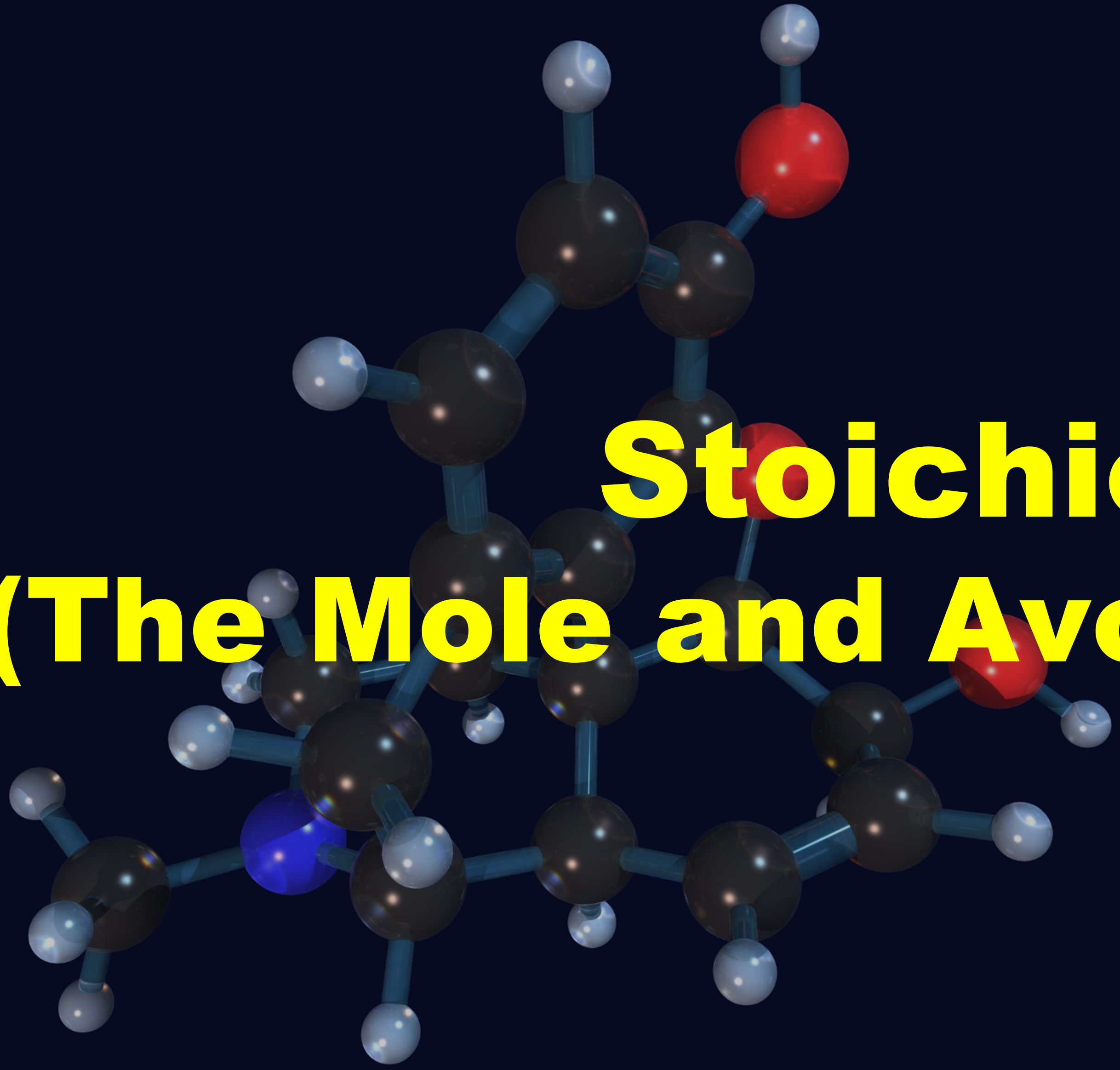


Stoichiometry
(The Mole and Avogadro's Number)

Avogadro's Number
(What in the world is a mole?)

We scientists use this thing called a ***mole*** to measure stuff. So, what is a mole? Strictly defined, a ***mole*** (also called ***Avogadro's number***) is **6.022×10^{23}**.

This is literally just a really big number, kind of like a dozen. For example, if I have a dozen eggs, then I have 12 eggs. If I have one ***mole*** of eggs, then I have **6.022×10^{23}** eggs (which, by the way, is a LOT of eggs).

So why do we care about such a bizarre number? Well, as it turns out, a mole is the number of atoms in a sample that weighs that element's atomic weight.

For example, carbon's ***atomic mass*** on the periodic table is ~12. This means that if you have a 12-gram sample of carbon, then you have **6.022×10^{23}** carbon atoms. Hydrogen's ***atomic mass*** is ~1, so if you have 1 gram of H, you have **6.022×10^{23}** hydrogen atoms.

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 **H** 1.008																	2 **He** 4.0026
3 **Li** 6.94	4 **Be** 9.0122											5 **B** 10.81	6 **C** 12.011	7 **N** 14.007	8 **O** 15.999	9 **F** 18.998	10 **Ne** 20.180
11 **Na** 22.990	12 **Mg** 24.305											13 **Al** 26.982	14 **Si** 28.085	15 **P** 30.974	16 **S** 32.06	17 **Cl** 35.45	18 **Ar** 39.948
19 **K** 39.098	20 **Ca** 40.078	21 **Sc** 44.956	22 **Ti** 47.867	23 **V** 50.942	24 **Cr** 51.996	25 **Mn** 54.938	26 **Fe** 55.845	27 **Co** 58.933	28 **Ni** 58.693	29 **Cu** 63.546	30 **Zn** 65.38	31 **Ga** 69.723	32 **Ge** 72.630	33 **As** 74.922	34 **Se** 78.97	35 **Br** 79.904	36 **Kr** 83.798
37 **Rb** 85.468	38 **Sr** 87.62	39 **Y** 88.906	40 **Zr** 91.224	41 **Nb** 92.906	42 **Mo** 95.95	43 **Tc** (98)	44 **Ru** 101.07	45 **Rh** 102.91	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.71	51 **Sb** 121.76	52 **Te** 127.60	53 **I** 126.90	54 **Xe** 131.29
55 **Cs** 132.91	56 **Ba** 137.33	57-71 *	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.84	75 **Re** 186.21	76 **Os** 190.23	77 **Ir** 192.22	78 **Pt** 195.08	79 **Au** 196.97	80 **Hg** 200.59	81 **Tl** 204.38	82 **Pb** 207.2	83 **Bi** 208.98	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** (226)	89-103 #	104 **Rf** (265)	105 **Db** (268)	106 **Sg** (271)	107 **Bh** (270)	108 **Hs** (277)	109 **Mt** (276)	110 **Ds** (281)	111 **Rg** (280)	112 **Cn** (285)	113 **Nh** (286)	114 **Fl** (289)	115 **Mc** (289)	116 **Lv** (293)	117 **Ts** (294)	118 **Og** (294)

