
Stoichiometry
Topic#10
AMSAT Chem 1H

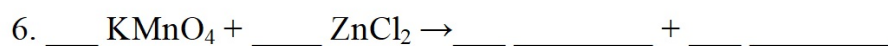
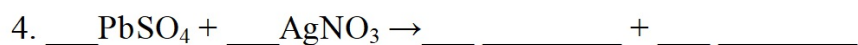
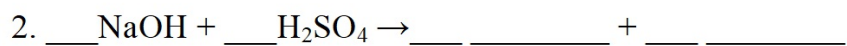
Teacher Edition

Stoichiometry

Reaction Review

Topic#10

Starter #1 - Reactions Review: Predict Products and balance.



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Moles - A Review

Molar mass (*MM*)

- the mass of 1 mole of a substance
 - Molar Mass (*MM*) of H₂O is 18.02 (2H + O = 2(1.01) + 16 = 18.02g/mol)
 - CuSO₄ • 5H₂O (copper (II) sulfate pentahydrate) = Cu + S + 4O + 10H + 5O = 63.55 + 32.07 + 4(16) + 10(1.01) + 5(16) = 249.72g/mol

Conversions (Go back to Topic#4 and fetch your conversion chart for moles!)

- mol to mass: multiply by *MM*
 - *label: g*
- mass to mol: divide by *MM*
 - *label: mol*
- mol to particles: multiply by 6.022x10²³parts/mol
 - *label: part* (ions, molecules, formula units, or atoms)
- particles to mol: divide by 6.022x10²³parts/mol
 - *label: mol*
- particles to grams: divide by 6.022x10²³parts then multiply by *MM*
 - *label: g*
- grams to particles: divide by *MM* then multiply by 6.022x10²³parts
 - *label: part* (ions, molecules, formula units, or atoms)

Stoichiometry

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Reaction Review

Re-Introduction to Moles

Groups of two on white board. Show work on whiteboard. Each member submits work on a piece of paper (3pts)

Starter #2 - Mole Practice Problems

1. Determine the number of moles of each element in 0.354 moles of $\text{Ba}(\text{NO}_3)_2$.

(Ans: 0.354 moles Ba, 0.708 moles N, and 2.124 moles of O)

2. What is the molar mass (*MM*) of $\text{Ba}(\text{NO}_3)_2$? (Ans: 261.35 g/mol)

3. How many grams are in 0.354 moles of $\text{Ba}(\text{NO}_3)_2$? (Ans: 92.5g)

4. How many moles are in 345.5 mg of $\text{Ba}(\text{NO}_3)_2$? (Ans:)

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Compositional Stoichiometry

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Composition Stoichiometry

- The use of a chemical formula, quantities of an element, and mole ratios to calculate the formula of a chemical species
- Keys to success
 - proper formula of a compounds
 - correct *MM* of compounds
 - using mole ratios between elements in a compound
 - conversion mass of an element to moles of an element

Empirical Formula vs Molecular Formula

Molecular Compounds (all nonmetals)

- Molecular formula: $C_6H_{12}O_6$ (each subscript can be divide by 6)
- Empirical formulas: CH_2O (empirical ratio (1:2:1))

Ionic Compounds (Starts with a metal or ammonium (NH_4^+))

- Ionic formula : titanium (IV) oxide $\rightarrow Ti_2O_4$ (simplify) $\rightarrow TiO_2$
 - always written as an empirical formula

Stoichiometry

Percent Composition

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Percent composition - the percent by mass of each element in a compound

$$\%X \text{ in } XY = \frac{\text{mass } X}{MM_{XY}} \times 100\%$$

Sample WS#1 - Percent Composition

1. Find the percentage composition of copper (I) sulfide, Cu_2S .

Ans: 79.9% Cu, 20.1% S

$$\% \text{ Cu} : \frac{2\text{Cu}}{\text{Cu}_2\text{S}} \times 100\% = \frac{2(63.55)}{2(63.55) + 32.07} \times 100\% = \frac{127.10}{159.17} \times 100\% = \boxed{79.9\%}$$

$$\% \text{ S} : \frac{\text{S}}{\text{Cu}_2\text{S}} \times 100\% = \frac{32.07}{2(63.55) + 32.07} \times 100\% = \frac{32.07}{159.17} \times 100\% = \boxed{20.1\%}$$

$$79.9 + 20.1 = 100\% \checkmark$$

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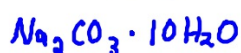
Percent Composition

Percent composition - the percent by mass of each element in a compound

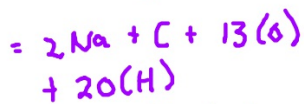
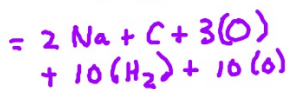
$$\%X \text{ in } XY = \frac{\text{mass } X}{MM_{XY}} \times 100\%$$

2. As some salts crystallize from a water solution, they bind water molecules in their crystal structure. Sodium carbonate forms such a *hydrate*, in which 10 water molecules are present for every formula unit of sodium carbonate. Find the mass percentage of water in sodium carbonate decahydrate, $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$, which has a *MM* of 286.14g/mol. Ans: 63.0% H_2O

