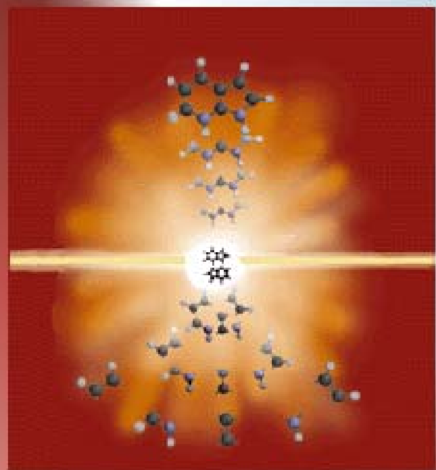
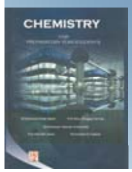
# Chapter 2
# Stoichiometry





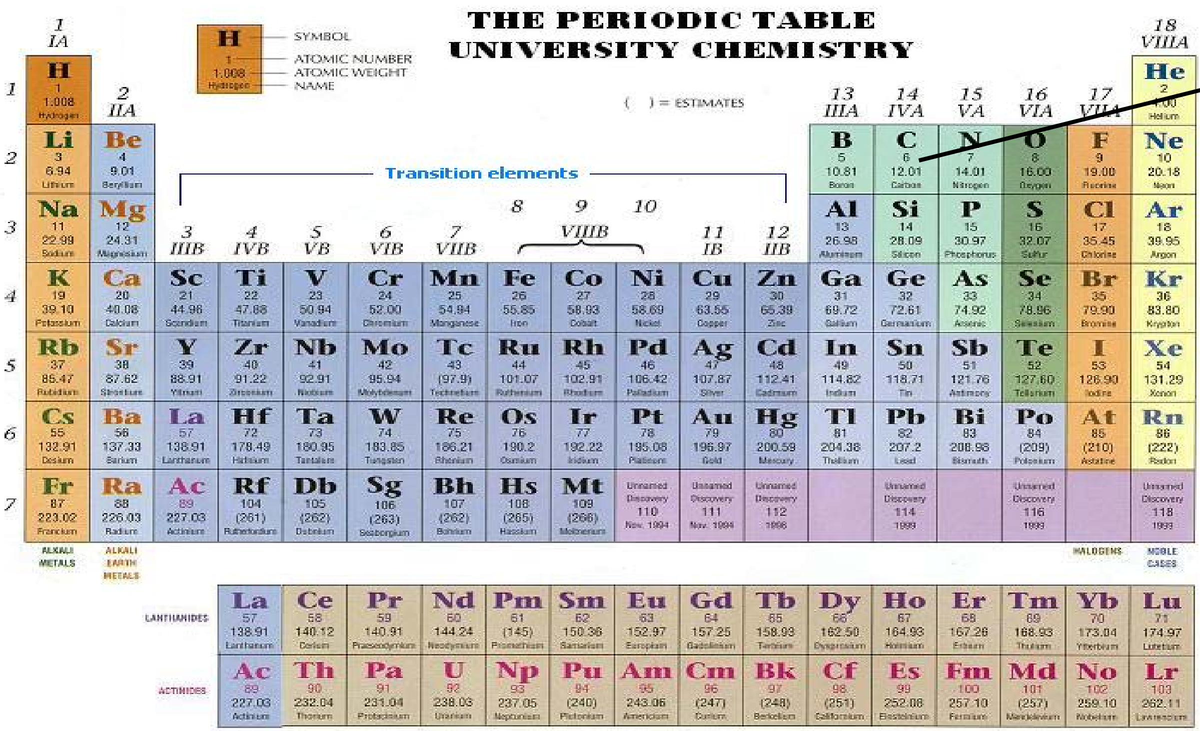
## Atomic Mass:

The **atomic mass** of an element is the mass average of the atomic masses of the different isotopes of an element.

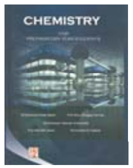
For example, naturally occurring carbon, for example, is a mixture of two isotopes, $^{12}C$ (98.89%) and $^{13}C$ (1.11 %).

