \text{ mol } H_2O \text{ produced}}{1 \text{ mol } NH_4NO_3 \text{ used}}\right)$$

$= 11.3 \text{ mol } H_2O \text{ produced}$

454 g of NH_4NO_3 f $N_2O + 2\ H_2O$

STEP 4 Convert amount of product (11.3 mol) to mass of product

Called the THEORETICAL YIELD

$$11.3 \text{ mol } H_2O \left(\frac{18.02 \text{ g}}{1 \text{ mol}} \right) = 204 \text{ g } H_2O$$

ALWAYS FOLLOW THESE STEPS IN SOLVING STOICHIOMETRY PROBLEMS!

454 g of NH_4NO_3 f $N_2O + 2\ H_2O$

STEP 5 What mass of N_2O is formed?

Total mass of reactants = total mass of products

454 g NH_4NO_3 = ___ g N_2O + 204 g H_2O

454 g of NH_4NO_3 f $N_2O + 2\ H_2O$

Amounts Table (from page 159)

1. Compound NH_4NO_3 N_2O H_2O
2. Initial (g)
3. Initial (mol)
4. Change (mol)
5. Final (mol)
6. Final (g)

454 g of NH_4NO_3 f $N_2O + 2\ H_2O$

Amounts Table

1. Compound	NH_4NO_3	N_2O	H_2O
2. Initial (g)	454 g	0	0
3. Initial (mol)	5.67 mol	0	0
4. Change (mol)	−5.67	+5.67	+2(5.67)
5. Final (mol)	0	5.67	11.3
6. Final (g)	0	250	204

Note that matter is conserved!

454 g of NH_4NO_3 f $N_2O + 2\ H_2O$

STEP 6 Calculate the percent yield

If you isolated only 131 g of N_2O, what is the percent yield?

This compares the theoretical (250.

$$454 \text{ g of } NH_4NO_3 \text{ f } N_2O + 2\ H_2O$$

STEP 6 Calculate the percent yield

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

$$\%\ \text{yield} = \left(\frac{131\ \text{g}}{250.\ \text{g}}\right) \times 100\% = 52.4\%$$

PROBLEM: Using 5.00 g of H_2O_2, what mass of O_2 and of H_2O can be obtained?

