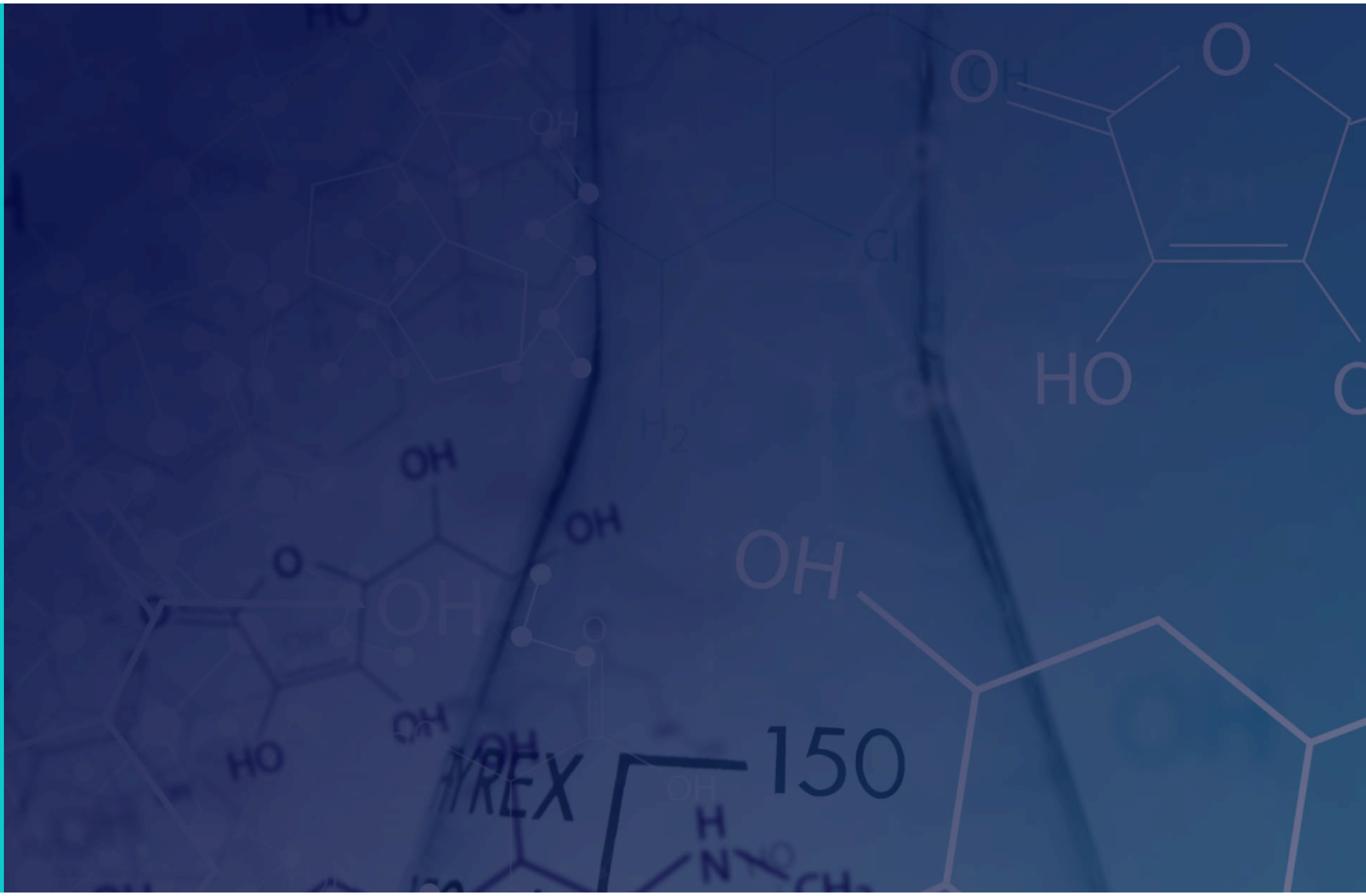

UWO

CHEM 1302

Fall 2024, Chapter 1 Notes

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1. Stoichiometry Review

1.1 Introduction to Chemical Equations and Balancing

1.1.1

Balancing Reactions

Reactions must be balanced before the stoichiometry can be analyzed.

What makes a reaction balanced?

The number of atoms of each type on the reactant side must equal the number of atoms of each type on the product side of the reaction.

Example:

$$\text{Unbalanced:} \qquad C_3H_{8(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$$

$$\text{Balanced:} \qquad C_3H_{8(g)} + 5\ O_{2(g)} \rightarrow 3\ CO_{2(g)} + 4\ H_2O_{(g)}$$

1.1.2

Balancing Chemical Reactions

Unbalanced Reaction

$$\mathbf{CH_4 + O_2 \longrightarrow CO_2 + H_2O}$$

The **subscripts** tell us the ratio of different atoms in one molecule.

