# Friday, September 19, 2025 AP Chemistry

Read 3.1–3.5 due Tu ~~9/22~~ 9/23
Read 3.6–3.7 due Th 9/~~24~~25
Paint Lab Th 9/25 Group Y M 9/29 Group X
Problem Set F 10/3
Quiz Tu 10/7

# Ch. 3 Reactions and Stoichiometry

## 3.1 Chemical Equations

(Reactants) $\xrightarrow[\text{into}]{\text{transform}}$ (Products)

The number of atoms is not changed in rxns.
We use coefficients in chem. eqns. to show this.
When balancing eqns Least Common Multiple (LCM) and guess-and-check work fine.

## 3.2 Simple Patterns of Chemical Reactions

Phases of matter help us to picture rxns as real events.

Solid (s) Liquid (l) Gas (g) Aqueous (aq) — dissolved in water

### Three Types of Rxns

Synthesis (combination) $A + B \rightarrow C$ (compound)

Decomposition $C \rightarrow A + B$

Combustion $C_xH_yO_z + O_2 \rightarrow CO_2 + H_2O$

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# SYNTHESIS RXN EXAMPLES $A + B \rightarrow C$

**NON-METALS COMBINE**

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$$

$$3H_{2(g)} + N_{2(g)} \rightarrow 2NH_{3(g)}$$

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**METALS + NON-METALS MAKE SALTS / (ANY IONIC COMPOUNDS)**

$$2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)} \quad Na^{+}\ Cl^{-}$$

$$Mg_{(s)} + Br_{2(g)} \rightarrow MgBr_{2(s)} \quad Mg^{2+}\ Br^{-}$$

NEUTRAL ELEMENTS → IONIC COMPOUNDS

$$4Fe_{(s)} + 3O_{2(g)} \rightarrow 2Fe_2O_{3(s)}$$

$$Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$$

(DEMO)

ALSO TRY $(Li + S)$ $(Ca + P_4)$ $(Al + Br_2)$

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**METAL OXIDE + WATER ⇓ BASE ($OH^-$)**

$$MgO_{(s)} + H_2O_{(l)} \rightarrow Mg(OH)_{2(aq)}$$

(DEMO)

MAKES PHENOLPHTHALEIN PINK

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**NON-METAL OXIDE + WATER ⇓ ACID ($H^+$)**

ACID RAIN $$N_2O_{5(g)} + H_2O_{(l)} \rightarrow 2HNO_{3(aq)}$$

CARBONATED DRINKS $$CO_{2(g)} + H_2O_{(l)} \rightarrow H_2CO_{3(aq)}$$

F 2025-09-19 AP CHEM

# Decomposition Reactions $C \rightarrow A + B$

Compounds decompose when exposed to heat, ultraviolet light, or catalysts.

$$2H_2O_{2(aq)} \rightarrow 2H_2O_{(l)} + O_{2(g)}$$

(DEMO)

Heating carbonates

Oxide + carbon dioxide

$$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$$

Limestone — Lime

$$2NaHCO_{3(s)} \rightarrow Na_2O_{(s)} + H_2O_{(l)} + CO_{2(g)}$$

Baking soda

$$(NH_4)_2CO_{3(s)} \rightarrow 2NH_{3(g)} + H_2O_{(g)} + CO_{2(g)}$$

(Demo)

Stopped here Group Y 2025-09-19 F without doing $(NH_4)_2CO_3$ demo

# Combustion Reactions

$$\boxed{\text{Hydrocarbon}} + O_2 \rightarrow CO_2 + H_2O$$

- $CH_4$ Methane
- $C_2H_6$ Ethane
- $C_2H_6O$ Ethanol

Gasoline
Diesel or wax!

# 3.3 Formula Weights

Add up atomic masses in a formula:

$H_2O$ $(2 \times H) + (1 \times O) = 2 \cdot 1.008 + 1 \cdot 15.999$

18.015 amu (or u)

for ONE molecule

How about $Ca(NO_3)_2$? You try!