$2\ H_2O_2(liq)$ f $2\ H_2O(g)\ +\ O_2(g)$

Reaction is catalyzed by MnO_2

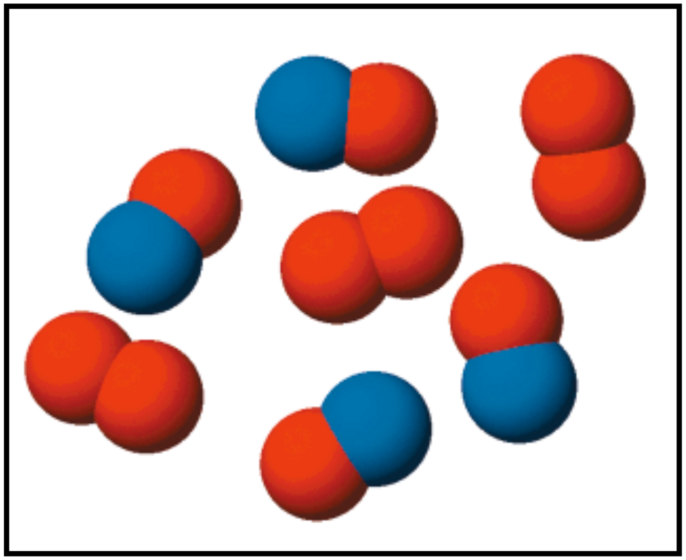
Step 1: amount (mol) of H_2O_2

Step 2: use STOICHIOMETRIC FACTOR to calculate amount (mol) of O_2

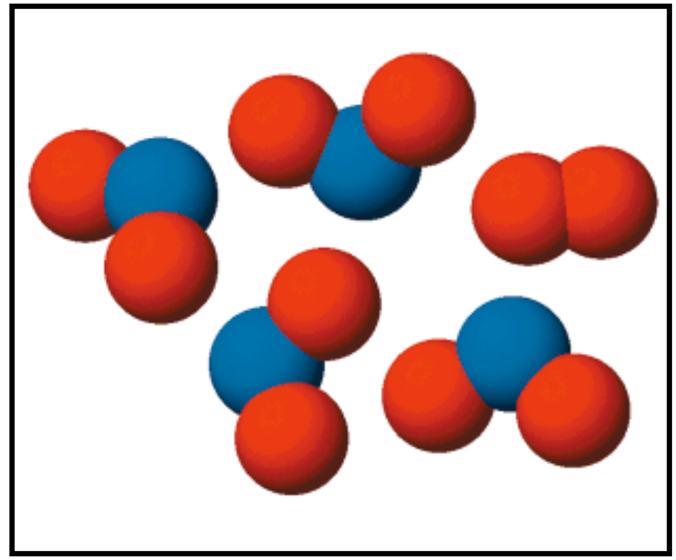
Reactions Involving a LIMITING REACTANT

1. In a given reaction, there is not enough of one reagent to use up the other reagent completely.
2. The reagent in short supply LIMITS the quantity of product that can be formed.

LIMITING REACTANTS



Reactants



Products

$$2\,NO(g) + O_2(g) \rightarrow 2\,NO_2(g)$$

Limiting reactant = ____________

Excess reactant = _____________

LIMITING REACTANTS

1 2 3

React solid Zn with 0.100 mol HCl (aq)

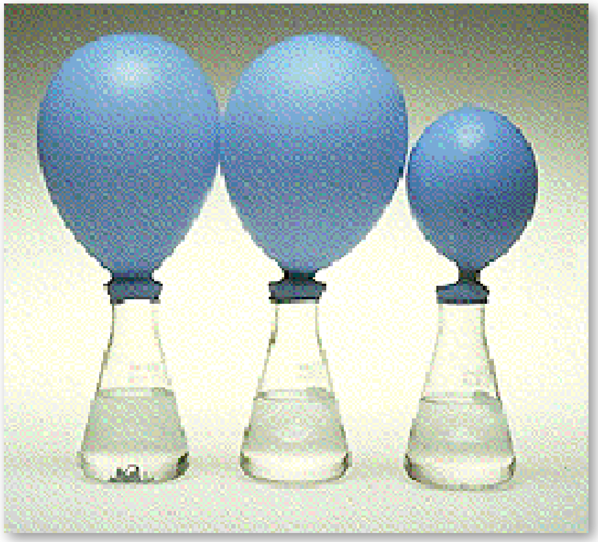
Zn + 2 HCl f $ZnCl_2$ + H_2

Rxn 1: Balloon inflates fully, some Zn left
* More than enough Zn to use up the 0.100 mol HCl

Rxn 2: Balloon inflates fully, no Zn left
* Right amount of each (HCl and Zn)

Rxn 3: Balloon does not inflate fully, no Zn left.

LIMITING REACTANTS



React solid Zn with 0.100 mol HCl (aq)