Individual carbon atoms therefore have a mass of either 12.000 or 13.03354 amu. But the average mass of the different isotopes of carbon is 12.011 amu.

$$\frac{98.89}{100} \text{ X } 12.00 \text{ amu} + \frac{1.11}{100} \text{ X } 13.003354 \text{ amu} = 12.011 \text{ amu}$$




C
6
12.01
Carbon

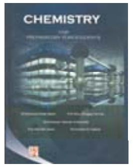


# Molecular Mass (Molar Mass, Molecular weight)

- The **molecular mass** of a compound is the sum of the atomic masses of the atoms in the molecules that form these compounds.

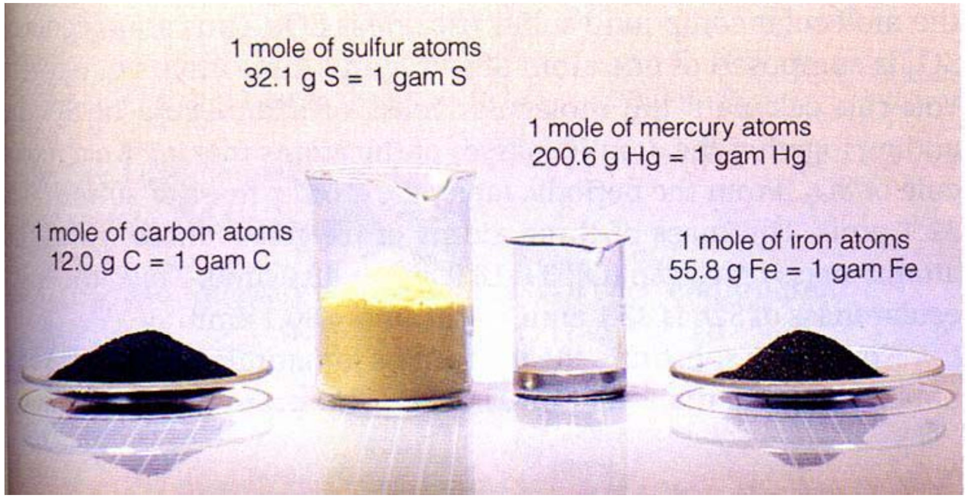
- **Calculate the molecular mass of the sugar molecule found in cane sugar ($C_{12}H_{22}O_{11}$).**

| Atom | Symbol | Number of atoms | Mass of one atom | Total mass (amu) | |
|---|---|---|---|---|---|
| **Carbon** | **C** | 12 | 12.011 amu | 12 x 12.011 | 144.132 |
| **Hydrogen** | **H** | 22 | 1.0079 amu | 22 x (1.0079) | 22.174 |
| **Oxygen** | **O** | 11 | 15.9994 amu | 11 x (15.9994) | 175.993 |
| | | | | | **342.299** |



# Molar mass and the mole

- one mole is defined as the number of carbon atoms in exactly 12.000000 grams of pure $^{12}C$.
- A mole of sugar ($C_{12}H_{22}O_{11}$) would have a mass of 342.299 grams.
- This quantity is known as the **molar mass**, a term that is often used in place of the terms *atomic mass* or *molecular mass.*




**Determine the molar mass of NaOH?**

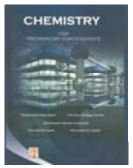
NaOH contains one Na atom + one oxygen atom + one hydrogen atom