Given



*
 $MM_{\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}}$



$$= 286.19 \text{ g/mol}$$

NTK

$$*MM_{\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}} = 286.19$$

$$mm_{10\text{H}_2\text{O}} = 10(18.02) = 180.2 \text{ g}$$

Unk

$$\% \text{H}_2\text{O} = \text{_____} \%$$

Solve:

$$\% \text{H}_2\text{O} = \frac{10\text{H}_2\text{O}}{\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}} \times 100\% = \frac{180.2}{286.19} \times 100\% = \boxed{63.0\%}$$

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Percent Composition

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Percent composition - the percent by mass of each element in a compound

$$\%X \text{ in } XY = \frac{\text{mass } X}{MM_{XY}} \times 100\%$$

3. (OYO) Find the percentage composition of the following. (Ans: (a) 74.5%Pb, 25.5%Cl (b) 52.5%Ba, 10.7%N, 36.7%O)

(a) PbCl_2

(b) $\text{Ba}(\text{NO}_3)_2$

4. (OYO) Find the mass percentage of water in $\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$.

(Ans: 43.9% H_2O)

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Percent Composition

5. Magnesium hydroxide is 54.87% oxygen by mass. How many grams of oxygen are in 175g of the compound? How many moles of oxygen is this? (Ans: 96.0g O, 6 moles)

Gvn
% Oxy by mass is 54.87%
175 g of compound

NTK
Mg(OH)₂

UNK
mass O: 96.0 g
moles O: 6 m

SOLVE:

$$\text{mass O: } 175 \text{ g} \times (0.5487) = \boxed{96.0 \text{ g O}}$$

$$\text{moles O: } \frac{96.0 \text{ g O}}{16 \text{ g}} = \boxed{6 \text{ moles O}}$$

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Empirical Formula

Empirical formula (EF) from Percent Composition.

- 1) Convert percents into grams (if given grams go to step 2)
- 2) Determine moles of each element by dividing the mass of the element by the atomic mass of each element.
- 3) Divide moles of each element by the smallest of them.
- 4) Convert to whole numbers by multiplying each by a factor.

Example: 1.333 is converted to 4 by multiplying by 3.

Stoichiometry Sample WS#2 - Empirical Formula Practice Problem:

1. Quantitative analysis shows that a compound contains 32.38% sodium, 22.65% sulfur, and 44.99% oxygen. Find the empirical formula of this compound. (Ans: Na_2SO_4)

<u>Gvn</u>	<u>SOLVE:</u>		
32.38% Na (32.38g)		Na_x	S_y
22.65% S (22.65g)	mass →	<u>32.38g</u>	<u>22.65g</u>
44.99% O (44.99g)	÷ atomic mass →	22.99	32.07
100 g Sample	moles →	<u>1.41</u>	<u>0.706</u>
<u>NTK</u>	÷ smallest value →	0.706	0.706
$\text{Na}_x\text{S}_y\text{O}_z$	mole ratio →	2	1
		<u>x</u>	<u>y</u>
			4
			<u>z</u>

$\text{Na}_x\text{S}_y\text{O}_z$

Na_2SO_4

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Empirical Formula

Empirical formula (EF) from Percent Composition.

- 1) Convert percents into grams (if given grams go to step 2)
- 2) Determine moles of each element by dividing the mass of the element by the atomic mass of each element.
- 3) Divide moles of each element by the smallest of them.
- 4) Convert to whole numbers by multiplying each by a factor.

Example: 1.333 is converted to 4 by multiplying by 3.

2. Analysis of a 10.150g sample of a compound known to contain only phosphorus and oxygen indicates a phosphorus content of 4.433g. What is the empirical formula of this compound?

(Ans: P₂O₅)

Given

$$m_T = 10.150 = m_P + m_O$$

$$m_P = 4.433$$

NTK

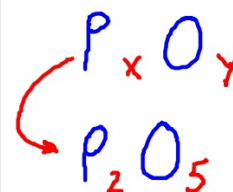
$$m_T - m_P = m_O$$

$$10.150 - 4.433 = m_O$$

$$m_O = 5.717$$

Solve:

	P _x	O _y
mass →	<u>4.433</u>	<u>5.717</u>
÷ atomic mass →	<u>30.97</u>	<u>16</u>
moles →	<u>0.143</u>	<u>0.357</u>
÷ smallest value →	<u>0.143</u>	<u>0.143</u>
mole ratio →	2(1)	2(2.5)
	X=2	Y=5



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Empirical Formula

3. (OYO) A compound is found to contain 63.52% iron and 36.48% sulfur. Find its empirical formula. (Ans: FeS)

<i>mass</i> →		
÷ <i>atomic mass</i> →		
<i>moles</i> →		
÷ <i>smallest value</i> →		
<i>mole ratio</i> →		

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Empirical Formula

4. (OYO) Find the empirical formula of a compound found to contain 26.56% potassium, 35.41% chromium, and the remainder oxygen. (Ans: $K_2Cr_2O_7$)

<i>mass</i> →			
÷ <i>atomic mass</i> →			
<i>moles</i> →			
÷ <i>smallest value</i> →			
<i>mole ratio</i> →			

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Empirical Formula

5. (OYO) Analysis of 20.0g of a compound containing only calcium and bromine indicates that 4.00g of calcium are present. What is the empirical formula of the compound formed?
(Ans: CaBr_2)

<i>mass</i> →		
÷ <i>atomic mass</i> →		
<i>moles</i> →		
÷ <i>smallest value</i> →		
<i>mole ratio</i> →		

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Molecular Formulas

Molecular Formula (MF) - the number and types of elements needed to make 1 molecule.