$$Zn + 2\ HCl\ f\ ZnCl_2 + H_2$$

0.10 mol HCl [1 mol Zn/2 mol HCl]
= 0.050 mol Zn

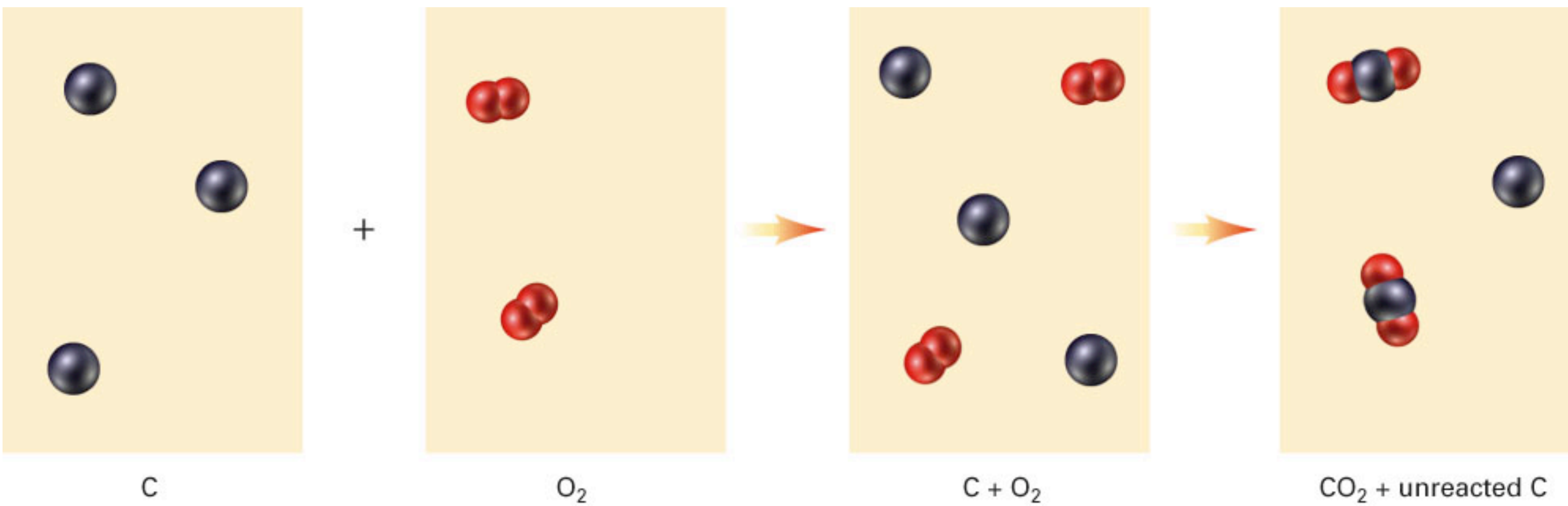
Rxn 1 Rxn
2 Rxn 3
mass Zn (g) 7.00 3.27
1.31
mol Zn 0.107
0.050 0.020

Limiting Reactants

Goal 4

Identify and describe or explain limiting reactants and excess reactants.

Limiting Reactants



Three atoms of carbon react with two molecules of oxygen:

Limiting Reactants

The reaction will stop when two molecules of oxygen are used up.
Two carbon dioxide molecules will form;
One carbon atom will remain unreacted:

	C	+	O_2	$\rightarrow$	CO_2
Start	3		2		0
Used (+) or Produced (–)	– 2		– 2		+ 2

Limiting Reactants

Limiting Reactant

The reactant that is completely used up.

(Oxygen)

Excess Reactant

The reactant initially present in excess.

(some will remain unreacted)

(Carbon)

Limiting Reactants

Goal 5

Given a chemical equation, or information from which it may be determined, and initial quantities of two or more reactants, (a) identify the limiting reactant, (b) calculate the theoretical yield of a specified product, assuming complete use of the limiting reactant, and (c) calculate the quantity of reactant initially in excess that remains unreacted.

Limiting Reactants

Comparison-of-Moles Method
How to Solve a Limiting Reactant Problem:

1. Convert the number of grams of each reactant to moles.
2. Identify the limiting reactant.
3. Calculate the number of moles of each species that reacts or is produced.
4. Calculate the number of moles of each species that remains after the reaction.
5. Change the number of moles of each species to grams.

Limiting Reactants

Example:

A 24.0-g piece of solid sodium is added to 24.0 g of liquid water. How many grams of hydrogen will be produced? How many grams of which reactant will be left over?

Solution:

Step 1 is to convert the number of grams of each reactant to moles.

$$24.0\text{ g Na} \times \frac{1\text{ mol Na}}{22.99\text{ g Na}} = 1.04\text{ mol Na}$$

$$\frac{1\text{ mol } H_2O}{18.02\text{ g } H_2O}$$

Limiting Reactants

Step 2 is to identify the limiting reactant. This requires a balanced chemical equation so that the mole ratio of the reacting substances can be identified.

$2\ Na + 2\ H_2O \rightarrow H_2 + 2\ NaOH$

For every 2 mol of Na that react, 2 mol of H_2O react, a 2:2 or 1:1 ratio. Thus, the 1.04 mol Na will be the limiting reactant.