# 3.4 Avogadro's Number and the Mole

Avogadro proposed that equal volumes of ~~gas~~ any gas at the same temp. & pressure have THE SAME number of molecules.

This was (eventually) used to allow chemists to get a count of the number of molecules.

At 0°C and 1 atmosphere of pressure the number of moleules in 22.41 L is $6.022 \times 10^{23}$.

In order to making counting atoms easy the amu (or u) was related to grams as:

$$6.022 \times 10^{23}\ u = 1\ g$$

So we can count atoms by mass:

$$\cancel{26.982\,g}\ Al \cdot \frac{6.022 \times 10^{23}\ \cancel{u}}{1\ \cancel{g}} \cdot \frac{1 \text{ atom of Al}}{26.982\ \cancel{u}} = 6.022 \times 10^{23} \text{ Al atoms}$$

(MASS)

164.1 amu

TO MAKE COUNTING MORE CONVENIENT WE USE A COUNTING UNIT CALLED <u>THE MOLE</u>.

THE MOLE IS LIKE THE DOZEN:

1 DOZEN = 12 ITEMS

1 MOLE = $6.022 \times 10^{23}$ ITEMS
(ABBR. mol)

So

26.982g Al = $6.022 \times 10^{23}$ ATOMS OF Al = 1 mol Al

26.982g Al = 1 mol Al

MOLAR MASS IS THE MASS IN GRAMS OF 1 mol OF A CHEMICAL
FORMULA WEIGHT IS THE MASS IN u OF <u>1</u> MOLECULE " " "
NUMERICALLY, THESE ARE THE SAME NUMBER!

## WHAT DO MOLES DO?

① COUNT THINGS BY WEIGHING

$$37\,g\ H_2O \cdot \frac{1\ mol\ H_2O}{18.015\,g} = 2.05\ mol\ H_2O$$

② FIND THE MASS FOR A NUMBER OF ITEMS

$$0.125\ mol\ H_2O \cdot \frac{18.015\,g}{1\ mol\ H_2O} = 2.25\,g\ H_2O$$

③ MOLES CAN BE USED IN CHEM. FORMULAS: <u>2 mol $NH_3$</u>
OR IF I FIND A RATIO BY NUMBER OF ATOMS IN MOLES, IT TELLS ME THE FORMULA.

2 mol N
6 mol H

(4) MOLES CAN RELATE AMOUNTS OF CHEMICALS IN CHEMICAL EQUATIONS

$$2H_2 + O_2 \rightarrow 2H_2O$$

2 mol $H_2$    1 mol $O_2$    2 mol $H_2O$

## 3.5 EMPIRICAL FORMULAS

EMPIRICAL MEANS DERIVED FROM EXPERIMENTAL DATA, LIKE YOUR DENSITY MEASUREMENTS.

IT MAY BE A CALCULATED RESULT BUT IT MUST BE BASED ON LAB MEAS. TO BE EMPIRICAL.

THEORETICAL IS THE OPPOSITE OF EMPIRICAL; IT MEANS DERIVED FROM THEORY, OFTEN A MATHEMATICAL MODEL OF SOMETHING. A THEORETICAL CALC. OF MASS USING A REFERENCE DENSITY VALUE AND A MEASURED VOL. IS MORE LIKE A THEOR. CALC.

EMPIRICAL CHEMICAL FORMULAS USE THE MASS OF COMPONENT ELEMENTS IN A COMPOUND TO CALC. AN INTEGER RATIO OF ELEMENTS USING MOLES.

FOR EX., A COMPOUND OF Fe AND O HAS FORMULA $Fe_xO_y$

WE HAVE 159.69g OF THIS COMPOUND AND ANALYSIS SHOWS:

$$111.69\text{g Fe} \cdot \frac{1\text{ mol Fe}}{55.845\text{g}} = 2\text{ mol Fe}$$