* Lanthanide series	57 **La** 138.91	58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.36	63 **Eu** 151.96	64 **Gd** 157.25	65 **Tb** 158.93	66 **Dy** 162.50	67 **Ho** 164.93	68 **Er** 167.26	69 **Tm** 168.93	70 **Yb** 173.05	71 **Lu** 174.97
# Actinide series	89 **Ac** (227)	90 **Th** 232.04	91 **Pa** 231.04	92 **U** 238.03	93 **Np** (237)	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (262)

Molecular (Formula) Weights

To calculate a molecule's entire mass, we add up the individual masses of each of the atoms in that molecule's formula. For example, fructose has a ***molecular formula*** of $C_6H_{12}O_6$. Each **C** weighs ~12, each **H** weighs ~1, and each **O** weighs about 16. Thus, as a whole, fructose weighs:

$$(6 \times 12) + (12 \times 1) + (6 \times 16) =$$

This value (**180**) is fructose's ***molecular weight***, also called its ***formula weight***. The units of ***molecular weight*** are either **amu** (atomic mass units) or **grams/mole**. The second set of units are VERY important to know.

Question

If you have a 150-gram sample of CrO_3:

- How many moles of CrO_3 do you have?
- How many oxygen atoms do you have?
- How many grams of oxygen do you have?

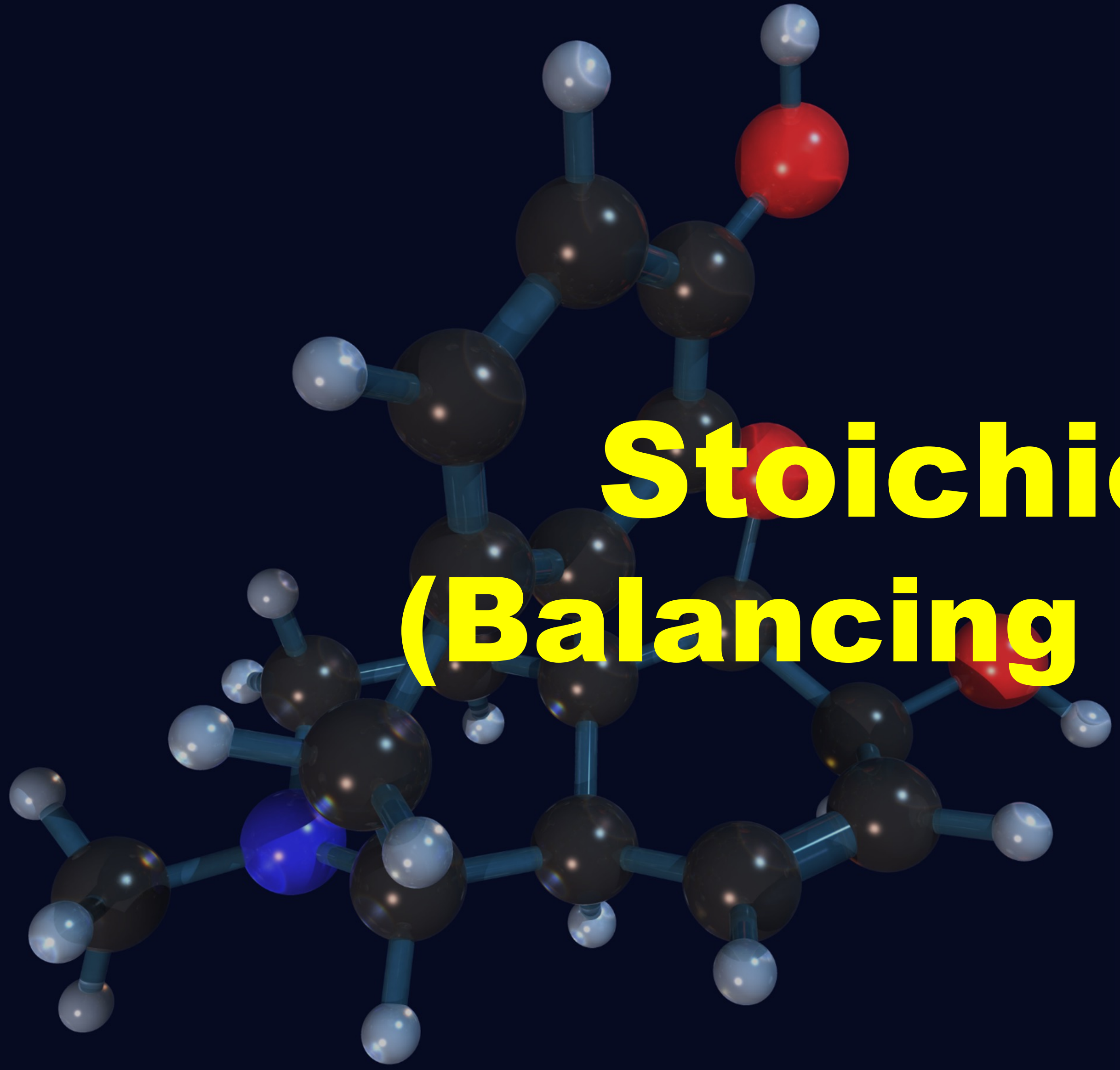
Questions

Arrange the following in order of increasing number of moles of Cl.

I. 1 g of KCl (74.6 g/mol)
II. 10 g of $MgCl_2$ (95.2 g/mol)
III. 0.1 mol of $AlCl_3$

A. II < I < III
B. III < I < II
C. I < II < III
D. II < III < I
E. III < II < I

Stoichiometry
(Balancing Reactions)



Balancing Chemical Equations

Before we can really use chemical equations, we have to balance them. We do this by adding coefficients to the left of each formula in the equation until the total number of each type of atom is equal on both sides.

When doing this it's more important to know where to start when balancing. Otherwise, you'll end up in a loop trying to balance everything.

Try to start with an element that only shows up in one compound on both sides of the equation. Then work your way toward other elements.