Limiting Reactants

Step 3 is to calculate the number of moles of each species that reacts or is produced. This can be done by assembling the data into a table.

2 Na + 2 H_2O → H_2 +2
NaOH

Grams at start 24.0 24.0 0 0

Molar mass, g/mol 22.99 18.02
2.016 40.00

Moles at start 1.04 1.33
0 0

Moles used (–), $\frac{1 \text{ mol } H_2}{2 \text{ mol Na}}$

Limiting Reactants

Step 4 is to calculate the number of moles of each species that remains after the reaction.

	$2\ Na$ +	$2\ H_2O$ →	H_2 +	$2\ NaOH$
Grams at start	24.0	24.0	0	0
Molar mass, g/mol	22.99	18.02	2.016	40.00
Moles at start	1.04	1.33	0	

Limiting Reactants

Step 5 is to change the number of moles of each species to grams.

$2\ Na + \quad 2\ H_2O \rightarrow \quad H_2 + \quad 2\ NaOH$

Moles at end $\quad 0 \quad 0.29 \quad 0.52 \quad 1.04$

$$0.29\ \text{mol}\ H_2O \times \frac{18.02\ \text{g}\ H_2O}{\text{mol}\ H_2O} = 5.2\ \text{g}\ H_2O$$

$$0.52\ \text{mol}\ H_2 \times \frac{2.016\ \text{g}\ H_2}{\text{mol}\ H_2} = 1.0\ \text{g}\ H_2$$

$$\frac{40.00\ \text{g NaOH}}{\text{mol NaOH}}$$

Limiting Reactants

Returning to the original problem statement:

A 24.0-g piece of solid sodium is added to 24.0 g of liquid water. How many grams of hydrogen will be produced? How many grams of which reactant will be left over?

The questions can now be answered:

1.0 g H_2 is produced

Limiting Reactants

Goal 5

Given a chemical equation, or information from which it may be determined, and initial quantities of two or more reactants, (a) identify the limiting reactant, (b) calculate the theoretical yield of a specified product, assuming complete use of the limiting reactant, and (c) calculate the quantity of reactant initially in excess that remains unreacted.

Limiting Reactants

Smaller-Amount Method
How to Solve a Limiting Reactant Problem:

1. Calculate the amount of product that can be formed by the initial amount of each reactant.
   a) The reactant that yields the smaller amount of product is the limiting reactant.
   b) The smaller amount of product is the amount that will be formed when all of the limiting reactant is used up.
2. Calculate the amount of excess reactant that is used by the total amount of limiting reactant.
3. Subtract from the amount of excess reactant present initially the amount that is used by all of

Limiting Reactants

Example:

A 24.0-g piece of solid sodium is added to 24.0 g of liquid water. How many grams of hydrogen will be produced? How many grams of which reactant will be left over?

Solution:

Step 1 is to calculate the amount of product that can be formed by the initial amount of each reactant.

Limiting Reactants

GIVEN: 24.0 g Na WANTED: g H_2

$2\ Na + 2\ H_2O \rightarrow H_2 + 2\ NaOH$

$$24.0\ g\ Na \times \frac{1\ mol\ Na}{22.99\ g\ Na} \times \frac{1\ mol\ H_2}{2\ mol\ Na} \times \frac{2.016\ g\ H_2}{mol\ H_2} = 1.05\ g\ H_2$$

GIVEN: 24.0 g H_2O WANTED: g H_2

$$24.0\ g\ H_2O \times \frac{1\ mol\ H_2O}{18.02\ g\ H_2O} \times \frac{1\ mol\ H_2}{2\ mol\ H_2O} \times \frac{2.016\ g\ H_2}{mol\ H_2}$$

Limiting Reactants

24.0 g Na and excess H_2O
produces 1.05 g H_2

24.0 g H_2O and excess Na
produces 1.34 g H_2

Step 1a: The reactant that yields the smallest amount of product is the limiting reactant, so Na is the limiting reactant.

Step 1b: The smaller amount of product is the amount that will be formed when all of the limiting reactant is

Limiting Reactants

Step 2 is to calculate the amount of excess reactant that is used by the total amount of limiting reactant.