• $MF = n(EF)$ and $n = MM/EM$

- MF: molecular formula
- n: ratio between the molecular and empirical mass
- EF: empirical formula
- MM: molecular mass
- EM: empirical formula mass

glucose

EF:

MF

titanium (IV) sulfide

EF:

Sample WS#3 - Molecular Formula Practice Problems

1. In an previous sample problem, the empirical formula of a compound of phosphorus and oxygen was found to be P_2O_5 . Experimentation shows that the *MM* of this compound is 283.89g/mol. What is the compound's molecular formula?

(Ans: P_4O_{10})

Given
 $EF = P_2O_5$
 $MM = 283.89 \text{ g/mol}$

NTK
 $MF = n(EF)$
 $EM = 2P + 5(O)$
 $= 2(30.97) + 5(16) = 141.94 \text{ g/mol}$
 $n = MM/EM$

UNK
 $MF = P_4O_{10}$

SOLVE:

$$n = \frac{283.89}{141.94} = 2$$

$$MF = n(P_2O_5) = 2(P_2O_5) = P_4O_{10}$$

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Molecular Formulas

2. Determine the molecular formula of the compound with the empirical formula CH and a formula mass of 78.110amu. (Ans: C₆H₆ (benzene))

$$\begin{aligned} \text{EF} &= \text{CH} \\ \text{MM} &= 78.110 \text{ amu}^* \end{aligned}$$

$$\begin{aligned} \text{EM} &= 12.01 + 1.01 = 13.02 \\ \text{MF} &= n(\text{EF}) = n(\text{CH}) \\ n &= \frac{\text{MM}}{\text{EM}} = \frac{\text{FM (amu)}}{\text{EM (amu)}} \end{aligned}$$

$$\text{MF} = \text{C}_6\text{H}_6$$

Solve:

$$n = \frac{78.110}{13.02} = 6$$

$$\text{MF} = 6(\text{CH}) = \boxed{\text{C}_6\text{H}_6}$$

* really formula mass, but we can use it has molar mass

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Molecular Formulas

3. A sample of a compound with a formula mass of 34.00 amu is found to consist of 0.44g H and 6.92g O. Find its molecular formula. (Ans: H₂O₂)

Given
0.44g H
6.92g O
FM: 34.00 amu*

* treat like MM

NTK
EF = HO
EM = (1.01) + 16 = 17.01
 $n = \frac{MM}{EM} = \frac{FM}{EM}$
MF = n(EF)

UNK
MF = H₂O₂

<u>Solve:</u>	H	O
mass →	<u>0.44</u>	<u>6.92</u>
÷ atomic mass →	<u>1.01</u>	<u>16</u>
moles →	<u>0.436</u>	<u>0.433</u>
÷ smallest value →	<u>0.433</u>	<u>0.433</u>
mole ratio →	≈ 1	1
	X = 1	Y = 1

$n = \frac{FM}{EM} = \frac{34}{17} = 2$
MF = n(EF) = 2(HO)
= H₂O₂

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Molecular Formulas

4. (OYO) Determine the molecular formula when the compound contains 64.27% carbon, 7.19% hydrogen, and 28.54% oxygen with an experimentally determined molar mass of 168.19.
(Ans: $C_9H_{12}O_3$)

<i>mass</i> →			
÷ <i>atomic mass</i> →			
<i>moles</i> →			
÷ <i>smallest value</i> →			
<i>mole ratio</i> →			

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Reaction Stoichiometry

Stoichiometry

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Reaction Stoichiometry

- The use of a balanced equation, quantities of chemical species, and mole ratios to calculate the amount of an unknown substance.
- Keys to success
 - balanced equation
 - correct *MM* of compounds
 - using mole ratios between chemical species
 - conversion of moles to mass and mass to moles
- Must obey the Law of Conservation of Mass

Stoichiometry

Mole to Mole

Process (Mole to Mole)

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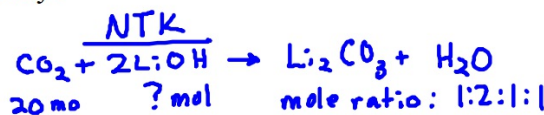
- (1) Balanced chemical equation and mole ratios
- (2) Identify given (circle) and unknown (box).
- (3) Using the balanced equation, write in given with label under known.
- (4) Using the balanced equation, write a (?) with label under unknown.
- (5) Above balanced equation, draw arrow from given to unknown.
- (6) Use molar ratios to solve for unknown

$$\frac{\text{moles}_{\text{gvn}}(\text{formula})}{\text{coefficient}_{\text{gvn}}(\text{formula})} = \frac{\text{coefficient}_{\text{unk}}(\text{formula})}{\text{coefficient}_{\text{gvn}}(\text{formula})} \times \text{moles}_{\text{unk}}$$