Balancing Chemical Equations

Balance the following equations:

$$CO\ (g) \ + \ O_2\ (g) \rightarrow CO_2\ (g)$$

$$CH_4\ (g) \ + \ Cl_2\ (g) \rightarrow CCl_4\ (l) \ + \ HCl\ (g)$$

Combustion Reactions

One last thing you should know is the general formula for ***combustion reactions.***

To explain, a ***combustion reaction*** is a rapid reaction in which a hydrocarbon (C_xH_y) or an alcohol (C_xH_yOH) reacts with oxygen (O_2). The products of ***combustion reactants*** are always CO_2 and H_2O:

$$C_2H_2OH\ (s) + 3O_2\ (g) \rightarrow 2CO_2\ (g) + 3H_2O\ (g)$$

$$2C_6H_{14}\ (g) + 19O_2\ (g) \rightarrow 12CO_2\ (g) + 14H_2O\ (g)$$

Balancing Chemical Equations

Please balance the following equations:

$$NH_4NO_3\ (s) \rightarrow N_2\ (g) \quad + \quad O_2\ (g) \quad + \quad H_2O\ (g)$$

$$Al(OH)_3\ (s) \quad + \quad H_2SO_4\ (aq) \rightarrow Al_2(SO_4)_3\ (aq) \quad + \quad H_2O\ (l)$$

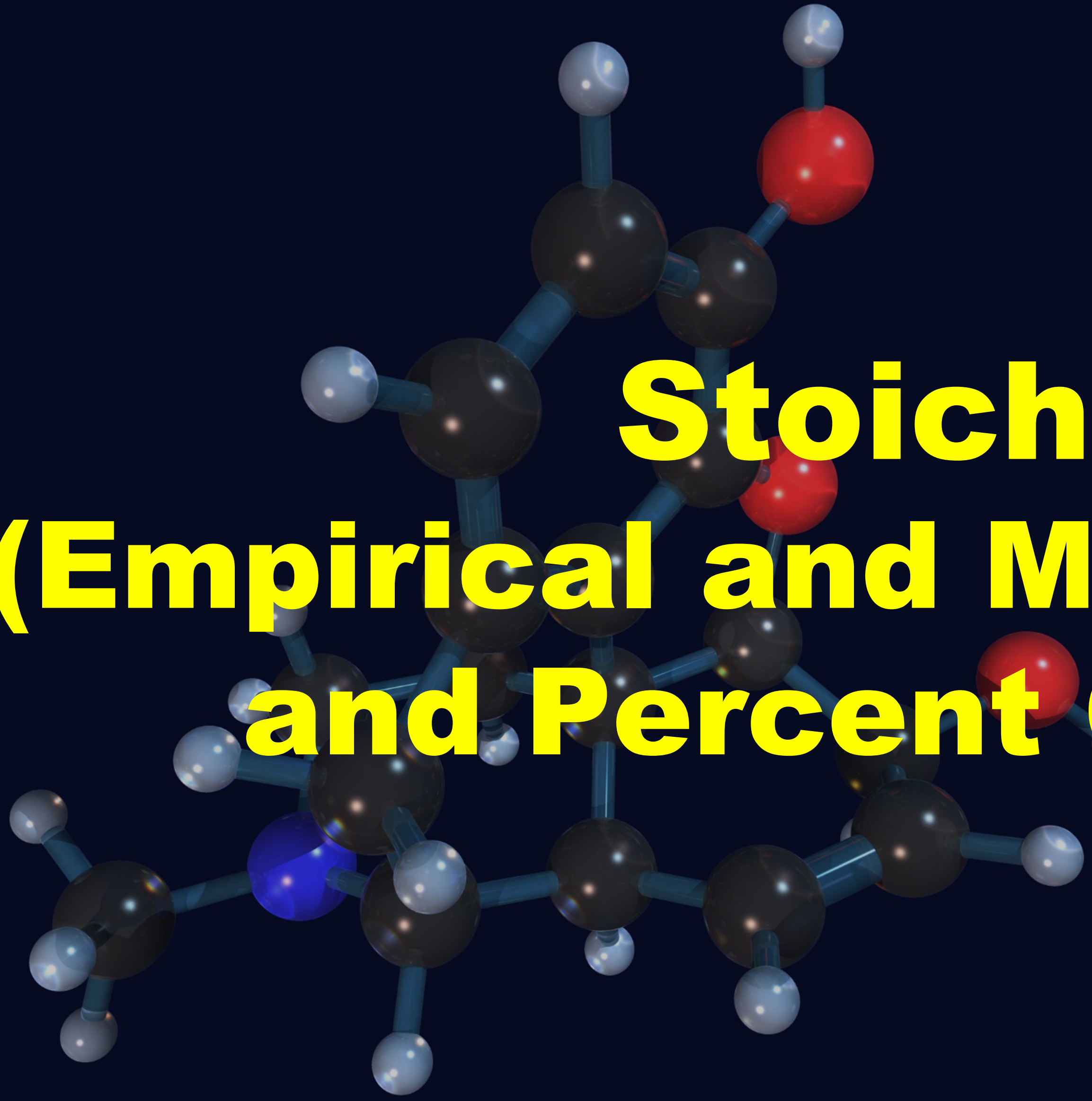
Balancing Chemical Equations

Balance the following equations and indicate whether they are combination, decomposition, or combustion reactions:

$$C_3H_6\ (g) + O_2\ (g) \rightarrow CO_2\ (g) + H_2O\ (g)$$

$$NH_4NO_3\ (s) \rightarrow N_2O\ (g) + H_2O\ (g)$$

$$PbCO_3\ (s) \rightarrow PbO\ (s) + CO_2\ (g)$$



Stoichiometry
(Empirical and Molecular Formulas and Percent Compositions)

Molecular & Empirical Formulas

Simply put, chemical formulas that show the *actual* elemental symbols and numbers of all the atoms in a molecule are called ***molecular formulas***.

Formulas that give only the *most-reduced whole number ratio* of each atom are called ***empirical formulas***.

For example, each glucose molecule *actually* contains six C atoms, twelve H atoms, and six O atoms. This means that glucose's ***molecular formula*** is $C_6H_{12}O_6$.

But, is it possible to divide each of those subscripts (6, 12, and 6) in a way that gives us a formula with a *smaller* whole-number ratio? Absolutely! If we divide them all by 6, then the formula becomes CH_2O. This is glucose's ***empirical formula***. Make sense?

So, why do we even talk about ***empirical formulas?*** The major reason is that in the lab, when we analyze an unknown sample by ***elemental analysis***, the machine (called an "elemental analyzer" or "combustion analyzer") doesn't give us our compound's ***molecular formula***.

Instead, it gives us a list of all the elements present in the sample, with their percentages. We then convert those percentages into an ***empirical formula***, and (if needed) we later use mass spectrometry to help convert that ***empirical formula*** into a ***molecular formula***.