$$48.0\text{g O} \cdot \frac{1\text{ mol O}}{15.999\text{g}} = 3\text{ mol O}$$

$Fe_2O_3$

FIND A RATIO BY MOLES

$$\frac{Fe}{O} \quad \frac{2}{3} = 0.6667$$

$$\frac{O}{Fe} \quad \frac{3}{2} = 1.5$$

Tu 2025-09-23 AP CHEM

# PERCENT BY MASS AND CALCULATING AN EMP. FORMULA

$Fe_xO_y$

69.942% Fe } ASSUME 100g
30.058% O }

KEEP 4-5 s.f. EVEN IF SMALL: 0.0012524

$$69.942g\ Fe \cdot \frac{1\ mol\ Fe}{55.845g} = 1.2524$$

$$30.058g\ O\ \frac{1\ mol\ O}{15.999g} = 1.8787$$

CALC. A RATIO BY MOLES BY DIVIDING ALL ATOMS BY THE SMALLEST NO. OF MOLES. THE ATOM WITH THE SMALLEST NUMBER GOES TO <u>ONE</u>.

$$\frac{O}{Fe} \quad \frac{1.8787}{1.2524} = 1.5001 \sim 1.5 \text{ so } ~~Fe_1O_{1.5}~~$$

$$\frac{1.50}{1\ Fe} \times \frac{2}{2} = \frac{3O}{2Fe} \text{ so } Fe_2O_3$$

| $\frac{1}{2}$ | $\frac{1}{3}$ | $\frac{1}{4}$ | $\frac{1}{5}$ | $\frac{1}{6}$ | $\frac{1}{7}$ | $\frac{1}{8}$ |
|---|---|---|---|---|---|---|
| 0.5 | 0.333 | 0.25 | 0.2 | 0.167 | 0.143 | 0.125 |

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PAUSE NOTES, DO $Ag_2O$ DEMO

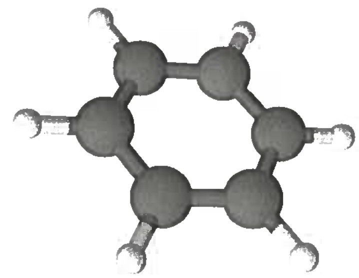
Tu 2025-09-23
AP Chem X and Y

# Student Worksheet for the Demonstration: The Empirical Formula of Silver Oxide

There are millions of chemical compounds described in the scientifc literature. In order to classify and organize all of these different substances it's necessary to specify the their elemental composition. References of chemical compounds list their properties along with a chemical formula showing how many atoms of which elements make up individual molecules of that compound. How are these formulas determined?

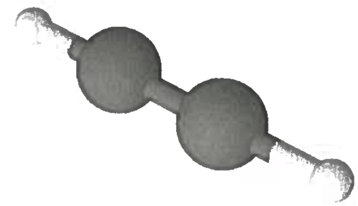
While there are a large variety of modern, highly technological methods for determining the formulas of molecules (such as NMR) there is a tried and true method for doing it, too. The idea is simple enough: break a compound down into its elements, count the atoms of each element, and calculate a ratio in terms of whole numbers. For example, in order to determine the formula of the flammable gas known as acetylene one could burn it and collect the carbon dioxide and water vapor that results. Since acetylene is composed of only carbon and hydrogen it's possible to count the atoms of carbon in the carbon dioxide and the atoms of hydrogen in the water to find out about the acetylene that was burned. When this is done the ratio of atoms is found to be 1 C for every 1 H. This is commonly written as a chemical formula: CH.

The same procedure can be used for other compounds, such as benzene. Benzene is physically and chemically distinct from acetylene as it is a poisonous liquid which evaporates to make a flammable vapor. And yet when its formula is determined as described above it is found to also be CH. So it appears that two compounds have the same formula!

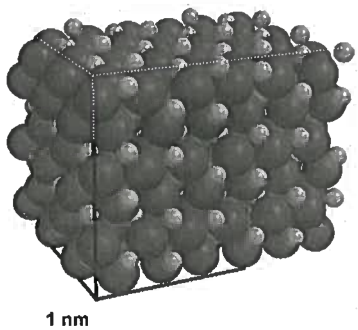


Benzene ($C_6H_6$)

It turns out these two compounds have the same lowest-whole-number-ratio of atoms but they have different molecular structures. Both benzene and acetylene are shown in images at right. There are many hydrocarbon compounds that have the same ratio of carbon and hydrogen atoms but which have different chemical and physical properties. This is because these compounds have different numbers of atoms in their molecular structures. Acetylene and benzene have the same **empirical formula**, which is a formula with the lowest whole number ratio of atoms in the compound. These compounds have different **molecular formulas**, which give the true number of atoms in each molecule of the substance.