Example: One CH_4 molecule has 4 hydrogen atoms and 1 carbon atom

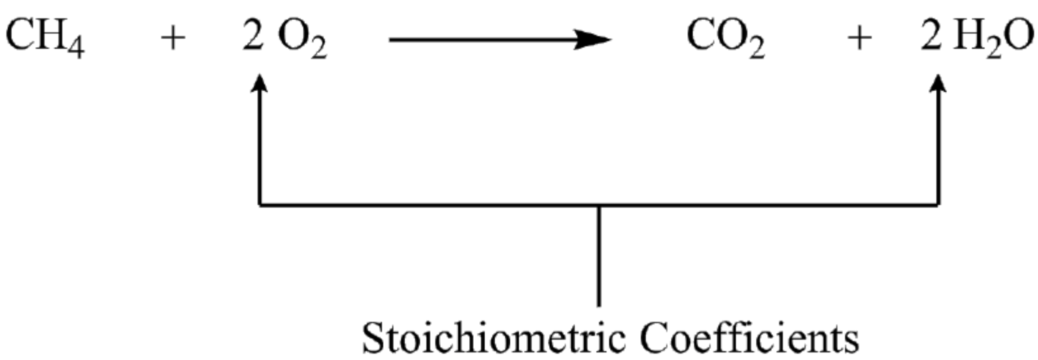
! **WATCH OUT!**

Sometimes you could be given unbalanced equations and you are expected to balance them!
We **need to use the balanced equation for calculations** in order to get the correct answer!

Steps to Balance Chemical Equations

1. Balance **elements that only appear once** on the reactant side and once on the product side (typically elements not including C, N, H, O)
2. **Group polyatomic ions** and balance them as a group (not as separate elements, *e.g.* balance NO_3 as 1 NO_3 group, not 1 N and 3 O)
3. **Balance all other elements**, starting with the least common elements to the most common elements
4. Make sure all **coefficients are whole numbers** AND are represented in the **simplest integer ratio possible**
5. **Check** to make sure all elements are balanced on each side of the chemical equation

Balanced Reaction



- The **stoichiometric coefficients** of the balanced equation tell us that we need 2 oxygen molecules for each methane (CH_4) molecule and we produce 2 waters for every methane consumed.
- It also tells us that for every 1 mole of methane used we produce 1 mole of CO_2.

WIZE CONCEPT

Multipy coefficients by subscripts to see exactly how much of an atom we have on each side.

WATCH OUT!

We want the ***lowest* possible coefficients.**

Example:

If I balanced an equation and got:

$$2MgCl_2\left(aq\right)+4NaOH\left(aq\right)\ \rightarrow 2Mg\left(OH\right)_2\left(s\right)\ +\ 4NaCl\left(aq\right)$$

The **greatest common factor** that can go into each coefficient is 2 so we need to divide each coefficient by 2 to get the lowest possible numbers!

$$MgCl_2\left(aq\right)+2NaOH\left(aq\right)\ \rightarrow Mg\left(OH\right)_2\left(s\right)\ +\ 2NaCl\left(aq\right)$$

1.1.3

Example: Balancing a Chemical Reaction

Balance the following equation:

$$4\,NO_{(aq)} \quad + \quad O_{2(g)} \quad + 2H_2O\ _{(l)} \quad \rightarrow\ 4HNO_{2(aq)}$$

WIZE CONCEPT

The physical state of each chemical compound is often included in chemical equations:

(s) = solid
(l) = liquid
(g) = gas
(aq) = aqueous

Polyatomic Ions

Note: There is no need to memorize the following table. It is just included here so you have an idea about what a polyatomic ion is!

Polyatomic Ion Cheatsheet			
Ion	**Name**	**Ion**	**Name**
NH_4^+	Ammonium	MnO_4^-	Permanganate
NO_2^-	Nitrite	CrO_4^{2-}	Chromate
NO_3^-	Nitrate	$Cr_2O_7^{2-}$	Dichromate
SO_3^{2-}	Sulfite	O_2^{2-}	Peroxide
SO_4^{2-}	Sulfate	O_2^-	Superoxide
HSO_4^-	Hydrogen sulfate*	$C_2O_4^{2-}$	Oxalate
HSO_3^-	Hydrogen sulfite	$HC_2O_4^-$	Hydrogen Oxalate
OH^-	Hydroxide	HCO_2^-	Formate
CN^-	Cyanide	$S_2O_3^{2-}$	Thiosulfate
PO_4^{3-}	Phosphate	$HS_2O_3^-$	Hydrogen Thiosulfate
HPO_4^{2-}	Hydrogen Phosphate	BrO_4^-	Perbromate
H_2PO^{4-}	Dihydrogen Phosphate	BrO_3^-	Bromate
CO_3^{2-}	Carbonate	BrO_2^-	Bromite
HCO_3^-	Hydrogen Carbonate**	BrO^-	Hypobromite
ClO^-	Hypochlorite	IO_4^-	Periodate
ClO_2^-	Chlorite	IO_3^-	Iodate
ClO_3^-	Chlorate	IO_2^-	Iodite
ClO_4^-	Perchlorate	IO^-	Hypoiodate
$C_2H_3O_2^-$	Acetate		

1.1.5

Example: Balance a Chemical Equation with Polyatomic Ions

$$(NH_4)_2SO_4\ (aq) \quad + \quad Ba(NO_3)_2\ (aq) \quad \rightarrow 2\,NH_4NO_3\ (aq) \quad + \quad BaSO_4\ (s)$$

1.1.6

Practice: Balancing the Combustion of Furan

Choose the answer that balances the reaction shown below:

$$\text{A } C_4H_4O(l) + \text{B } O_2(g) \rightarrow \text{C } CO_2(g) + \text{D } H_2O(g)$$