Stoichiometry Sample WS#4 - Mole and Mass Stoichiometry Practice Problem:

1. In a spacecraft, the carbon dioxide exhaled by astronauts can be removed by its reaction with lithium hydroxide, LiOH, according to the following chemical equation. $\text{CO}_2(\text{g}) + 2\text{LiOH}(\text{s}) \rightarrow \text{Li}_2\text{CO}_3(\text{s}) + \text{H}_2\text{O}(\text{l})$. How many moles of lithium hydroxide are required to react with 20mol of CO_2 , the average amount exhaled by a person each day? (Ans: 40 mol LiOH)

Gvn
20 moles CO_2



UNK
moles LiOH = 40 mol

Solve:

$$\frac{20 \text{ moles } \text{CO}_2}{1 \text{ CO}_2} \times \frac{2 \text{ LiOH}}{1 \text{ CO}_2} = \boxed{40 \text{ moles LiOH}}$$

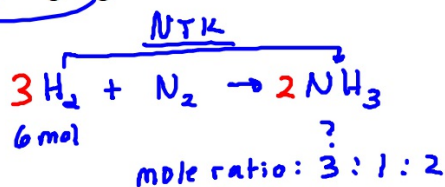
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Mole to Mole

$$\frac{\text{moles}_{\text{gvn}}(\text{formula})}{\text{coefficient}_{\text{gvn}}(\text{formula})} = \frac{\text{coefficient}_{\text{unk}}(\text{formula})}{\text{coefficient}_{\text{gvn}}(\text{formula})} \text{ moles}_{\text{unk}}$$

2. Ammonia, NH₃, is widely used as a fertilizer and in many household cleaners. How many moles of ammonia are produced when 6 mol of hydrogen gas react with an excess of nitrogen gas. (Ans: 4 mol NH₃)

Gvn
6 mol H₂



UNK
moles NH₃ = 4 mol NH₃

Solve:

$$\frac{6 \text{ mol H}_2}{3 \text{ H}_2} \left| \frac{2 \text{ NH}_3}{3 \text{ H}_2} \right. = \boxed{4 \text{ mol NH}_3}$$

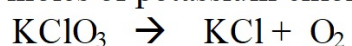
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Mole to Mole

$$\frac{\text{moles}_{\text{gvn}}(\text{formula})}{\text{coefficient}_{\text{gvn}}(\text{formula})} = \frac{\text{coefficient}_{\text{unk}}(\text{formula})}{\text{coefficient}_{\text{gvn}}(\text{formula})} \times \text{moles}_{\text{unk}}$$

3. (OYO) The decomposition of potassium chlorate, KClO_3 , is used as a source of oxygen in the laboratory. How many moles of potassium chlorate are needed to produce 15 mol of oxygen?



(Ans: 10. mol KClO_3)

Stoichiometry
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Mass to Mass

Process (Mass to Mass)

- (1) Balanced chemical equation and mole ratios
- (2) Identify given (circle) and unknown (box).
- (3) Using the balanced equation, write in given with label under known.
- (4) Using the balanced equation, write a (?) with label under unknown.
- (5) Above balanced equation, draw arrow from given to unknown.
- (6) Divide given by its *MM*.
- (7) Use molar ratios to solve for unknown
- (8) Multiply by *MM* of unknown. Write answer in correct s.f.

$$\frac{\text{mass}_{\text{gvn}}(\text{formula})}{\text{MM}_{\text{gvn}}} \times \frac{1 \text{ mol}}{\text{coefficient}_{\text{gvn}}(\text{formula})} \times \frac{\text{coefficient}_{\text{unk}}(\text{formula})}{1 \text{ mol}} \times \text{MM}_{\text{unk}} = \text{grams}_{\text{unk}}$$

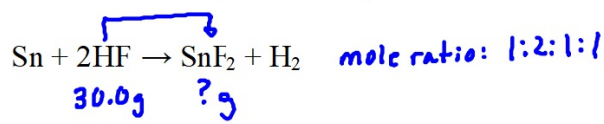
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Mass to Mass

$$\frac{\text{mass}_{\text{gvn}}(\text{formula})}{\text{MM}_{\text{gvn}}} \times \frac{1 \text{ mol}}{\text{coefficient}_{\text{gvn}}(\text{formula})} = \frac{\text{coefficient}_{\text{unk}}(\text{formula})}{1 \text{ mol}} \times \text{MM}_{\text{unk}} = \text{grams}_{\text{unk}}$$

Mass to Mass Practice Problems

4. Tin (II) fluoride, SnF₂, is used in some toothpaste. It is made by the reaction of tin with hydrogen fluoride according to the following equation. Sn + 2HF → SnF₂ + H₂



How many grams of SnF₂ are produced from the reaction of 30.00g of HF with Sn? (Ans: 117.5g SnF₂)