The total amount of limiting reactant is 24.0 g Na.

The excess reactant is H_2O.

GIVEN: 24.0 g Na WANTED: g H_2O

$2\ Na + 2\ H_2O \rightarrow H_2 + 2\ NaOH$

$$\frac{1\ \text{mol Na}}{22.99\ \text{g Na}} \quad \frac{2\ \text{mol}\ H_2O}{2\ \text{mol Na}} \quad \frac{18.02\ \text{g}\ H_2O}{\text{mol}\ H_2O}$$

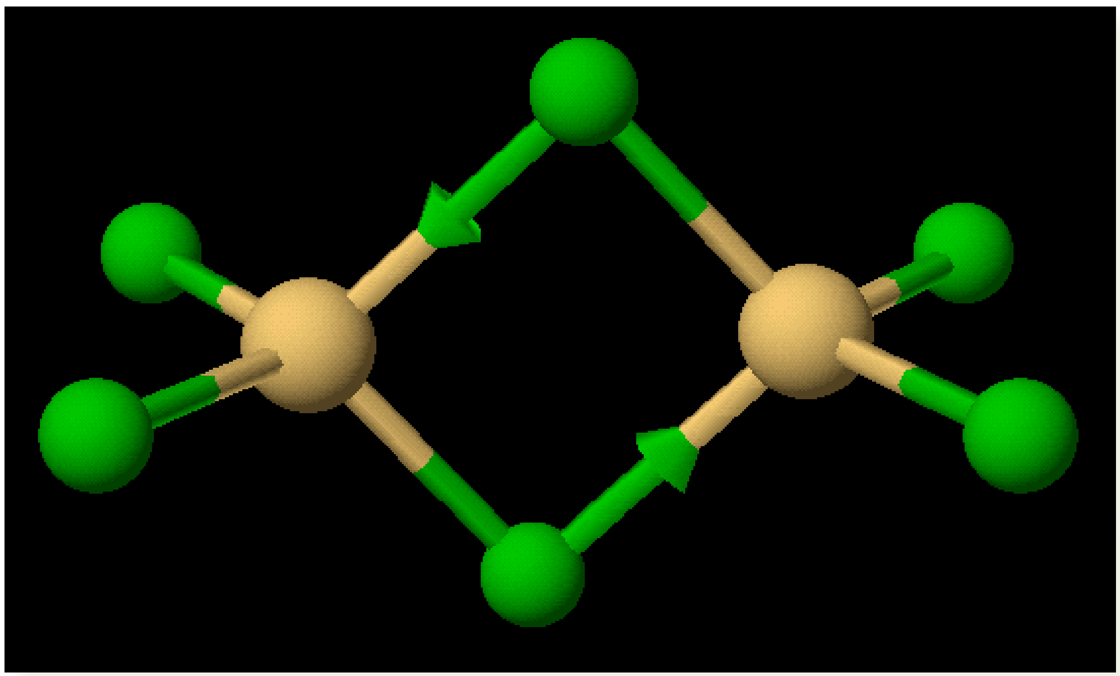
Limiting Reactants

Step 3 is to subtract from the amount of excess reactant present initially the amount that is used by all of the limiting reactant. The difference is the amount of excess reactant left.