Percent Composition

If you're given a molecule's formula, you can calculate what percent of it is hydrogen, oxygen, carbon, and so forth. For example, glucose's molecular formula is $C_6H_{12}O_6$. Each carbon has an atomic mass of ~12, each hydrogen of ~1, and each oxygen of ~16. This means that the total mass or ***molecular weight*** of glucose is:

$$(6\times12) + (12\times 1) + (6\times16) =$$

The carbon atoms contribute 72 (6×12) to this total ***molecular weight***. This means that each glucose molecule is . . .

$$(72 / 180)\times100 =$$

. . . ***<u>Forty percent carbon</u>***. We say, then, that 40% of glucose is composed (made up) of carbon. Thus, glucose's ***percent composition*** is 40% carbon, 6.67% hydrogen, and 53.33% oxygen.

Percent Composition

Percent composition questions appear in two ways on the EXAM:

- **Way 1:** You are given a molecule's formula and are asked for a specific element's percent.

- **Way 2:** You are given the mass percent of each element in a compound and are asked to determine the compound's formula.

Percent Composition (Way 1)

What is the ***percent composition*** of nitrogen in morphine, whose ***molecular formula*** is $C_{17}H_{19}NO_3$?

Answer: First, calculate the total mass of $C_{17}H_{19}NO_3$. Next, calculate the mass of the nitrogen in that formula. Last, divide the mass of nitrogen by the total mass and multiply it by 100.

Total mass of $C_{17}H_{19}NO_3$: **(17×12) + (19×1) + (14) + (3×16) = 285**
Mass of N: **14**
Percent composition of N in $C_{17}H_{19}NO_3$: **(14 / 285)×100 = 4.9%**

Percent Composition (Way 2)

What is the ***empirical formula*** of a compound that is 17.3% H and 82.7% C?

<u>Answer</u>: First, assume 100 g of total sample. By doing this, you can **convert each % into grams** directly:

17.3 g of H, 82.7 g of C

Next, **divide each number of grams by the element's atomic mass**:

H: 17.3 ÷ 1 = 17.3
C: 82.7÷ 12 = 6.89

Percent Composition (Way 2)

What is the ***empirical formula*** of a compound that is 17.3% H and 82.7% C?

Next, **divide the answers from the previous step by whichever number of the two was smaller**:

H: 17.3 ÷ 1 = 17.3
C: 82.7÷ 12 = 6.89

If the answers are close to whole numbers, go to the next step. If not, then multiply by 2 or 3 to change them into whole numbers:

H: 2.5
C: 1

Percent Composition (Way 2)

What is the ***empirical formula*** of a compound that is 17.3% H and 82.7% C?

These new numbers are the subscripts of your ***empirical formula:***

H: 5
C: 2

Empirical formula: **C_2H_5**

Empirical to Molecular Formulas

To convert an ***empirical formula*** to a ***molecular formula:***

- **<u>Step 1</u>:** Divide your ***molecular weight*** (which is obtained using ***mass spectrometry*** and will be given to you in your problem) by the weight of your ***empirical formula***.

- **<u>Step 2</u>:** Multiply each subscript in your ***empirical formula*** by the answer your got in Step 1. The new formula is your ***molecular formula***.

Empirical to Molecular Formulas

In our prior example, we calculated an ***empirical formula*** of C_2H_5. If your compound's actual ***molecular weight*** is 58.123 amu, then what is its ***molecular formula?***

Answer: Divide your ***molecular weight*** by the weight of your ***empirical formula***.

Empirical formula weight: (2×12) + (5×1) = 29
Molecular weight: 58.123
Molecular ÷ Empirical = 58.123 / 29 = 2

Multiply each subscript in your ***empirical formula*** by the answer your got in Step 1:

$$C_2H_5 \times 2 = C_4H_{10}$$

This new formula is your ***molecular formula***.

Questions

The empirical formula of a compound with molecules containing 12 carbon atoms, 14 hydrogen atoms, and 6 oxygen atoms is ____________.

A. $C_{12}H_{14}O_6$

B. $C_6H_7O_3$

C. CHO

D. CH_2O

E. C_2H_4O

Questions

Calculate the ***percent composition***, by mass, of carbon in the following compounds:

Cocaine, $C_{17}H_{21}NO_4$

Vancomycin, $C_{55}H_{75}Cl_2N_9O_{24}$

Questions

What is the ***molecular formula*** of each of the following compounds?

Empirical formula CH_2, molar mass = 84 g/mol

Empirical formula HCO_2, molar mass = 90.0 g/mol

Questions

What is the empirical formula of a compound that contains 12.1% C, 16.2% O, and 71.7% Cl?

A. $COCl$
B. $COCl_2$
C. C_2OCl
D. CO_2Cl
E. $ClCO_4$

Questions

An ore containing gold was processed so that all of the gold is converted to $AuCl_3$ (303 g/mol). If a 4.0 g sample of an ore produced 3.03 g of $AuCl_3$, what is the percentage of Au (197 g/mol) in the ore?