Acetylene ($C_2H_2$)




iron(III) oxide ($Fe_2O_3$) crystal

Some combinations of elements have one or more possible ratios but all samples with the same empirical formula are the same substance. For example, iron and oxygen form two main compounds: iron(II) oxide (FeO) and iron(III) oxide ($Fe_2O_3$). There are no compounds with the formula $Fe_2O_2$ or $Fe_4O_6$. This is because these compounds are ionic in nature, that is, they are composed of a specific ratio of two oppositely charged ions. As a result, they have two types of smallest particle, not just one as molecular compounds do. In the solid state such compounds form a repeating pattern in three dimensions called a crystal lattice. There is an image of the lattice of iron(III) oxide at left. Throughout the structure the ratio of positive ions to negative ions is the same. Ionic compounds are defined only in terms of the lowest whole number ratio of the ions of which they are composed. In other words, ionic compounds have only an empirical formula and do not have a molecular formula.

In this demonstration your teacher will show you a method by which the empirical formula of silver oxide may be found. Silver oxide will decompose on heating to form silver metal and oxygen gas. By carefully weighing a sample, placing it in a crucible, heating it over a bunsen burner, and weighing the resulting metal it will be possible to determine the ratio of silver to oxygen atoms in the original compound.

**Questions**

1. Use the data in the table at right to calculate the amount of iron oxide produced by rusting the steel wool. Also, calculate the amount of oxygen added to the steel wool in the formation of the iron oxide. (The data were generated by rapidly rusting steel wool, which is a form of iron. Rusting is a result of combination with oxygen.)

| Empirical Formula of Iron Oxide (or Rust) | |
|---|---|
| crucible | 12.988 g |
| steel wool | 1.175 g |
| rusted steel wool + crucible | 14.668 g |

14.668
− 12.988
1.680 g RUST

1.680 g RUST
− 1.175 g IRON
0.505 g OXYGEN

2. Calculate the moles of iron from the mass of the steel wool. Also, calculate the moles of oxygen atoms from the mass of oxygen you calculated in the previous step.

$1.175g\ Fe \frac{1\ mol}{55.845g} = 0.02\ \ 0.021\ \ 2.104 \times 10^{-2}\ mol\ Fe$

$0.505g\ O \frac{1\ mol}{15.999g} = 3.156 \times 10^{-2}\ mol\ O$

3. Calculate a ratio of moles of oxygen to moles of iron. Convert this ratio to a ratio of whole numbers.

$\frac{O}{Fe}\ \frac{3.156 \times 10^{-2}}{2.104 \times 10^{-2}} = \frac{1.5}{1} = \frac{3}{2}\ \ Fe_2O_3$

4. Write the chemical formula of iron oxide based on your calculations. Also, write a balanced chemical equation for the reaction of iron with oxygen to form this compound.

$Fe_2O_3$ $\quad Fe + O_2 \rightarrow Fe_2O_3$

5. As the demonstration proceeds, record the data in the table at right. Use the data to calculate the mass of silver oxide, the mass of silver at the end of the experiment, and the mass of oxygen that was driven off in the decompostion reaction.

| Empirical Formula of Silver Oxide | |
|---|---|
| crucible | 12.761 g |
| crucible + silver oxide | 13.288g |
| crucible + silver metal | 13.252g |

SILVER OXIDE: 0.527g
SILVER ALONE: 0.491g
OXYGEN LOST: 0.036g

6. Calculate the moles of silver and the moles of oxygen. Use these numbers to calculate a molar ratio for mol Ag/mol O in whole numbers.