- A = 2; B = 9 ;C = 8; D = 4
- A = 1; B = 3 ;C = 5; D = 2
- A = 1; B = 1 ;C = 1; D = 1
- A = 4; B = 25 ;C = 20; D = 10

Step 1: Balanced C atoms

$$C_4H_4O_{(l)} + O_{2(g)} \longrightarrow 4\,CO_{2(g)} + H_2O_{(g)}$$

Step 2: Balanced H atoms

$$C_4H_4O_{(l)} + O_{2(g)} \longrightarrow 4\,CO_{2(g)} + 2H_2O_{(g)}$$

Step 3: Balanced O atoms

$$C_4H_4O_{(l)} + 4\tfrac{1}{2}O_{2(g)} \longrightarrow 4\,CO_{2(g)} + 2H_2O_{(g)}$$

$$1 + 2B \Rightarrow 10 \qquad 2B = 9 \qquad B = 4\tfrac{1}{2}$$

Step 4: Eliminate Fraction

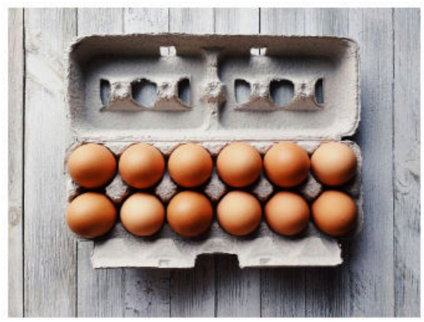
$$2\,C_4H_4O_{(l)} + 9O_2 \longrightarrow 8\,CO_{2(g)} + 4\,H_2O_{(g)}$$

1.2 Atoms, Molecules, and the Mole!

1.2.1

Moles

When someone says they want a dozen donuts, roses, or eggs we know they mean they want 12. A **mole** in chemistry also tells us the amount of something.



- Just like how a dozen of something = 12, 1 mole of something =$6.02x10^{23}$ molecules
- $6.02x10^{23}$ is referred to as **Avagadro's number (N_A)**

WIZE TIP

You should memorize Avagadro's number since it is not always provided on exams! We'll soon see that you need to know it to solve problems :)

$$\boxed{\mathbf{N_A = 6.02 \times 10^{23}}\ \text{molecules/mole}}$$

We use the unit "**moles**" to help make very small amounts more measurable. A mole is just like any other unit!

There are 2 Equations Related to Moles:

$$\boxed{n = \frac{m}{M}} \qquad \boxed{n = \frac{N}{N_A}}$$

n=# of moles

m=mass (g)

M=molar mass (g/mol)

N=# of molecules or atoms

N_A=Avagadro's number=$6.02\text{x}10^{23}$ molecules

Example #1:

If we are told that a sample of CO_2(s) weighs 11g, how many moles are present?

$m = 11\,g \quad n = ? \quad M_{CO_2(s)} = 12\,g/mol + 2(16\,g/mol)$

$n = \frac{m}{M} = \frac{11}{44} = 0.25\text{ mols} \qquad = 44\,g/mol$

Example #2

We have 2 moles of CO_2 present in our sample, how many molecules are there?

$n = 2 \text{ mols} \quad N = ? \quad N_A = 6.02 \times 10^{23}$

$$n = \frac{N}{N_A}$$

$N = N_A(n) = 2 \times 6.02 \times 10^{23} = 1.2 \times 10^{24}$ molecules

Example #3

How many oxygen atoms there in 1 mole of CO_2?

1) Find Number of Molecules In the Sample

$$n = \frac{N}{N_A} \quad N = ?$$

$$N = n N_A$$

$$= 1(6.02 \times 10^{23})$$

$$= 6.02 \times 10^{23} \text{ molecules}$$

2) Find the Number of O atoms in the Sample

In each molecule there are 2 O atoms.

Therefore to find the number of O atoms:

# of molecules in sample x 2 O atoms/molecule =# of O atoms

# of O atoms= $6.02 \times 10^{23} \times 2 = 1.2 \times 10^{24}$ atoms O

1.2.2

Example: Calculate the Number of Molecules of Ethyl Mercaptan

The volatile liquid ethyl mercaptan, $\mathrm{C_2H_6S}$, is one of the most odoriferous substances known. It is added to natural gas to make gas leaks detectable. (***d*** = 0.84 g/mL; **MW** = 62.1 g/mol)

1. How many $\mathrm{C_2H_6S}$, molecules are contained in a 3.0 μL sample?

$$N = ? \quad V = 3\mu L = 3.0\times10^{-6}\,L = 3.0\times10^{-3}\,mL \quad M = 62.1\,g/mol \quad N_A = 6.02\times10^{23}$$

$$d = 0.84\,g/mL$$

$$m = dV = (0.84\,g/mol)(3.0\times10^{-3}\,mL) = 2.52\times10^{-3}\,g$$

$$n = \frac{m}{M} = \frac{2.52\times10^{-3}\,g}{62.1\,g/mol} = 4.05797\times10^{-5}\text{ moles}$$

$$N = nN_A = (4.05797\times10^{-5}\text{ moles})(6.02\times10^{23}\text{ molecules/mols})$$
$$= 2.4\times10^{19}\text{ molecules}$$

2. In the same 3.0 μL sample, how many C and H atoms are there?

$$\# \text{ of C atoms} = 2.4 \times 10^{19} \text{ molecules} \times \frac{2 \text{ C atom}}{\text{molecules}} = 4.8 \times 10^{19} \text{ C atoms}$$

$$\# \text{ of H atoms} = 2.4 \times 10^{19} \text{ molecules} \times \frac{6 \text{ H atoms}}{\text{molecules}} = 1.4 \times 10^{20} \text{ H atoms}$$

1.2.3

Practice: Converting Mass to Number of Atoms

Calculate the number of nitrogen atoms in 2.25 g of bismuth(III) nitrate.