Gvn
30.00g HF

NTK

UNK
mass SnF₂ = 117.5g

mm HF = 20.01 g/mol
MM SnF₂ = 156.71 g/mol

Solve:

$$\frac{30.00 \text{ g HF}}{20.01 \text{ g}} \times \frac{1 \text{ mol}}{2 \text{ HF}} \times \frac{1 \text{ SnF}_2}{1 \text{ mol}} \times \frac{156.71 \text{ g}}{1 \text{ mol}} = 117.5 \text{ g SnF}_2$$

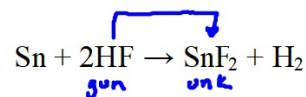
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Mass to Mass

Mass to Mass Practice Problems

4. Tin (II) fluoride, SnF₂, is used in some toothpaste. It is made by the reaction of tin with hydrogen fluoride according to the following equation. Sn + 2HF → SnF₂ + H₂
How many grams of SnF₂ are produced from the reaction of 30.00g of HF with Sn? (Ans: 117.5g SnF₂)

$$\text{mass}_{\text{gvn}} \times \left[\frac{(\text{coeff}_{\text{unk}})(MM_{\text{unk}})}{(\text{coeff}_{\text{gvn}})(MM_{\text{gvn}})} \right]$$



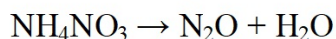
$$30.00\text{g HF} \left(\frac{(1) 156.71\text{g}}{(2) 20.01\text{g}} \right) = \boxed{117.5\text{g SnF}_2}$$

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Mass to Mass

5. Laughing gas (nitrous oxide, N₂O) is sometimes used as an anesthetic in dentistry. It is produced when ammonium nitrate is decomposed according to the following reaction.



(a) How many grams of NH₄NO₃ are required to produce 33.0g of N₂O? (Ans: 60.0g NH₄NO₃)

(b) How many grams of water are produced in this reaction? (Ans: 27.0g H₂O)

Given
33.0g N₂O

$$\begin{array}{ccc} \text{NTR} & & \text{(B)} \\ \text{(A)} & & \\ \text{NH}_4\text{NO}_3 \rightarrow \text{N}_2\text{O} + 2\text{H}_2\text{O} \\ \text{? g} & & \text{33.0g} \quad \text{? g} \end{array}$$

UNK
(A) mass_{NH₄NO₃} = 60.0 g
(B) mass_{H₂O} = 27.0 g

MM_{NH₄NO₃} = 80.06 g/mol
MM_{N₂O} = 44.02 g/mol
MM_{H₂O} = 18.02 g/mol

Solve:

(A) $33.0\text{g} \left(\frac{(1)(80.06)}{(1)(44.02)} \right) = \boxed{60.0\text{g NH}_4\text{NO}_3}$

(B) $33.0\text{g} \left(\frac{(2)(18.02)}{(1)(44.02)} \right) = \boxed{27.0\text{g H}_2\text{O}}$

$$\text{mass}_{\text{gvn}} \times \left[\frac{(\text{coeff}_{\text{unk}})(\text{MM}_{\text{unk}})}{(\text{coeff}_{\text{gvn}})(\text{MM}_{\text{gvn}})} \right]$$

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Mass to Mass

$$\text{mass}_{\text{gvn}} \times \left[\frac{(\text{coeff}_{\text{unk}})(MM_{\text{unk}})}{(\text{coeff}_{\text{gvn}})(MM_{\text{gvn}})} \right]$$

6. (OYO) When copper metal is added to silver (I) nitrate in solution, silver metal and copper (II) nitrate are produced. What mass of silver is produced from 100.g of Cu? (Ans: 339g Ag)

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Mass to Mass

$$\text{mass}_{\text{gvn}} \times \left[\frac{(\text{coeff}_{\text{unk}})(MM_{\text{unk}})}{(\text{coeff}_{\text{gvn}})(MM_{\text{gvn}})} \right]$$

7. (OYO) What mass of aluminum is produced by the decomposition of 5.0kg of Al_2O_3 ?

(Ans: 2.6kg of Al)

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Mass to Mass

Limiting reactant (LR) - the reactant A₁ or A₂ that limits the amount of B.

Excess reactant (ER) - the reactant A₁ or A₂ that remains after all of the LR is consumed.

Jessica wants to make cookies for her classmates and her recipe makes 24 cookies per batch. She has 120 classmates. How many batches does she need to make? The recipe calls for 2 cups of sugar, 4 cups of flour, 2 eggs, 1 tablespoon of vanilla extract, and 1 cup of semi-sweet chocolate chips. When Jessica gathers all of the ingredients, she has 13 cups of sugar, 29 cups of flour, 11 eggs, 11 tablespoons of vanilla, and 7 cups of chocolate chips. Will Jessica be able to make any extra cookies? How many?

	Need	Have	Extra			
2 cups sugar	x 5 = 10	13	3	1 egg	x $\frac{1 \text{ batch}}{2 \text{ eggs}}$	x $\frac{24 \text{ cookies}}{1 \text{ batch}} = 12 \text{ cookies}$
4 cups flour	x 5 = 20	29	9			
2 eggs	x 5 = 10	11	1			
1 Tbsp	x 5 = 5	11	6			
1 cup chips	x 5 = 5	7	2			
				Total cookies = 120 + 12 = 132 cookies		

$$120 \text{ cookies} \times \frac{1 \text{ batch}}{24 \text{ cookies}} = 5 \text{ batches of cookies}$$

Stoichiometry
Topic#10

Limiting/Excess Reactant

A₁ to Unknown Moles

$$\frac{A_1(\text{moles})}{\text{coefficient}_{\text{unk}}} = \text{moles of unk}$$

The LR produces the LESSER of the two calculated unknown masses.