$0.491g \frac{1\ mol}{107.868g} = 4.55 \times 10^{-3}\ mol\ Ag$

$0.036g \frac{1\ mol}{15.999g} = 2.25 \times 10^{-3}\ mol\ O$

$\frac{Ag}{O}\ \frac{4.55 \times 10^{-3}}{2.25 \times 10^{-3}} = 2.02 / 2$

7. Write the chemical formula of silver oxide based on your calculations. Also, write a balanced chemical equation for the decomposition of silver oxide to form oxygen and silver.

$2Ag_2O \rightarrow 4Ag + O_2$

2:1 RATIO    4:1 RATIO

8. In your own words, define the terms empirical formula and molecular formula.

9. In your own words, explain why ionic compounds have empirical formulas but not molecular formulas.

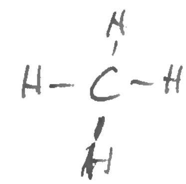
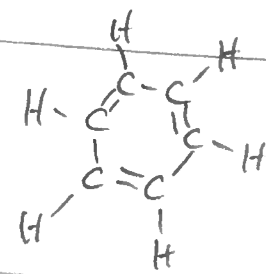
# Thursday, September 25, 2025 AP Chemistry

Other things to know about empirical formulas.

Ionic compounds' formulas are all empirical formulas as they show the smallest whole-number ratio which gives a neutral ion combination.

For molecular compounds, the emp. formula and molecular formula may be the same or different.

For ex.

| | Emp Formula | Molecular Formula | Structure |
|---|---|---|---|
| Methane | $CH_4$ | $CH_4$ |  |
| Acetylene | CH | $C_2H_2$ | H−C≡C−H |
| Benzene | CH | $C_6H_6$ |  |

We use molar mass to identify the molecular formula once we know the empirical formula.

Acetylene  CH 13 g/mol  26 g/mol  $\frac{26}{13} = 2$ so $C_2H_2$

Benzene  CH 13 g/mol  78 g/mol  $\frac{78}{13} = 6$ so $C_6H_6$

# SOLVING STOICHIOMETRY PROBLEMS

STOICHIOMETRY IS SOLVING PROPORTIONS USING MOLES WITH RATIOS BASED ON BALANCED CHEM. EQNS. WE'LL USE DIMENSIONAL ANALYSIS TO ORGANIZE OUR CALC.

RATIOS WE USE COME FROM BAL. CHEM. EQNS:

$$2H_2 + O_2 \rightarrow 2H_2O$$

MOLAR RATIOS

- MOLAR NOT MORAL
- NO STRESS B/C NO DENTISTRY

$$\frac{2\text{ mol }H_2}{1\text{ mol }O_2} \quad \frac{1\text{ mol }O_2}{2\text{ mol }H_2} \quad \frac{2H_2}{1O_2}$$

$$\frac{2H_2}{2H_2O} \quad \frac{2H_2O}{1O_2}$$

THESE ARE DIRECT PROPORTIONS. WE USE THEM AS CONVERSION FACTORS.

THE MOST FUNDAMENTAL STOICHIOMETRY CALCULATION CONVERTS <u>MOLES</u> OF ONE SUBSTANCE INTO MOLES OF ANOTHER.

MOLAR RATIOS CONVERT mol A to mol B

THEY CAN'T BE USED DIRECTLY FOR MASSES:

$$2H_2 + O_2 \rightarrow 2H_2O$$

$$2(2.02\text{ g/mol}) \quad 1(32.0\text{ g/mol}) \quad 2(18.0\text{ g/mol})$$

$$(4\text{g})\,H_2 + (32\text{g})\,O_2 \rightarrow (36\text{g})\,H_2O$$

→ MASS RATIOS COULD BE USED BUT THEY'RE NOT AS CONVENIENT.

(ASIDE: THE LAW OF CONSERVATION OF MASS IS THE UNDERLYING PRINCIPLE).