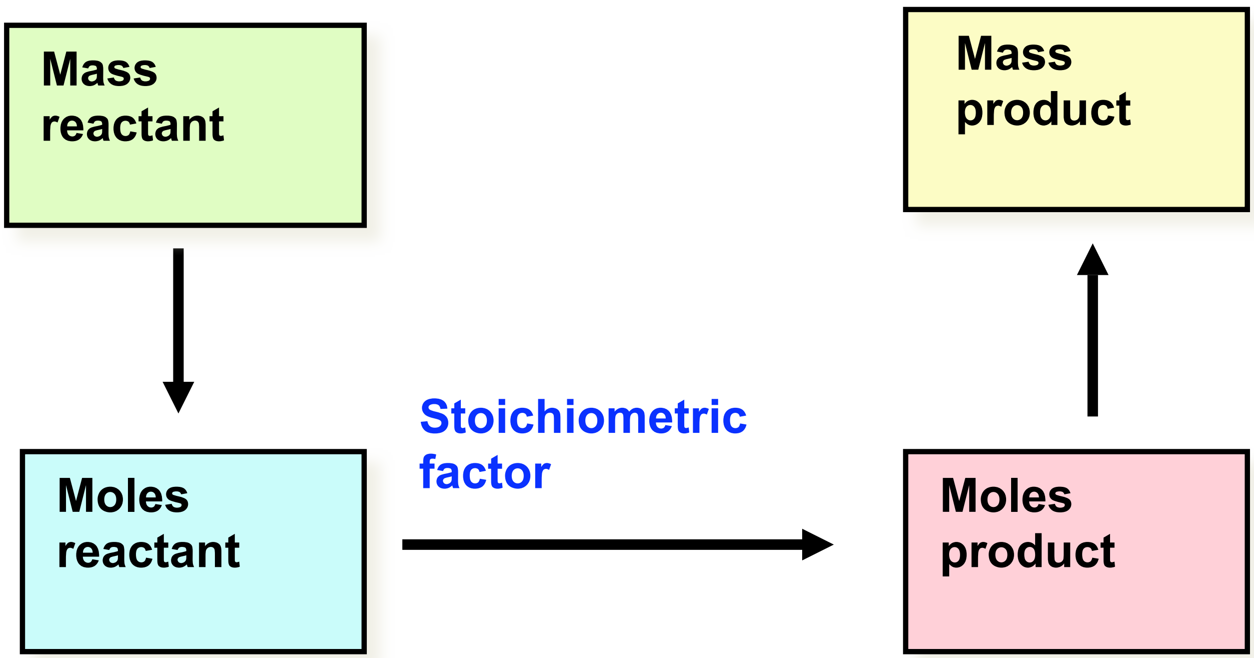
$$\begin{array}{r} 24.0 \text{ g } H_2O \text{ initially} \\ -\ 18.8 \text{ g } H_2O \text{ reacted} \\ \hline \end{array}$$

Reaction to be Studied

$$2\ Al(s) \ + \ 3\ Cl_2(g)\ f\ Al_2Cl_6(s)$$



PROBLEM: Mix 5.40 g of Al with 8.10 g of Cl_2. What mass of Al_2Cl_6 can form?



Step 1 of LR problem: compare actual mole ratio of reactants to theoretical mole ratio.

$$2\ Al(s)\ +\ 3\ Cl_2(g)\ f\ Al_2Cl_6(s)$$

Reactants must be in the mole

$$\frac{mol\ Cl_2}{mol\ Al} = \frac{3}{2}$$

Step 2 of LR problem: Calculate moles of each reactant

We have 5.40 g of Al and 8.10 g of Cl_2

$$5.40 \text{ g Al} \left(\frac{1 \text{ mol}}{27.0 \text{ g}}\right) = 0.200 \text{ mol Al}$$

$$8.10 \text{ g } Cl_2 \left(\frac{1 \text{ mol}}{70.9 \text{ g}}\right) = 0.114 \text{ mol } Cl_2$$

Find mole ratio of reactants

$2\ Al(s) + 3\ Cl_2(g)$ f $Al_2Cl_6(s)$

1. 0.200 mol Al X $\frac{3\ \text{mol}\ Cl_2}{2\ \text{mol Al}}$ = 0.300 mol Cl_2
2. needed to react with 0.200 mol Al, but we only have 0.114 mole of Cl_2, so

Limiting reactant is Cl_2

Mix 5.40 g of Al with 8.10 g of Cl_2. What mass of Al_2Cl_6 can form?

$2\ Al(s) + 3\ Cl_2(g)$ f $Al_2Cl_6(s)$

Limiting reactant = Cl_2

Base all calcs. on Cl_2

mass Cl_2

moles Cl_2

$\xrightarrow{\frac{1\ mol\ Al_2Cl_6}{3\ mol\ Cl_2}}$

moles Al_2Cl_6

mass Al_2Cl_6

CALCULATIONS: calculate mass of Al_2Cl_6 expected.