- a) 1.03×10^{22} atoms ✓
- b) 1.03×10^{21} atoms
- c) 3.43×10^{21} atoms
- d) 3.43×10^{22} atoms

Step 1: Determine the molar mass of Bismuth(III) Nitrate

Bi: 208.98 g/mol

N: 14.01 g/mol

O: 16.00 g/mol

$$Bi(NO_3)_3 = 208.98 + 14.01 \times 3 + 16.00 \times 9$$

$$= 395.01 \text{ g/mol}$$

Step 2: Calculate the moles

$$n = \frac{m}{M} = \frac{2.25g}{395.01 g/mol} = 0.0057 \text{ mols}$$

Step 3: Determine mol of Nitrogen

$$n_N = 0.0057 \times 3 = 0.0171 \text{ mols}$$

Step 4: Convert to # of atoms

N atoms: $0.0171 \times 6.02 \times 10^{23} \approx 1.03 \times 10^{22}$ molecules

1.2.4

Practice: Finding the Number of Moles of Iron

Calculate the number of moles of iron atoms in 14.1 g of iron oxide, Fe_2O_3

- A) 0.177 mol ✓
- B) 0.0821 mol
- C) 0.0906 mol
- D) 0.0451 mol

Step 1: Determine the molar mass of Fe_2O_3

$$M_{Fe_2O_3} = 2 \times 55.85 + 3 \times 16 = 159.70\,g/mol$$

Step 2: Calculate moles amount

$$n_{Fe_2O_3} = \frac{mass}{molar\ mass} = \frac{14.1}{159.70} \approx 0.0883\ mole$$

Step 3: Calculate the Fe atoms

$$n_{Fe} = 0.0833\ moles = \frac{2\ mol\ Fe}{1\ mol\ Fe_2O_3} = 0.1766\ moles$$

1.3 Stoichiometry

1.3.1

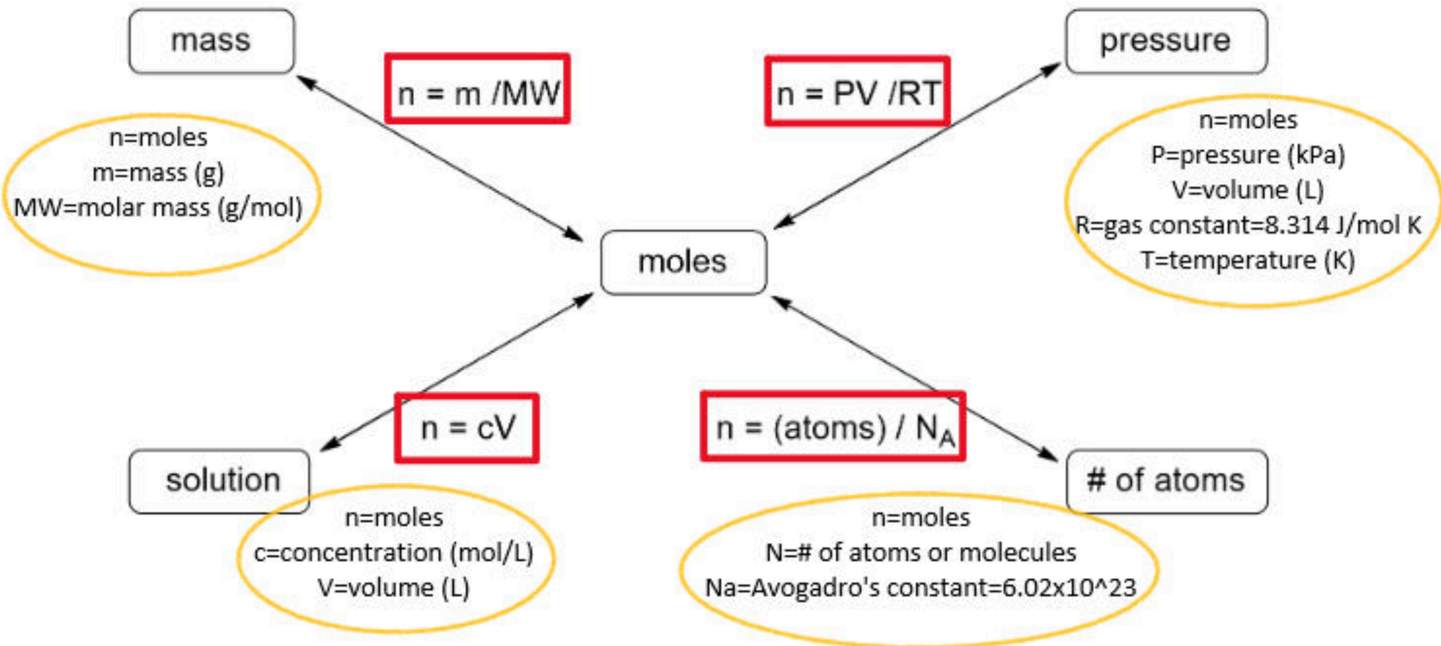
Intro to Stoichiometry

Stoichiometry allows us to predict the quantities of products or reagents across a chemical reaction.