Th 2025-09-25 AP CHEM

## STOICH. CALC.

$$2H_2 + O_2 \rightarrow 2H_2O$$

### MOLES TO MOLES

0.75 mol $H_2$ — HOW MANY mol $O_2$ ARE NEEDED?

$$0.75 \text{ mol } H_2 \cdot \frac{1\ O_2}{2\ H_2} = \boxed{0.375 \text{ mol } O_2}$$

### GRAMS TO GRAMS

THIS REQUIRES A g → mol CONVERSION
WHICH MEANS YOU HAVE TO SEPARATELY CALC.
THE MOLAR MASS

YOU HAVE 17g $H_2$, HOW MANY g $H_2O$ CAN BE MADE?

GIVEN

$$17\text{ g } H_2 \cdot \frac{1 \text{ mol } H_2}{2.0 \text{ g } H_2} \cdot \frac{2\ H_2O}{2\ H_2} \cdot \frac{18.0 \text{ g } H_2O}{1 \text{ mol } H_2O} = \boxed{153 \text{ g } H_2O}$$

- 17 ↑ NO.
- g ↑ UNIT
- $H_2$ ↑ FORMULA

| g A → mol A USING MOLAR MASS | mol A → mol B | mol B → g B USING MOLAR MASS |
|---|---|---|
| | STOICH. STEP | |

YOU TRY: THERE ARE TWO WAYS TO CALC. MASS OF $O_2$ IN GRAMS FOR THIS SCENARIO. DO BOTH.

$$17\text{g } H_2 \cdot \frac{1 \text{ mol } H_2}{2.0\text{g } H_2} \frac{1\ O_2}{2\ H_2} \cdot \frac{32.0\text{g}}{1 \text{ mol } O_2} = 136\text{g } O_2$$

USE THE LAW OF CONSERVATION OF MATTER!

$$17 + x = 153 \quad \text{so} \quad x = 136\text{g } O_2$$

# STOICHIOMETRY WITH LIMITING REACTANT

[CHEESE SANDWICHES]

CHEMICALS ARE FREQUENTLY REACTED SO THAT ONE IS USED UP BEFORE THE OTHER.

THE REACTANT WHICH IS USED UP IS CALLED THE LIMITING REACTANT.

THE REACTANT WHICH LEAVES SOME LEFT OVER IS CALLED THE EXCESS REACTANT.

IN ORDER TO SOLVE STOICH. PROBLEMS IN WHICH AMOUNTS ARE GIVEN FOR BOTH REACTANTS YOU NEED TO IDENTIFY WHICH REACTANT IS LIMITING AND WHICH IS IN EXCESS. TO DO THIS, AMOUNTS NEED TO BE EXPRESSED IN MOLES.

SCENARIOS

$$2H_2 + O_2 \rightarrow 2H_2O$$

| | $H_2$ | $O_2$ |
|---|---|---|
| EX. 1 | 5 mol | 3 mol |
| EX. 2 | 7 mol | 2 mol |
| EX. 3 | 0.48 mol | 0.34 mol |

## THREE WAY TO ID THE LR

### METHOD 1 FOR EX. 1

COMPARE GIVEN RATIO TO STOICH. RATIO

GIVEN — NEED (ACCORDING TO EQN)

$$\frac{H_2}{O_2} \quad \frac{5}{3} \quad \frac{2}{1}$$

$$\left(\frac{1.67}{1}\right) < \frac{2}{1}$$

THE AMOUNT OF $H_2$ PER $O_2$ IS LESS THAN THE STOICH. RATIO SO $H_2$ IS THE LR.

WHAT IF... GIVEN — NEED

$$\frac{O_2}{H_2} \quad \frac{3}{5} > \frac{1}{2}$$

$$0.6 \quad 0.5$$

THE AMOUNT OF $O_2$ PER $H_2$ IS GREATER THAN THE STOICH. RATIO SO $O_2$ IS IN EXCESS ∴ **$H_2$ IS LR**

## Method 2 for ex. 2

For $H_2$

What if $H_2$ is LR?

If so, how much $O_2$ is used?

$$7\text{ mol } H_2 \cdot \frac{1\ O_2}{2\ H_2} = 3.5\text{ mol } O_2$$

We only have 2 mol $O_2$ so

**$O_2$ is LR.**

For $O_2$

What if $O_2$ is LR?