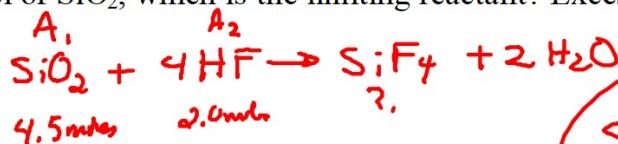
A₂ to Unknown Moles

$$\frac{A_2(\text{moles})}{\text{coefficient}_{\text{unk}}} = \text{moles of unk}$$

The excess reactant produces the GREATER of the two calculated unknown masses

Stoichiometry Sample WS#5 - Limiting/Excess Reactant Practice Problem:

1. Silicon dioxide (quartz) is usually quite unreactive but reacts readily with hydrogen fluoride according to the following equation: $\text{SiO}_2(s) + 4\text{HF}(g) \rightarrow \text{SiF}_4(g) + 2\text{H}_2\text{O}(l)$. If 2.0 mol of HF are exposed to 4.5 mol of SiO_2 , which is the limiting reactant? Excess reactant? (Ans: HF/ SiO_2)



$$(A_1) \quad 4.5 \text{ moles SiO}_2 \times \frac{\text{SiF}_4}{\text{SiO}_2} = 4.5 \text{ moles SiF}_4$$

SiO_2 is ER

$$(A_2) \quad 2.0 \text{ moles HF} \times \frac{\text{SiF}_4}{4\text{HF}} = 0.5 \text{ moles SiF}_4$$

HF is LR

Stoichiometry

Topic#10

Limiting/Excess Reactant

A₁ to Unknown Moles

$$\frac{A_1(\text{moles})}{A_1 \text{ coefficient}_{\text{gvn}}} \times \text{coefficient}_{\text{unk}} = \text{moles of unk}$$

A₂ to Unknown Moles

$$\frac{A_2(\text{moles})}{A_2 \text{ coefficient}_{\text{gvn}}} \times \text{coefficient}_{\text{unk}} = \text{moles of unk}$$

2. Some rocket engines use a mixture of hydrazine, N₂H₄, and hydrogen peroxide, H₂O₂, as the propellant. The reaction is given by the following equation: N₂H₄ + 2H₂O₂ → N₂ + 4H₂O.

- (a) Which is the LR in this reaction when 0.750mol of N₂H₄ is mixed with 0.500mol of H₂O₂? (Ans: H₂O₂)
 (b) How much of the ER, in moles, remains unchanged? (Ans: 0.500mol N₂H₄)
 (c) How much of each product, in moles, is formed? (Ans: 0.250mol N₂ and 1.00mol H₂O)

NTL

Given
 0.750 mol N₂H₄
 0.500 moles H₂O₂

N₂H₄ + 2H₂O₂ → N₂ + 4H₂O
 0.750 mol 0.500 mol

(A) 0.750 mol N₂H₄ $\left(\frac{(4) \text{ H}_2\text{O}}{(1) \text{ N}_2\text{H}_4} \right) = 3.00 \text{ moles H}_2\text{O}$ (ER N₂H₄)

0.500 mol H₂O₂ $\left(\frac{(4) \text{ H}_2\text{O}}{(2) \text{ H}_2\text{O}_2} \right) = 1.00 \text{ moles H}_2\text{O}$ (LR H₂O₂)

Stoichiometry

Topic#10

Limiting/Excess Reactant

A₁ to Unknown Moles

$$\frac{A_1(\text{moles})}{A_1 \text{ coefficient}_{\text{gvn}}} \cdot \text{coefficient}_{\text{unk}} = \text{moles of unk}$$

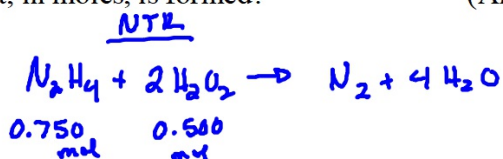
A₂ to Unknown Moles

$$\frac{A_2(\text{moles})}{A_2 \text{ coefficient}_{\text{gvn}}} \cdot \text{coefficient}_{\text{unk}} = \text{moles of unk}$$

2. Some rocket engines use a mixture of hydrazine, N₂H₄, and hydrogen peroxide, H₂O₂, as the propellant. The reaction is given by the following equation: N₂H₄ + 2H₂O₂ → N₂ + 4H₂O.