A. 10.0 %

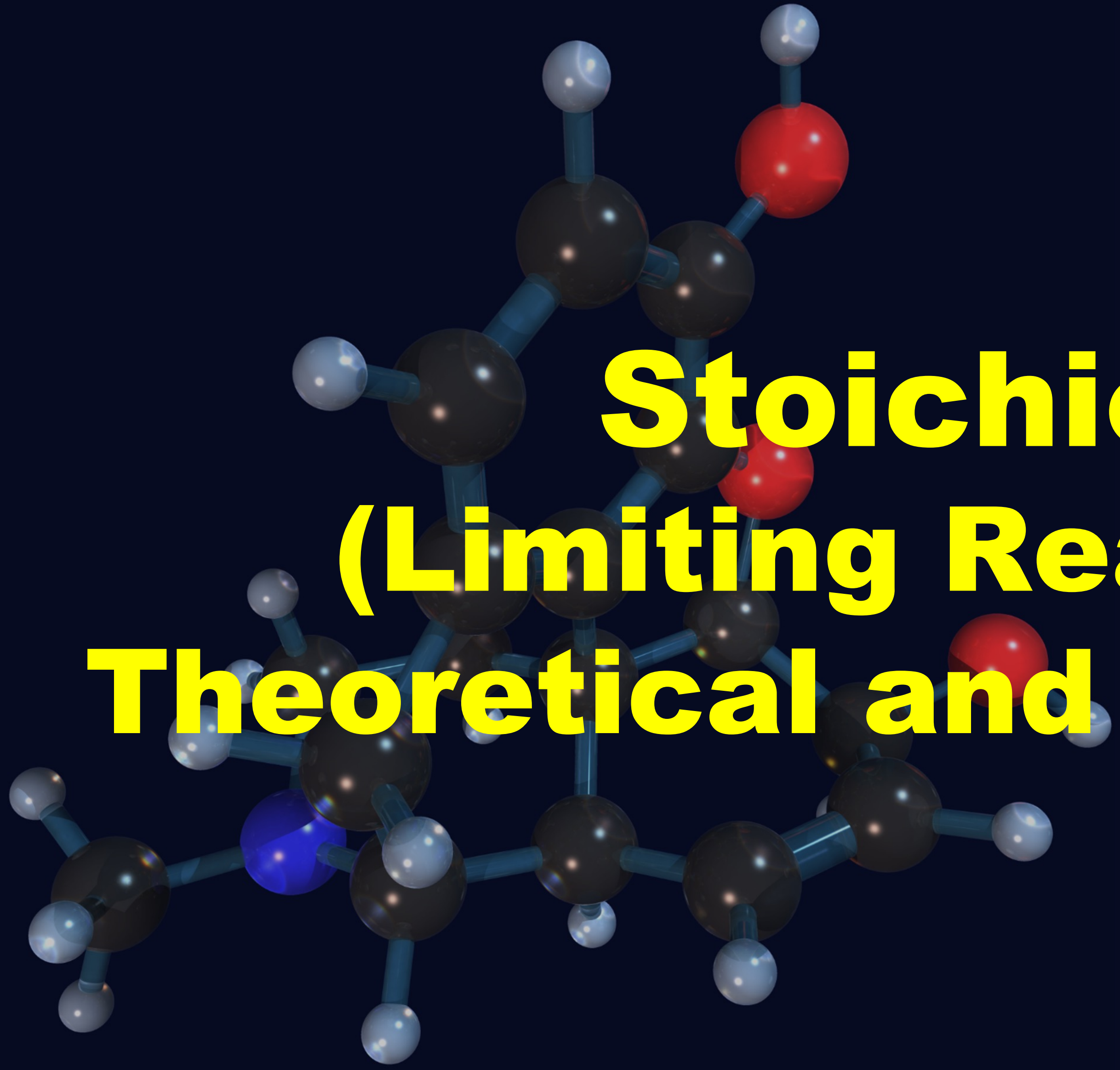
B. 49.3 %

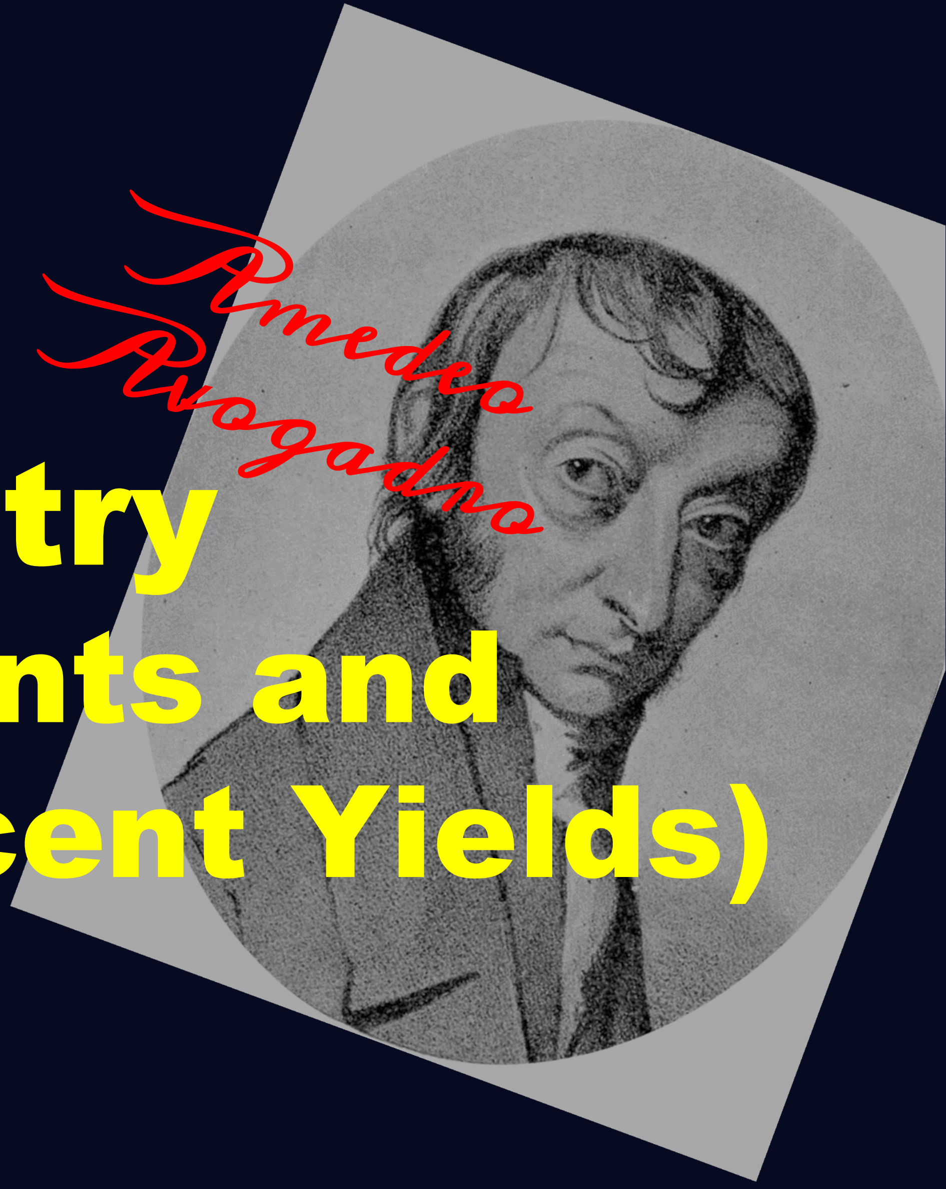
C. 75.8 %

D. 98.5 %

E. 24.6 %

Stoichiometry (Limiting Reactants and Theoretical and Percent Yields)





Product & Reactant Amounts

In real life, we can use balanced chemical equations to predict the amounts of products formed and reactants needed in a given chemical reaction, much like using a recipe from a cookbook. To do this, remember **BCC**:

Step 1: **B**alance the equation (if necessary).

Step 2: **C**onvert to moles (or millimoles).

Step 3: Use the **C**oefficients in the equation to transform the moles of things you know into moles of things you don't know.

Questions

How many moles of C_2H_6 are needed to react with excess O_2 to form 22 moles of CO_2 in the following reaction?

$$C_2H_6\ (g) \ + \ O_2\ (g) \rightarrow CO_2\ (g) \ + \ H_2O\ (g)$$

How many grams of C_2H_6 does this represent?