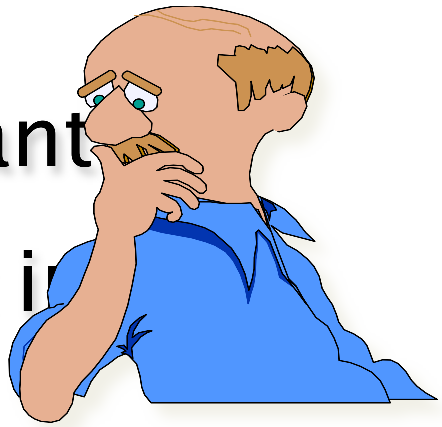
Step 1: Calculate moles of Al_2Cl_6 expected based on LR.

$$0.114 \text{ mol } Cl_2 \left(\frac{1 \text{ mol } Al_2Cl_6}{3 \text{ mol } Cl_2} \right) = 0.0380 \text{ mol } Al_2Cl_6$$

Step 2: Calculate mass of Al_2Cl_6 expected based on LR.

$$0.0380 \text{ mol } Al_2Cl_6 \left(\frac{266.4 \text{ g } Al_2Cl_6}{\text{mol}} \right) = 10.1 \text{ g } Al_2Cl_6$$

How much of which reactant will remain when reaction is complete?



1. Cl_2 was the limiting reactant. Therefore, Al was present in excess. But how much?
2. First find how much Al was required.

Calculating Excess Al

$$2\,Al + 3\,Cl_2 \longrightarrow \text{products}$$

0.200 mol

0.114 mol = LR

$$0.114 \text{ mol } Cl_2 \left(\frac{2 \text{ mol Al}}{3 \text{ mol } Cl_2}\right) = 0.0760 \text{ mol Al req'd}$$

Excess Al = Al available - Al required

= 0.200 mol - 0.0760 mol

= 0.124 mol Al in excess

Energy

Goal 6

Given energy in one of the following units, calculate the other three: joules, kilojoules, calories, and kilocalories.

Energy

Energy
The ability to do work or transfer heat.

Energy Unit: Joule

Joule (J)
The amount of energy exerted when a force of one newton (force required to cause a mass of 1 kg to accelerate at a rate of
1 m/s^2) is applied over a displacement of one meter:

$$1 \text{ joule} = 1 \text{ newton} \times 1 \text{ meter}$$

$$1 \text{ J} = 1 \text{ kg} \bullet \text{m}^2/\text{sec}^2$$

Energy

Energy Unit: Calorie

calorie (historical definition):
Amount of energy required to raise the temperature of 1 g of water by 1°C.

calorie (modern definition)
$1 \text{ cal} \equiv 4.184 \text{ J}$

A food Calorie (Cal) is a thermochemical kilocalorie.
In scientific writing it is capitalized;
in everyday writing, interpret the context.

$1 \text{ kcal} = 4.184 \text{ kJ}$

Energy

Example:

How many joules are equivalent to 2.4×10^2 food Calories?

Solution:

A food Calorie is a thermochemical kilocalorie.

GIVEN: 2.4×10^2 kcal WANTED: J

PER: 1000 cal/kcal 4.184 J/cal

PATH: kcal $\longrightarrow$ cal $\longrightarrow$ J

$$\frac{1000 \text{ cal}}{\text{kcal}} \quad \frac{4.184 \text{ J}}{\text{cal}}$$

Thermochemical Equations

Goal 7

Given a chemical equation, or information from which it may be written, and the heat (enthalpy) of reaction, write the thermochemical equation either (a) with ΔH to the right of the conventional equation or (b) as a reactant or product.

Thermochemical Equations

Enthalpy
Heat content of a system.
Symbol: H

ΔH is the enthalpy of reaction, or,
as chemists commonly call it,
the heat of reaction.

Δ means "change in":
It is determined by subtracting the initial quantity from the final quantity:
ΔA = final value of A – initial value of A

Thermochemical Equations

Enthalpy of Reaction, ΔH

If the final heat content of a system is less than its
initial heat content, ΔH is negative,
and the change is exothermic.

If the final heat content of a system is greater than its
initial heat content, ΔH is positive,
and the change is endothermic.