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 (c) How much of each product, in moles, is formed? (Ans: 0.250mol N₂ and 1.00mol H₂O)

Given
 0.750 mol N₂H₄
 0.500 moles H₂O₂



(B) ER = ?

$$0.500 \text{ mol } \cancel{\text{H}_2\text{O}_2} \left(\frac{(1) \text{N}_2\text{H}_4}{(2) \cancel{\text{H}_2\text{O}_2}} \right) = 0.250 \text{ mol } \text{N}_2\text{H}_4 \text{ (used)}$$

total - used = excess
 0.750 - 0.250 = 0.500 mol N₂H₄ in excess

Stoichiometry

Topic#10

Limiting/Excess Reactant

A₁ to Unknown Moles

$$\frac{A_1(\text{moles})}{A_1 \text{ coefficient}_{\text{gvn}}} \times \text{coefficient}_{\text{unk}} = \text{moles of unk}$$

A₂ to Unknown Moles

$$\frac{A_2(\text{moles})}{A_2 \text{ coefficient}_{\text{gvn}}} \times \text{coefficient}_{\text{unk}} = \text{moles of unk}$$

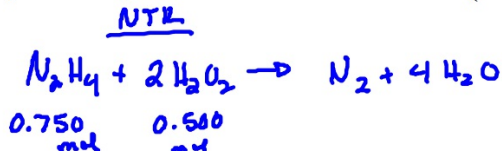
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(b) How much of the ER, in moles, remains unchanged? (Ans: 0.500mol N₂H₄)

(c) How much of each product, in moles, is formed? (Ans: 0.250mol N₂ and 1.00mol H₂O)

Gvn
0.750 mol N₂H₄
0.500 moles H₂O₂



(c) use LR (0.500 mol H₂O₂)

$$\text{N}_2 : 0.500 \text{ mol H}_2\text{O}_2 \left(\frac{(1)(\text{N}_2)}{(2)(\text{H}_2\text{O}_2)} \right) = \boxed{0.250 \text{ mol N}_2}$$

$$\text{H}_2\text{O} : 0.500 \text{ mol H}_2\text{O}_2 \left(\frac{(4)(\text{H}_2\text{O})}{(2)(\text{H}_2\text{O}_2)} \right) = \boxed{1.00 \text{ mol H}_2\text{O}}$$

Stoichiometry

Topic#10

Limiting/Excess Reactant

A₁ to Unknown (grams)

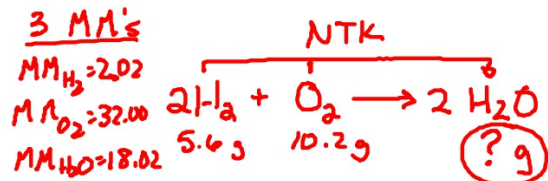
A₂ to Unknown (grams)

$$\frac{A_1(\text{grams})}{\text{coefficient}_{A_1} \times MM_{A_1}} = ? \text{ grams of unk}$$

$$\frac{A_2(\text{grams})}{\text{coefficient}_{A_2} \times MM_{A_2}} = ? \text{ grams of unk}$$

3. If 5.6g of H₂ is reacted with 10.2g O₂, what mass of H₂O will be produced? (Ans: 11g H₂O)

Given
A₁: 5.6g H₂
A₂: 10.2g O₂



Unk
mass_{H₂O} = 11.5 g

Solve:

$$A_1: 5.6g \left(\frac{(2)(18.02)}{(2)(2.02)} \right) = 50.0g \text{ H}_2\text{O}$$

$$A_2: 10.2g \left(\frac{(2)(18.02)}{(1)(32.00)} \right) = \boxed{11.5g \text{ H}_2\text{O}}$$

* smallest amount produced is the correct answer.

Stoichiometry

Topic#10

Limiting/Excess Reactant

A₁ to Unknown (grams)

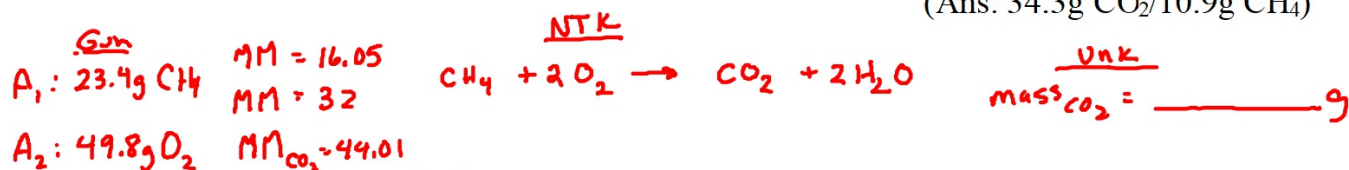
$$\frac{A_1(\text{grams})}{\text{coefficient}_{A_1} \times \text{MM}_{A_1}} = ? \text{ grams of unk}$$

A₂ to Unknown (grams)

$$\frac{A_2(\text{grams})}{\text{coefficient}_{A_2} \times \text{MM}_{A_2}} = ? \text{ grams of unk}$$

4. What mass of CO₂ will be produced from the combustion of 23.4g CH₄ with 49.8g O₂? The other product is H₂O. How many grams of the excess reactant is left after the reaction?