Questions

How many grams of CO_2 (44.01 g/mol) are produced from the complete combustion of 0.25 mol of butane, C_4H_{10} (58 g/mol)?

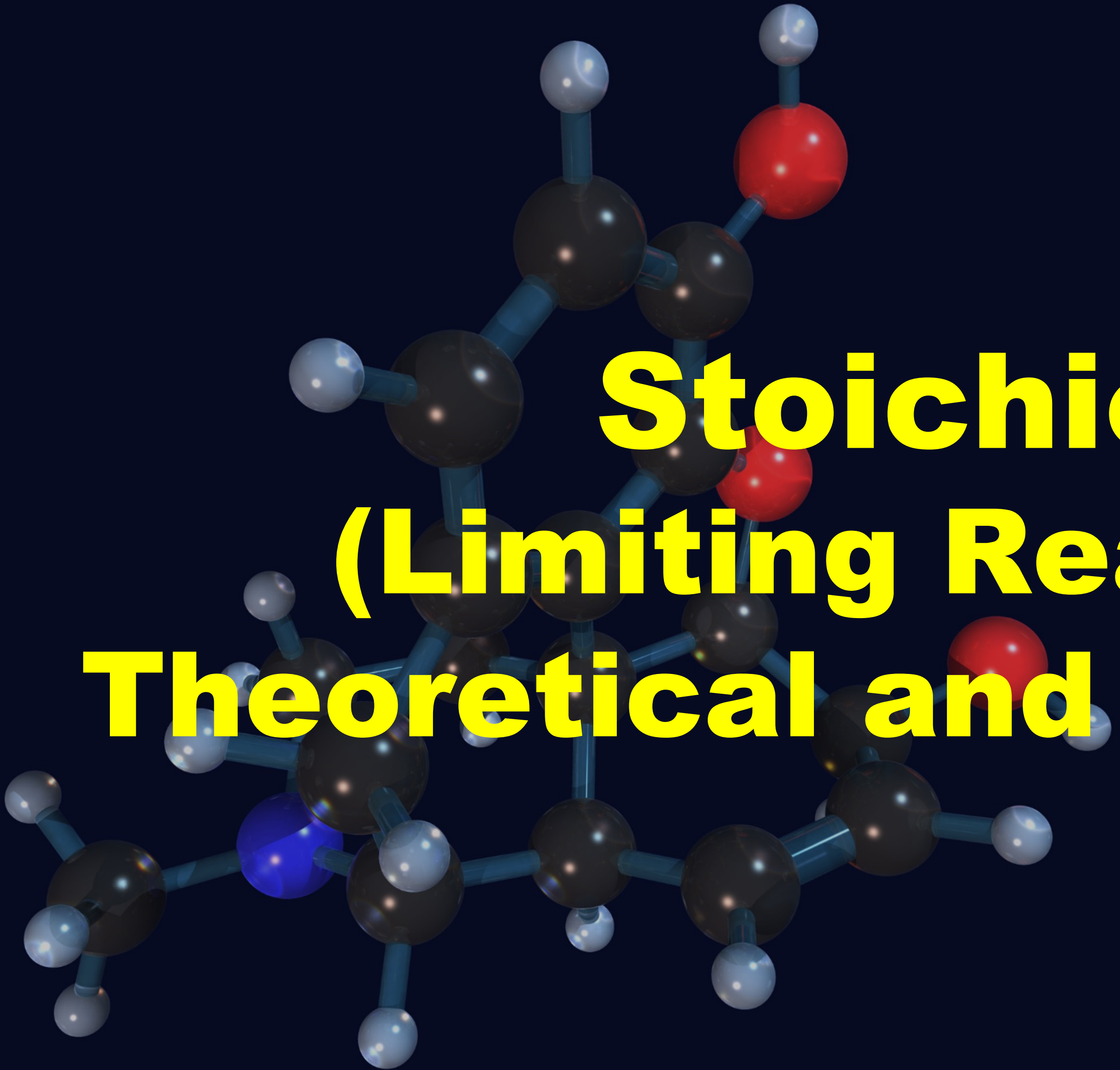
A. 11 g
B. 44 g
C. 22 g
D. 88 g
E. 15 g

Questions

How many grams of Mg (24 g/mol) are needed to react with excess HCl (36.5 g/mol) to obtain 0.90 mol of $MgCl_2$ (95.2 g/mol) if the reaction has 90% yield?

$$\mathbf{Mg + 2\,HCl \rightarrow MgCl_2 + H_2}$$

A. 22 g
B. 2.4 g
C. 24 g
D. 12 g
E. 48 g



Stoichiometry
(Limiting Reactants and Theoretical and Percent Yields)

Limiting Reactants

When doing problems like the previous example, we are sometimes asked to identify the ***limiting reactant***. Simply stated, this is <u>the reactant that runs out first</u>.

It's often VERY important to identify the ***limiting reactant*** because the amount of product you get depends on the amount of ***limiting reactant***.

If a question says that a reactant is added "in excess", it means that this reactant is NOT the ***limiting reactant***. In other words, the "excess" reactant is NOT the ***limiting reactant***. In such problems, the <u>*other*</u> reactant (the non-"excess" reactant) is the ***limiting reactant***.

excess (or “xs”) = NOT the limiting reactant

Problems that say “excess” in them

- Balance the chemical reaction
- Convert everything to moles
- Use Coefficients

$$A + B \rightarrow C$$

“xs”

limiting

How to Identify the Limiting Reactant

To identify the ***limiting reactant*** (for questions that do NOT say "excess" in them) just remember **BCDS**:

Problems that do NOT say "excess" in them

Balance the chemical reaction

Convert everything to moles

Divide: moles / coefficients

Smallest one wins!