Typically this is done in a few steps:

1. Calculating the number of moles of the reagents
2. Calculating the number of moles of the products using the stoichiometric ratio
3. Calculating the mass or pressure of the products.

Moles are the central unit!



In addition we can convert between mass and volume of a pure substance by looking at it's **density**,

$$\boxed{\delta = m/V}$$

1.3.2

Stoichiometry of a Reaction

We use the **coefficients** of the **balanced reaction** along with the our equations that convert mass, volume, and concentration into moles to predict the quantities of reactants and products in a chemical reaction.

To answer any stoichiometry problem, focus on converting to and from moles!

General Steps to Solving a Stoichiometry Problem:

1. Convert the values given in the problem about a reactant or product to a **number of moles**
2. Use the **stoichiometric coefficients** from the **balanced** reaction to **find the number of moles of the unknown** you are being asked for
3. Convert the number of moles of your unknown to a mass, or whatever quantity you are being asked for

Example

2.6g of sodium metal (Na) reacts with water to form NaOH and H_2 according to the unbalanced reaction below. Once the reaction is complete how many grams of NaOH are formed?

$$Na_{(s)} + \;\; H_2O_{(g)} \rightarrow \;\; NaOH_{(s)} + \;\; H_{2(g)}$$

WIZE TIP

On an exam, your prof will NOT specify if an equation is balanced or unbalanced.
Always double check that the equation is balanced. If it's not, balance the equation before continuing!
The equation must be balanced in order to get the correct answer!

Step 1: Balance the reaction

$$2\ Na_{(s)} + 2\ H_2O_{(g)} \rightarrow 2\ NaOH_{(s)} + H_{2(g)}$$

Step 2: Find the number of moles of sodium

$$n_{Na} = \frac{m}{MW} = \frac{2.6\ g}{22.990\ g/mol} = 0.113\ mol$$

Step 3: Find the number of moles of NaOH

$$n_{NaOH} = n_{Na} \times \frac{coefficient\ NaOH}{coefficient\ Na} = 0.113\ mol \times \frac{2}{2} = 0.113\ mol$$

Step 4: Convert the number of moles of NaOH to a mass

$$n = \frac{m}{M}$$

$$m_{NaOH} = n \times MW = 0.113\ mol \times 39.997\ g/mol = 4.520\ g = 4.5\ g$$

1.3.3

Example: Converting Mass of Reactant to Mass of Product

Calculate the mass of hydrogen gas produced if 0.550 g of iron powder is reacted with an excess amount of sulfuric acid.

$$Fe\ (s)\ +\ H_2SO_4\ (aq)\ \rightarrow\ Fe_2(SO_4)_3\ (aq)\ +\ H_2\ (g)$$

Step 1: Balanced Equation

$$2Fe_{(s)} + 3H_2SO_{4(aq)} \longrightarrow Fe_2(SO_4)_{3\,(aq)} + 3H_{2\,(g)}$$

Step 2: Calculate the molar amount of $Fe_{(s)}$

$$n = \frac{m}{M} = \frac{0.550\ g}{55.84\ g/mol} = 0.00985\ moles$$

Step 3: Determine # of moles of $H_{2(g)}$

$$0.00985\ moles\ Fe \times \frac{3\ moles\ H_2}{2\ moles\ Fe} = 0.01477\ moles\ H_2$$

Step 4: Calculate mass of Hydrogen gas

$$m = nM = (0.01477\ moles)(2.02\ g/mol) = 0.0298\ g$$

Example: Solution Stoichiometry to Determine Mass of Product

Lithium metal was added to a 25mL of a 1.3M solution of Ag_2SO_4. The unbalanced chemical reaction is shown below.

$$Ag_2SO_{4(aq)} + Li_{(s)} \rightarrow Li_2SO_{4(aq)} + Ag_{(s)}$$

Once the reaction has gone to completion, what mass of silver metal is produced?

a) 0.0070 g
b) 7.0 g ✓
c) 3.5 g
d) 0.0035 g

Step 1: Balanced Reaction.

$$Ag_2SO_{4(aq)} + 2Li_{(s)} \longrightarrow Li_2SO_{4(aq)} + 2Ag_{(s)}$$

Step 2: Determine moles of Ag_2SO_4 in solution

$n = MV \quad V = 0.025L$

$$n_{Ag_2SO_4} = (1.3\,mol/L)(0.025\,L) = 0.0325\,mol$$

Step 3: Determine the moles Ag produced

$$n_{Ag} = n_{Ag_2SO_4} \times \frac{2}{1} = 0.065\,mol$$

Step 4: Determine mass of Ag

$$m = nW = (0.065\,mol)(107.87g/mol) = 7.01\,g$$

1.3.5

Practice: Calculate the Mass of Product

16.0 mL of Hg (mp = -38.8°C, density = 13.56 g/mL) reacts with fluorine gas and the reaction goes to completion as shown below.

$$Hg_{(l)} + F_{2(g)} \rightarrow HgF_{2(s)}$$

Once the reaction has gone to completion, what mass of mercury fluoride, HgF_2, is produced?