(Ans: 34.3g CO₂/10.9g CH₄)



Solve:

$$A_1: 23.4\text{g CH}_4 \left(\frac{(1)(44.01)}{(1)(16.05)} \right) = 64.2\text{g}$$

$$A_2: 49.8\text{g O}_2 \left(\frac{(1)(44.01)}{(2)(32.00)} \right) = \textcircled{34.2\text{g}}^*$$

Stoichiometry

Topic#10

Limiting/Excess Reactant

A₁ to Unknown (grams)

$$\frac{A_1(\text{grams})}{\text{coefficient}_{A_1} \times \text{MM}_{A_1}} = ? \text{ grams of unk}$$

A₂ to Unknown (grams)

$$\frac{A_2(\text{grams})}{\text{coefficient}_{A_2} \times \text{MM}_{A_2}} = ? \text{ grams of unk}$$

4. What mass of CO₂ will be produced from the combustion of 23.4g CH₄ with 49.8g O₂? The other product is H₂O. How many grams of the excess reactant is left after the reaction?

(Ans: 34.3g CO₂/10.9g CH₄)

$\text{MM}_{\text{CH}_4} = 16.05$
 $\text{MM}_{\text{O}_2} = 32.00$
 $\text{MM}_{\text{CO}_2} = 44.01$

A
 CH₄ 23.4g
 B
 + 2 O₂ 49.8g → CO₂ ? + 2 H₂O

CH₄ O₂
 64.2 - 34.2 = 30g CO₂

excess: $30.0 \left(\frac{(1)(16.05)}{(1)(44.01)} \right) = 10.9\text{g CH}_4 \text{ in excess}$

Excess product related to ER

Stoichiometry

Topic#10

Limiting/Excess Reactant

A₁ to Unknown (grams)

$$\frac{A_1(\text{grams})}{\text{coefficient}_{A_1} \times MM_{A_1}} = ? \text{ grams of unk}$$

A₂ to Unknown (grams)

$$\frac{A_2(\text{grams})}{\text{coefficient}_{A_2} \times MM_{A_2}} = ? \text{ grams of}$$

5. (OYO) If 0.345g Sb are reacted with 0.928g of C₂ to produce Sb₂O₅, how much Sb₂O₅ can be produced? (The reaction is a synthesis reaction)
- (Ans: 0.458g Sb₂O₅)
(Ans: 0.814g O₂ extra)

Stoichiometry

Percent Yield

Topic#10

Theoretical yield - calculated using masses (calculated)

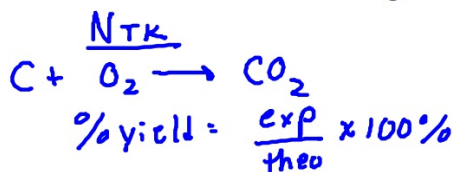
Actual yield - from an experiment (experimental)

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

Stoichiometry Sample WS#6 - Percent Yield Practice Problem:

1. Determine the percent yield for the reaction between 12.1g of C with excess O₂ if 32.2g of CO₂ is recovered? (Ans: 72.7%)

Given
12.1g C (MM=12.01)
Excess O₂
32.2g CO₂ (MM=44.01)
(experimental)



Unk

$$\% \text{ yield} = \underline{72.7\%}$$

Solve:

$$12.1\text{g C} \left(\frac{(1)(44.01)}{(1)(12.01)} \right) = 44.3\text{g CO}_2 \text{ (theo)}$$

$$\% \text{ yield} = \frac{\text{exp}}{\text{theo}} \times 100\% = \frac{32.2}{44.3} \times 100\% = \boxed{72.7\%}$$

Stoichiometry
Topic#10

Percent Yield

Theoretical yield - calculated using masses (calculated)

Actual yield - from an experiment (experimental)

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

2. If a reaction has a percent yield of 85.2% and the theoretical yield is 23.2g. What is the actual yield for the reaction? (Ans: 19.8g)

Given

$$\% \text{ yield} = 85.2\%$$
$$\text{theo} = 23.2 \text{ g}$$

NTK

$$\% \text{ yield} = \frac{\text{exp}}{\text{theo}} \times 100\%$$
$$\text{dec} = \frac{\text{exp}}{\text{theo}}$$

Unk

$$\text{actual (exp)} = \underline{19.8} \text{ g}$$

Solve:

$$0.852 = \frac{\text{exp}}{23.2}$$
$$\text{exp} = (23.2)(0.852) = \boxed{19.8 \text{ g}}$$

Stoichiometry
Topic#10

Percent Yield

Theoretical yield - calculated using masses (calculated)

Actual yield - from an experiment (experimental)

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

3. If the reaction has a percent yield 74.5% and the actual yield is 58.7g. What is the theoretical yield for this reaction? (Ans: 78.8g)

Given
%yield = 74.5%
actual(exp) = 58.7g

NRK
%yield = $\frac{\text{exp}}{\text{theo}} \times 100\%$

UNK
theo = 78.8 g

Solve: dec = $\frac{\text{exp}}{\text{theo}}$

$$(0.745) = \frac{58.7}{\text{theo}}$$

$$\text{theo} = \frac{58.7}{0.745} = \boxed{78.8\text{g}}$$

Stoichiometry

Topic#10

Percent Yield

Theoretical yield - calculated using masses (calculated)

Actual yield - from an experiment (experimental)

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

4. (OYO) Determine the percent yield for the reaction between 3.74g of Na and 4.89g C₂ if 4.34g of Na₂O₂ is recovered. (Ans: 68.5%)