Questions

If you react 6 grams of C_2H_6 (M.W. = 30 g/mol) with 9.6 grams of O_2 (M.W. = 32 g/mol), which reactant will run out first?

$$C_2H_6\ (g) + O_2\ (g) \rightarrow CO_2\ (g) + H_2O\ (g)$$

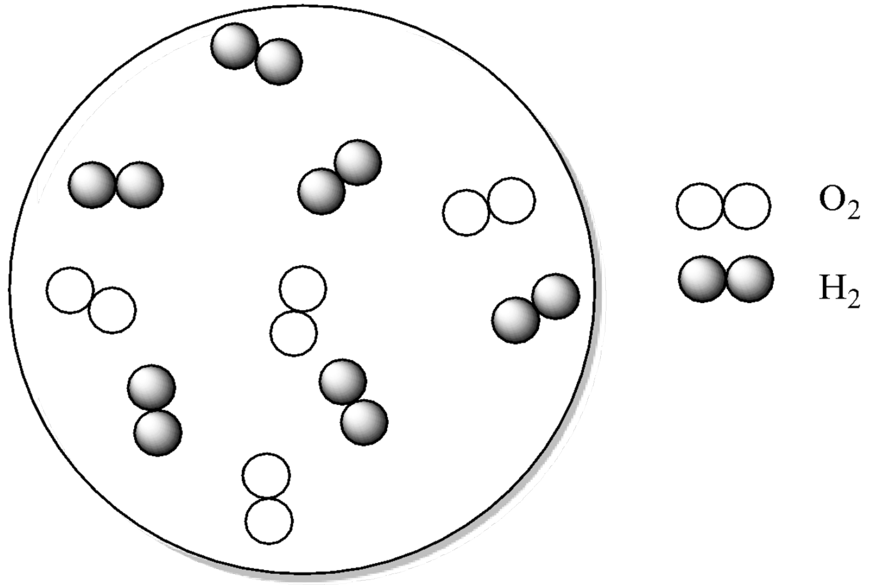
Under these conditions, how many grams of H_2O will this reaction produce?

Questions

Hydrogen gas reacts with oxygen gas to form water.

$$H_2 \quad + \quad O_2 \rightarrow H_2O$$

When the molecules in the diagram to the right react to form the maximum amount of water, what is the limiting reactant and how many molecules of water can be formed?



A. O_2 is limiting and 8 molecules of water will be formed
B. O_2 is limiting and 4 molecules of water will be formed
C. H_2 is limiting and 6 molecules of water will be formed
D. H_2 is limiting and 3 molecules of water will be formed
E. H_2 is limiting and 12 molecules of water will be formed

amount of product given in the problem

$$\% \text{ yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}}$$

amount of product calculated from your <u>limiting reactant</u>

excess (or “xs”) = NOT the limiting reactant

Problems that say “excess” in them

- Balance the chemical reaction
- Convert everything to moles
- Use Coefficients

$$\underset{\text{“xs”}}{A} + \underset{\text{limiting}}{B} \rightarrow C$$

Problems that do NOT say "excess" in them

<u>B</u>alance the chemical reaction

<u>C</u>onvert everything to moles

<u>D</u>ivide: moles / coefficients

<u>S</u>mallest one wins!

Questions

Ammonia is a major component of household cleaner. It is synthesized industrially via the Haber-Bosch process according to the reaction:

$$N_2 \quad + \quad 3\, H_2 \rightarrow \; 2\, NH_3$$

If a mixture of 14 g of N_2 and 10 g of H_2 produced 0.40 moles of ammonia, what is the percent yield of the reaction?