- A) 1.48 g
- B) 514 g
- C) 142 g
- D) 258 g ✓

Step 1: Calculate the mass of Hg

$\delta = 13.56\ g/mL \quad V = 16.0\ mL$

$m = V\delta = 16\ mL \times 13.56\ g/mL = 216.96\ g$

Step 2: Calculate the moles of Hg

$$n = \frac{m}{M} = \frac{216.96g}{200.59g/mol} \approx 1.082\ mol$$

Step 3: Determine the moles of HgF_2

$$Hg_{(l)} + F_{2(g)} \longrightarrow HgF_{2(g)}$$

$n_{HgF_2} = n_{Hg} = 1.082\ mol$

Step 4: Calculate the mass of HgF_2

$M_{Fe_2} = 18.998 \times 2 = 37.996\ g/mol \quad M_{Hg} = 200.59\ g/mol$

$M_{Fe_2Hg} = 238.586\ g/mol$

$m = nM = 1.082mol \times 238.586\ g/mol \approx 258.0g$

1.3.6

Practice: Stoichiometry Calculations Application

A solution that is 183mM (millimolar) NaCl(aq) is isoosmotic with plasma. This means that cells don't swell or shrink when in this solution. How many grams of sodium chloride are required to make 150mL of isoosmotic NaCl(aq)? The molar mass of NaCl is 58.44g/mol.

- 0.2g
- 1.6g ✓
- 2.2g
- 4.9g

Step 1: Calculate the moles of NaCl

$n = cV$ concerntration $c = 183\ mmol/L$ $V = 150\ mL$

$n = 183\ mmol/L \times 0.150\ L = 27.45\ mmol$

$n = 0.02745\ mol$

Step 2: Calculate the mass of NaCl

$m = nM = 0.02745\ mol \times 58.44\ g/mol$

$\approx 1.604\ g$

1.4 Limiting Reagents

1.4.1

Introduction to Limiting Reagents

Anytime **reactant species are in limited supply** and not present in perfectly proportional amounts, a chemical reaction will have a **limiting reagent** which will be totally **consumed before any other reactant**

The quantity of the limiting reagent available directly **determines the maximum number of product molecules** that can be formed!

Let's Consider an Example:

When making smores (yum!!) the "reaction" looks something like:

2 graham crackers + 1 piece of chocolate + 1 marshmellow → 1 smore

If I had 10 graham crackers, 6 pieces of chocolate, and 6 marshmellows, what would be the limiting reagent? In other words, what would I run out of first? And how many smores could I make?

- The "chemistry way" to figure this out would be to **take the # of moles of each reactant and divide by its stoichiometric coefficient:**

Graham crackers: 10/2=5
Pieces of chocolate: 6/1=6
Marshmellows: 6/1=6

- Now, **to figure out the limiting reagent, look at which of the above numbers are the smallest!!**
  - 5 is smallest, therefore Graham crackers are the limiting reagent!!

Now how many smores could we make?

We know graham crackers will determine how much product we get since we'll run out of the crackers first.

$$moles\ of\ graham\ crac\ker\ x\ \frac{1\ mol\ smores}{2\ moles\ graham\ crac\ker} = moles\ of\ smores$$

10 x (1/2)= 5 moles of smores created!

WIZE CONCEPT

It is necessary to **determine the limiting reagent** whenever we are **given the amounts of two or more reactants** in a chemical reaction.

The limiting reagent is **totally consumed in the reaction**. **Any other reactant is in excess.**

Limiting reagent stops the reaction by **running out first.**

The **quantity of the limiting reagent** available directly **determines the maximum number of products** molecules that can be formed

Example: Determine the Mass of Product in a Limiting Reagent Problem

Iron and chlorine gas react to form iron (III) trichloride. If 110 g of iron and 105 g of chlorine gas are reacted, which species is the limiting reagent? What is the maximum mass of $FeCl_3$ that can be formed?

$$2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s)$$

WIZE TIP

Steps for Solving Limiting Reagent Problems:

Step 1 – Write & **balance** the equation.

Step 2 – Calculate **moles of each reactant.**

Step 3 – Use the molar ratios of the present reactants (from the balanced equation) to determine the limiting reactant (LR).

- Take the **# of moles of each reactant** from step 2 and **divide by the stoichiometric coefficient** for that reactant
  - **Smallest value** from this calculation tells us the **LR**

Step 4 – From the limiting reactant, use the **molar ratio** (from the balanced equation) to **calculate moles of the desired product.**

Step 5 – Convert moles to the desired units (density, molarity, grams, etc...)

$$2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s)$$

Step 1 – Write & balance the equation.

Is the equation given balanced? Yes

Step 2 – Calculate moles of each reactant.

i) Find moles of Fe

$$n = \frac{m}{M} \quad m = 110g \quad M = 55.845 g/mol$$

$$n = \frac{110\ g}{55.845\ g/mol} = 1.9697\ mol$$

ii) Find moles of Cl_2

$$n = \frac{m}{M} \quad m = 105g \quad M = 70.906 g/mol$$

$$= \frac{105\ g}{70.906 g/mol} = 1.48083\ mol$$

Step 3 – Use the molar ratios of the present reactants (from the balanced equation) to determine the limiting reactant.