If so, how much $H_2$ is used?

$$2\text{ mol } O_2 \frac{2\ H_2}{1\ O_2} = 4\text{ mol } H_2 \text{ is used up}$$

We have 7 mol $H_2$ so

**$O_2$ is LR.**

## Method 3 for ex. 3

Find out which reactant makes less of a product — that one is the LR.

$$0.48\text{ mol } H_2 \frac{2\ H_2O}{2\ H_2} = 0.48\text{ mol } H_2O \text{ if } H_2 \text{ is LR}$$

$$0.34\text{ mol } O_2 \frac{2\ H_2O}{1\ O_2} = 0.68\text{ mol } H_2O \text{ if } O_2 \text{ is LR}$$

So $H_2$ is LR b/c $0.48 < 0.68$

This method has the advantage of telling us how much product is made.

See Demo: Stoichiometry and the Limiting Reactant

(Choose your favorite method and practice it in the PSet)

Stopped here Th 2025-09-25 X

# GENERAL OUTLINE FOR STOICH. PROBLEMS

1. CONVERT MASS TO MOLES $(g \rightarrow mol)$
   USING MOLAR MASSES
2. IDENTIFY THE LR
3. USE MOLAR RATIOS TO CALC.
   a. mol PRODUCT(S)
   b. mol OF EXCESS REACTANT WHICH IS USED UP (AND LEFT OVER BY SUBTRACTION)
4. CONVERT mol → g

## FULLY WORKED EXAMPLE

GIVEN

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

| | $3H_2$ | $N_2$ | $2NH_3$ |
|---|---|---|---|
| START | 5.7g | 35.1g | Øg |
| CHANGE | −5.7g | −26.4g | +31.8g |
| END | Øg (LR) | 8.7g | 31.8g |

1. g → mol

$$5.7g\ H_2 \cdot \frac{1\ mol\ H_2}{2.016g} = 2.8\ mol\ H_2$$

$$35.1g\ N_2 \cdot \frac{1\ mol\ N_2}{28.014\ g} = 1.25\ mol\ N_2$$

2. FIND LR

$$2.8\ mol\ H_2 \cdot \frac{2NH_3}{3H_2} = 1.87\ mol\ NH_3$$ (IF $H_2$ IS LR)

$$1.25\ mol\ N_2 \cdot \frac{2NH_3}{1N_2} = 2.50\ mol\ NH_3$$ (IF $N_2$ IS LR)

$H_2$ IS LR SO CHANGE IN MASS OF $H_2$ IS −5.7g

PRODUCT

3. a. 1.87 mol $NH_3$ ✓

b. $$5.7g\ H_2 \cdot \frac{1\ mol\ H_2}{2.016g} \cdot \frac{1N_2}{3H_2} \cdot \frac{28.014g}{1\ mol\ N_2}$$

26.4g $N_2$ IS USED UP

35.1 − 26.4 = 8.7g $N_2$ LEFT OVER

4. $$1.87\ mol\ NH_3 \cdot \frac{17.031\ g}{1\ mol\ NH_3} = 31.8g\ NH_3$$

CONS. OF MASS: TOTAL MASS

5.7 + 35.1 = 40.8g    8.7 + 31.8 = 40.5g (CLOSE ENOUGH! ROUNDING)

THEORETICAL YIELD: THE AMOUNT OF A PRODUCT THAT COULD BE MADE IF ALL OF THE LIMITING REACTANT IS USED UP.

ACTUAL YIELD: THE AMOUNT MEASURED TO HAVE BEEN MADE IN A PARTICULAR EXPERIMENT.