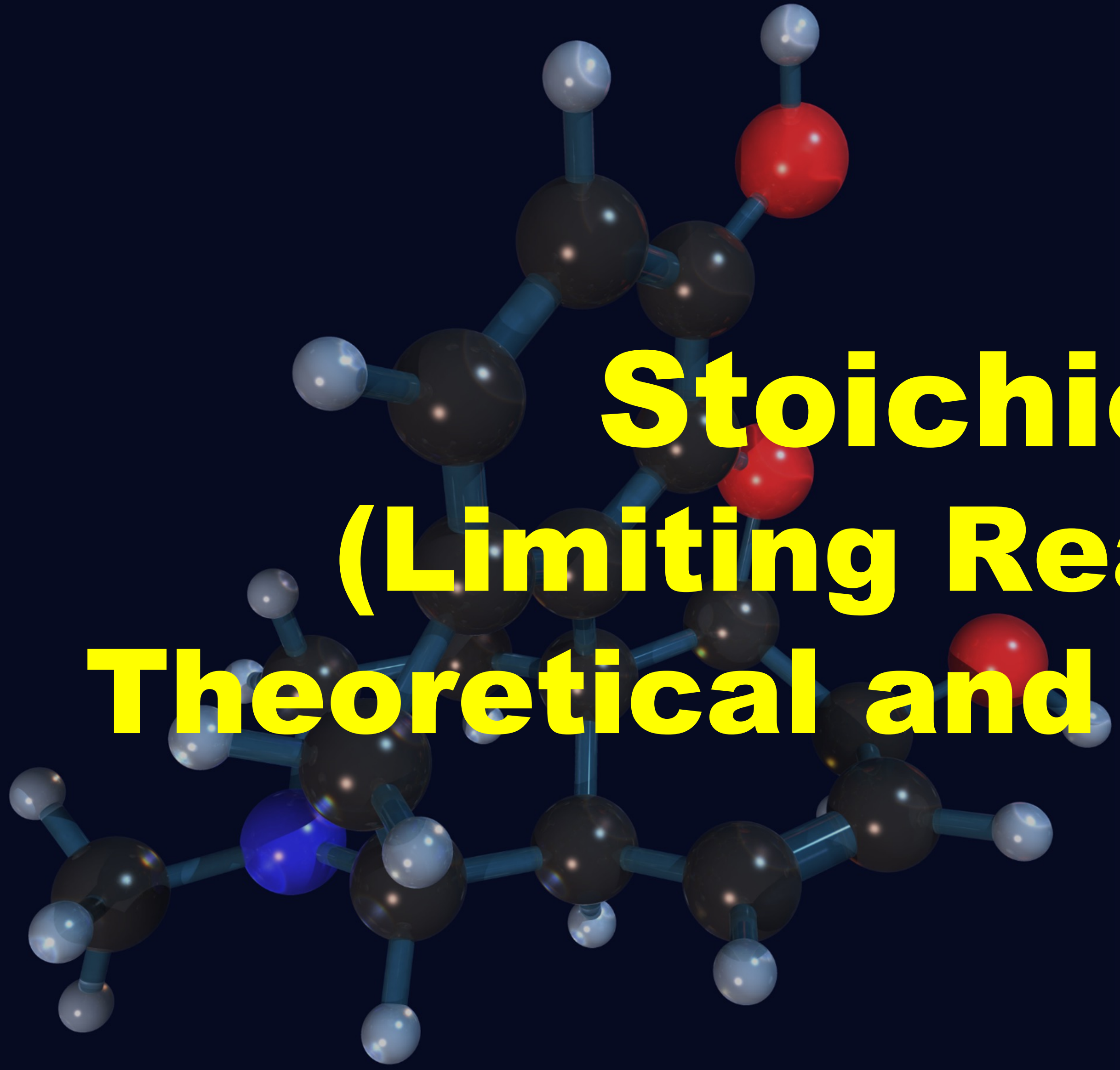
A. 24 %

B. 20 %

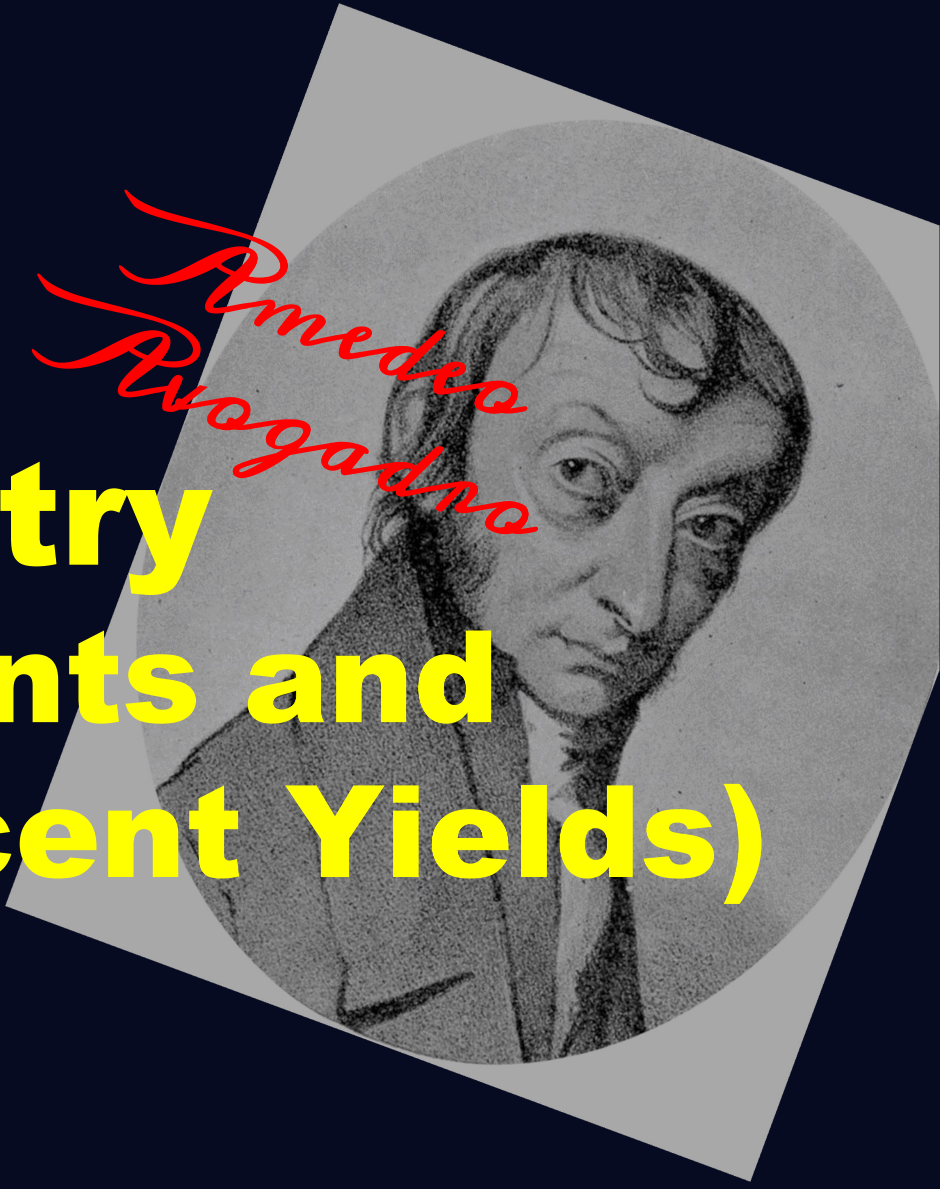
C. 80 %

D. 40 %

E. 60 %



Stoichiometry (Limiting Reactants and Theoretical and Percent Yields)



Theoretical and Percent Yields

The ***theoretical yield*** is the amount of product you *should* get in a given reaction, if everything goes perfectly. This really is the amount of product that you would *calculatedly* obtain.

So, to calculate a ***theoretical yield***, just identify the limiting reactant and then calculate—from that reactant—the amount of product you would form. That's it!

Theoretical and Percent Yields

Unfortunately, in real life, you rarely get the *actual* amount of product that you should calculatedly obtain. The amount of product that you actually get when you run a reaction is called the ***actual yield.***

Chemists often report something called a ***percent yield*** as a measure of how efficient a reaction is. Simply defined:

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

Theoretical and Percent Yields

Unfortunately, in real life, you rarely get the *actual* amount of product that you should calculatedly obtain. The amount of product that you actually get when you run a reaction is called the ***actual yield.***

Chemists often report something called a ***percent yield*** as a measure of how efficient a reaction is. I often define it as this:

$$\%\ \text{yield} = \frac{\text{what you actually got}}{\text{what you } \underline{\textit{would}} \text{ have gotten if everything went perfectly}} \times 100$$

Questions

Suppose you react 81 grams of **Al** (atomic mass = 27 g/mol) with 213 grams of **Cl_2** (M.W. = 71 g/mol), and you obtain 133.5 grams of **$AlCl_3$**, according to the following equation:

$$Al\ (s) + Cl_2\ (g) \rightarrow AlCl_3\ (s)$$

Please calculate:

- The limiting reactant
- The theoretical yield
- The percent yield
- How many grams of the excess reagent are left over, once the reaction is completed?

Questions

Iron (Fe, 55.8 g/mol) is an abundant element on Earth. An iron ore called hematite contains iron in the form of Fe_2O_3 (160 g/mol). To extract the iron from hematite, the Fe_2O_3 is reacted with carbon monoxide according to the reaction:

$$Fe_2O_3 + 3\,CO \rightarrow 2\,Fe + 3\,CO_2$$

How much hematite ore, which contains 64% Fe_2O_3 by mass, is required to produce 0.20 mol of elemental iron?

A. 17.4 g
B. 34.9 g
C. 8.72 g
D. 25.0 g
E. 50.0 g

Questions

Pentane is a component of gasoline. How many liters of air, which contains 20% O_2 by volume, is required to completely combust 2.0 L of pentane?

$$C_5H_{12}\,(g) \quad + \quad 8\ O_2\,(g) \rightarrow 5\ CO_2\,(g) \quad + \quad 6\ H_2O\,(g)$$

A. 20 L
B. 40 L
C. 60 L
D. 80 L
E. 100 L

Questions

A 6.55-gram sample of an unknown element Q reacts with excess fluorine gas to form 10.35 grams of QF_4. The unknown element is most likely which of the following?

A. P

B. S

C. Xe

D. C

E. Te