Fe mole consumed ratio : $1.9697/2 = 0.98495$

Cl_2 mole consumed ratio $1.48083/3 = 0.49361$

Therefore the limiting reagent is Cl_2 ! (because it had the smallest number)

Step 4 – From the limiting reactant, use the molar ratio (from the balanced equation) to calculated moles of the desired product.

- The maximum mass of $FeCl_3$ that is formed depends on how much Cl_2 we have (LR).

$n_{FeCl_3} = ?$ $\quad$ $n_{Cl_2} = 1.48083\ mol$

$$n_{Cl_2} \times \frac{2\ mol_{FeCl_3}}{3\ mol\ _{Cl_2}} = 0.98722\ mol$$

Step 5 – Convert moles to the desired units (density, molarity, grams, etc...)

$m_{FeCl_3} = ?$

$$m = nM = 0.98722\ mol \times 162.2\ g/mol$$

$$= 160.13\ g$$

1.4.3

Example: What is the Limiting Reagent?

2.0 g of aqueous Barium hydroxide ($Ba(OH)_2(aq)$, 171.32 g/mol) and 1.5 g of liquid hydrogen bromide ($HBr(l)$, 80.91g/mol) react together to give aqueous barium bromide ($BaBr_2(aq)$) and liquid water ($H_2O(l)$). Which species would be the limiting reagent of this reaction?

$$Ba(OH)_{2\,(aq)} + 2HBr_{(l)} \longrightarrow BaBr_{2\,(aq)} + 2H_2O_{(l)}$$

Step 1: Calculate moles of each reactant

$$n_{Ba(OH)_2} = \frac{2g}{171.32\,g/mol} = 0.011674\,mol$$

$$n_{HBr} = \frac{1.5g}{80.91\,g/mol} = 0.01854\,mol$$

Step 2: Divide the moles by Stoichiometric coefficient

$Ba(OH_2)$ Ratio: $0.011674/1 = 0.011674$ mole

HBr Ratio: $0.01845/2 = 0.00927$ mole

∴ HBr is the limiting reagent.

Example: How Many Grams of the Excess Reagent Will Remain?

The reaction between P_4 and Br_2 is very exothermic and produces PBr_5 as the only product. If 7.0 g of P_4 react with 12.0 g of Br_2 how many grams of the excess reagent will remain?

a) 0.4 g
b) 4.7 g
c) 6.1g
d)11.1 g

Balanced the Reaction $P_{4(s)} + 10 Br_{2(l)} \longrightarrow 4 PBr_{5(s)}$

$$P_{4(s)} + 10 Br_{2(l)} \longrightarrow 4 PBr_{5(s)}$$

m	7.0 g	12.0 g
MM	123.90 g/mol	159.81 g/mol
n	0.0565 mol	0.0751 mol

Determine the LR (limiting reagent)

$P_4 = 0.0565/1 = 0.0565$ mole

$Br_2 = 0.0751/10 = 0.00751$ mole

$$n_{P_4\,remained} = n_{P_4\,start} - n_{P_4\,consumed} = 0.0565\text{ mol} - 0.0751\text{ mol}$$
$$= 0.0490\text{ mol}$$

$$m_{P_4\,remained} = nM = 0.0490\text{ mol} \times 123.90\text{ g/mol} = 6.07\text{ g}$$

1.4.5

Practice: What is the Limiting Reagent

If 12 g of sodium is mixed with 0.64 g of H_2 and 17 g of O_2 which reagent will be limiting in the reaction shown below?

$$2\ Na_{(s)} + O_{2(g)} + H_{2(g)} \rightarrow 2\ NaOH$$

- A) Na ✓
- B) O_2
- C) H_2
- D) None, they will all be completely consumed

Step 1: Calculate the mole amount of Na, H_2, O_2

$n_{Na} = ?$ $m_{Na} = 12g$ $M_{Na} = 22.99 g/mol$

$n_{Na} = \frac{12}{22.99} \approx 0.522 mol$

$n_{H_2} = ?$ $m_{H_2} = 0.64g$ $M = 2.016 g/mol$

$n_{H_2} = \frac{0.64}{2.016} \approx 0.317 mol$

$n_{O_2} = ?$ $m_{O_2} = 17g$ $M = 32.0 g/mol$

$n_{O_2} = \frac{17}{32.0} \approx 0.531 mol$

Step 2: After divide by Stoichiometric coefficient

Na Ratio: 0.261 H_2 Ratio: 0.531 O_2 Ratio: 0.317

∴ The Na is the LR (limiting reagent)

1.5 Percent Yield

1.5.1

Percent Yield

So far we have been assuming that the reaction proceeds 100% to products. In reality, this is rarely the case!

When we are doing experiments in the lab we often lose some of our product by spilling it or residues of it get left on glassware and spatulas.

So far we have been calculating the **theoretical yield.**

Theoretical Yield - the maximum amount of product produced based on the quantity of the limiting reagent (*e.g.* the amount of product produced in a calculation).