(ACTUAL) < (THEOR.) USUALLY

- SOME MATERIAL MAY SIMPLY NOT BE ABLE TO BE RECOVERED
- SOME REACTIONS REACH EQUILIBRIUM SO THAT NOT ALL OF THE LR REACTS

PERCENT YIELD: $\frac{\text{ACTUAL YIELD}}{\text{THEOR. YIELD}} \times 100\%$

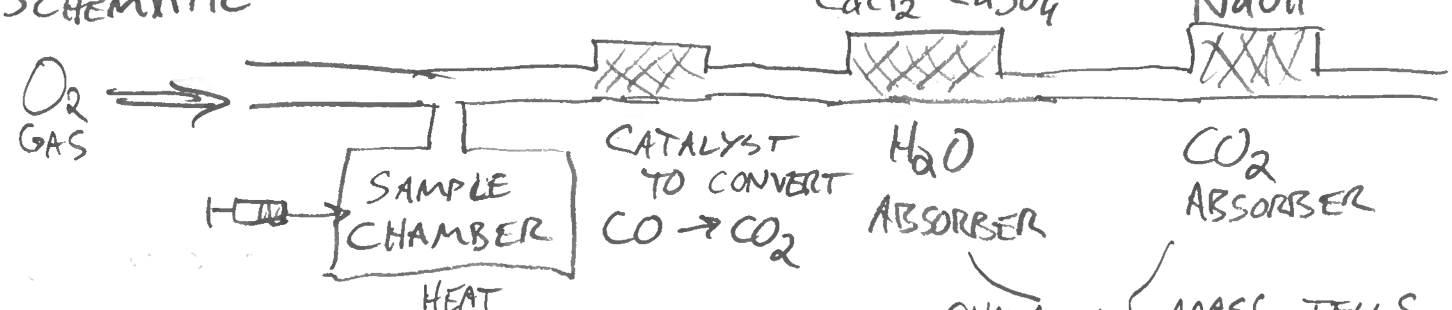
HOW MUCH DID WE ACTUALLY GET COMPARED TO THE AMOUNT WE EXPECTED?

## COMBUSTION ANALYSIS

A PROCEDURE FOR DETERMINING THE EMPIRICAL FORMULA OF A HYDROCARBON BY BURNING IT.

$C_xH_y$ OR $C_xH_yO_z$ OR $C_xH_yN_z$ OR $C_xH_yCl_z$

SCHEMATIC




CHANGE IN MASS TELLS YOU HOW MUCH $H_2O$ AND $CO_2$ ARE PRODUCED.

FOR COMB. ANALYSIS CALC.

YOU START WITH 1. MASS OF SAMPLE

2. MASS OF $H_2O$

3. MASS OF $CO_2$

FOR EXAMPLE, FIND THE EMP. FORMULA OF ISOPROPYL ALCOHOL

$C_xH_yO_z$

GIVEN DATA: 0.255g ISOPROPYL ALCOHOL

⇓

0.306g $H_2O$ AND 0.561g $CO_2$

GOAL: FIND LOWEST WHOLE-NUMBER RATIO OF ELEMENTS

mol H: $0.306g\ H_2O \cdot \frac{1\ mol\ H_2O}{18.015g} \cdot \frac{2\ H}{1\ H_2O} = 0.03397\ mol\ H$

(KEEP 4 s.f. TO AVOID ROUNDING ERRORS) ✱

mol C: $0.561g\ CO_2 \cdot \frac{1\ mol\ CO_2}{44.009g} \cdot \frac{1C}{1CO_2} = 0.01275\ mol\ C$

MASS O (g) BY SUBTRACTION

$0.03397\ mol\ H \cdot \frac{1.008g}{1\ mol\ H} = 0.03424g\ H$

$0.01275\ mol\ C \cdot \frac{12.011g}{1\ mol\ C} = 0.153g\ C$

$0.255 - 0.03424 - 0.153 = 0.06776g\ O$

($C_xH_yO_z$ − H − C = O)

mol O: $0.06776g\ O \cdot \frac{1\ mol\ O}{15.999g} = 0.004235\ mol\ O$ ($4.235 \times 10^{-3}$)

DIVIDE ALL mol BY A COMMON DENOMINATOR (WHICHEVER ELEMENT HAS FEWEST mol)

C: $\frac{0.01275}{0.004235} \sim 3$

H: $\frac{0.03397}{0.004235} \sim 8$

O: $1$

$C_3H_8O$

9/29 X – GAVE COMB. ANALYSIS PACKET

STOPPED HERE GROUP X M2025-09-29