Thermochemical Equations

Thermochemical Equation
A conventional equation plus its associated ΔH term.

Thermochemical equations are typically written in one of two different ways, depending on the context for which it is needed.

Example:
When one mole of acetylene gas reacts with two moles of hydrogen gas, one mole of ethane gas is formed and 311 kJ of energy is released. Write the thermochemical equation in two forms.

Solution:

$$C_2H_2(g) + 2\ H_2(g) \rightarrow C_2H_6(g) + 311\ kJ$$

$$C_2H_2(g) + 2\ H_2(g) \rightarrow C_2H_6(g) \qquad \Delta H = -311\ kJ$$

Thermochem Stoichiometry

Goal 8

Given a thermochemical equation, or information from which it may be written, calculate the amount of energy released or added for a given amount of reactant or product; alternately, calculate the mass of reactant required to produce a given amount of energy.

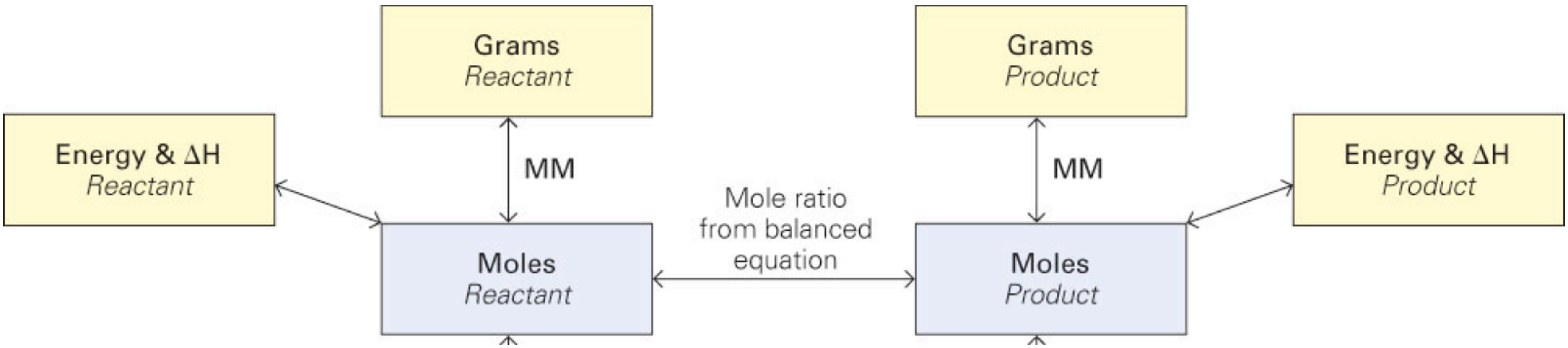
Thermochem Stoichiometry

The energy term in a thermochemical equation
is directly proportional to the number of moles of
all
species in the reaction.

$$C_2H_2(g) + 2\ H_2(g) \rightarrow C_2H_6(g) + 311\ kJ$$

$$311\ kJ \propto 1\ mol\ C_2H_2$$

Thermochem Stoichiometry



Thermochem Stoichiometry

Example:
When propane, $C_3H_8(g)$, is burned to form gaseous carbon dioxide and liquid water, 2.22×10^3 kJ of heat is released for every mole of propane burned. What quantity of heat is released when 1.00×10^2 g of propane is burned?

Solution:
First, write and balance the thermochemical equation to determine the stoichiometric relationships.

Thermochem Stoichiometry

GIVEN: 1.00×10^2 g C_3H_8 WANTED: quantity of heat

(assume kJ)

PER: 1 mol C_3H_8/44.09 g C_3H_8

PATH: g C_3H_8 $\longrightarrow$ mol C_3H_8 $\longrightarrow$

2.22×10^3 kJ/1 mol C_3H_8

$$\cancel{\text{g } C_3H_8} \times \frac{1 \text{ mol } \cancel{C_3H_8}}{44.09 \cancel{\text{ g } C_3H_8}} \times \frac{2.22 \times 10^3 \text{ kJ}}{1 \cancel{\text{ mol } C_3H_8}} = \text{kJ}$$