Actual Yield - the actual amount of product produced in the reaction (e.g. the amount of product obtained in a laboratory experiment).

$$\boxed{\%\ yield = \frac{actual\ yield}{theoretical\ yield}\ x100}$$

Test Your Understanding

1) In a lab, you measure your product and record 5.222g. This is the actual yield.

2) If everything went 100% perfectly according to the reaction we would get the theoretical yield

3) If I determined the limiting reagent, and then calculated the number of moles of product and the mass of product based on this, I've just calculated the theoretical yield

Example: What is the Percent Yield

When 49.00 g of a hydrocarbon fuel with formula $C_7H_{10}O_2$ is reacted with excess oxygen, a total of 21.56 g of water is collected. What was the percent yield of the reaction?

$$\underline{\ 2\ }\ C_7H_{10}O_2 + \underline{\ 17\ }\ O_2 \rightarrow \underline{\ 10\ }\ H_2O + \underline{\ 14\ }\ CO_2$$

actual yield $= 21.56g_{H_2O} = m_{H_2O}$

$m = 49\,g_{C_7H_{10}O_2}$ $\quad M_{C_7H_{10}O_2} = 126\,g/mol$

$n_{C_7H_{10}O_2} = \frac{49}{126} = 0.38889$ mol

convert to $n_{H_2O} = \frac{10\ mol_{H_2O}}{2\ mol_{C_7H_{10}O_2}} = 1.945$ mol

$M_{H_2O} = 18\,g/mol$

$m_{H_2O} = nM = 1.945(18) = 35.01\,g_{H_2O}$

% yield $= \frac{\text{actual yield}}{\text{theoritical yield}} = \frac{21.56}{35.01} = 62\%$

1.5.3

Practice: Calculate Moles Given Percent Yield

If the yield for the following reaction is 45%, how many moles of $KClO_3$ are needed to produce 1 mol of O_2?

$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

- A) 0.4 moles
- B) 1.1 moles
- C) 1.5 moles ✓
- D) 4.6 moles

$$n_{KClO_3} = \frac{2\,mol\ KClO_3}{3\,mol\ O_2} \times n_{O_2} = \frac{2}{3} = 0.667\,mol$$

$$n_{require} = \frac{n_{stoich}}{yield} = \frac{0.667}{0.45} \approx 1.482\,mol$$

Practice: Limiting Reagent & Percent Yield

Iron oxidizes when exposed to oxygen to form iron oxide according to the following equation:

$$4\,Fe + 3O_2 \rightarrow 2Fe_2O_3$$

352g of pure iron is exposed to 12.0 mols of O_2, and after a period of time 46.7 g of iron oxide (rust) is collected.

Part 1

What is the limiting reagent in this reaction?

- A) Fe ✓
- B) O_2
- C) Fe_2O_3
- D) There is no limiting reagent

$$4Fe + 3O_2 \longrightarrow 2Fe_2O_3$$

4 moles of Fe react with 3 moles of O_2

$$n_{Fe} = \frac{352\,g}{55.85\,g/mol} \approx 6.30\,mol_{Fe}$$

$$n_{O_2} = 12.0\,mol_{O_2}$$

$$n_{O_2\,required} = \frac{6.3 \times 3}{4} = 4.73\,mol$$

$\therefore$ Fe is the limiting reagent.

Practice: Limiting Reagent & Percent Yield

Iron oxidizes when exposed to oxygen to form iron oxide according to the following equation:

$$Fe + O_2 \rightarrow Fe_2O_3$$

352g of pure iron is exposed to 12.0 mols of O_2, and after a period of time 46.7 g of iron oxide (rust) is collected.

Part 2

What is the percent yield?

A) 2.2%

B) 8.8%

C) 9.4% ✓

D) 33.1 %

$$n_{Fe_2O_3 \text{ produced}} = \frac{6.3 \times 2}{4} = 3.15 \text{ mol } Fe_2O_3$$

$$M_{Fe_2O_3} = 159.7 \text{ g/mol}$$

$$m_{Fe_2O_3} = 3.15 \text{ mol} \times 159.7 \text{ g/mol} \approx 503.1 \text{ g}$$

$$\% \text{ yield} = \frac{46.7}{503.1} \times 100\% = 9.28\%$$

Practice: Limiting Reagent & Percent Yield

Iron oxidizes when exposed to oxygen to form iron oxide according to the following equation:

$$Fe + O_2 \rightarrow Fe_2O_3$$

352g of pure iron is exposed to 12.0 mols of O_2, and after a period of time 46.7 g of iron oxide (rust) is collected.

Part 3

How many Fe atoms are present in the rust produced?

- A) 0.89 x 10^{17} atoms
- B) 1.77 x 10^{23} atoms
- C) 3.56 x 10^{23} atoms ✓
- D) 4.33 x 10^{23} atoms

$$n_{Fe_2O_3} = \frac{46.7g}{159.7g/mol} \approx 0.292 \, mol \, Fe_2O_3$$

$$n_{Fe} = \frac{2 \, mol \, Fe}{1 \, mol \, F_2O_3} = 0.584 \, mol$$

$$\# \, Fe \, atom = 0.584 \, mol_{Fe} \times 6.02 \times 10^{23} \, atoms/mol$$

$$\approx 3.52 \times 10^{23} \, atoms$$