Molar mass = 1 x mass of Na atom + 1 x mass of O atom + 1 x mass of H atom

**The masses of the elements can be obtained from the periodic table.**

$$= 1 \times 22.99 + 1 \times 16.00 + 1 \times 1.008 = 39.99\ \text{g}$$

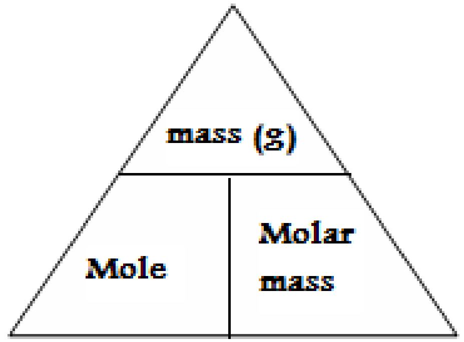
**Molar mass of NaOH = 39.99 g**



# Number of moles

- **To determine the number of moles use the following formula or triangles:**

$$number\ of\ moles = \frac{mass\ (g)}{molar\ mass\ (g\ /\ mole)}$$




**How many moles are there in 22.99 g of sodium?**

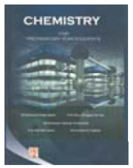
$$\text{number of moles} = \frac{\text{mass (g)}}{\text{molar mass (g / mole)}} = \frac{22.99\,\text{g}}{22.99\,\text{g / mole (from the periodic table)}}$$

**number of moles = 1 mole.**

**How many moles are there in 1 g of chlorine?**

$$\text{number of moles} = \frac{\text{mass (g)}}{\text{molar mass (g / mole)}} = \frac{1\,\text{g}}{35.45\,\text{g / mole (from the periodic table)}}$$

**number of moles = 0.028 mole.**



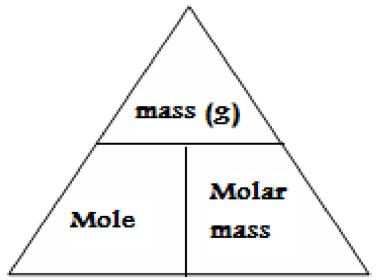
## How many grams are there in 0.10 mole of $CH_4$?

**First calculate the molar mass of $CH_4$**

**Molar mass of $CH_4$ = 1 x mass of C atom + 4 x mass of H atoms**

**= 1 x 12.01 + 4 x 1.008 = 16.02 g / mole**

**Then use the formula:**




**mass of $CH_4$ = number of moles × molar mass of $CH_4$**

**= 0.10 mole x16.02 g/ mole = 1.602 g**

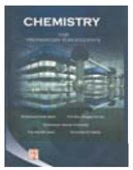
## Which one is the lightest in mass: one mole of hydrogen, one mole of sodium, one mole of iron, one mole of sulfur?

One mole for an element contains the atomic mass of the element.

Atomic mass of H = 1.008 g / mole, Atomic mass of Na = 22.99 g / mole,
Atomic mass of Fe = 55.85 g / mole, Atomic mass of S = 32.07 g / mole.

**The lightest one is one mole of hydrogen**

**THE HEAVIEST ONE MOLE IS THE IRON**



# AVOGADRO'S NUMBER & THE MOLE

**1 mole of anything contains the Avogadro 's Number ($N_A$) of this thing**

**Avogadro 's Number (NA) = *6.02214 x $10^{23}$***

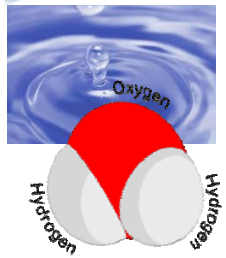
***1 mole of particles= 6.02214 x $10^{23}$ particles for any substance***

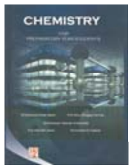
***1 mole of shoes= 6.02214 x $10^{23}$ shoes***

***1 mole of cars = 6.02214 x $10^{23}$ car***

***1 mole of carbon atoms= 6.02214 x $10^{23}$ carbon atoms***

***1 mole of water molecules = 6.02214 x $10^{23}$ water molecules***

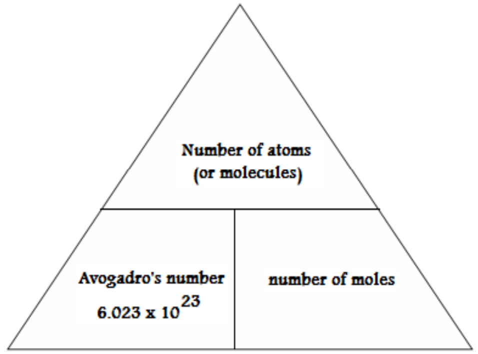



**Number of particles = number of moles x Avogadro's number**

**To calculate the number of particles (atoms, molecules, shoes….etc) use the following formula:**

**Number of particles = number of moles x Avogadro's number**




**Calculate the number of atoms in 2 mole of hydrogen?**

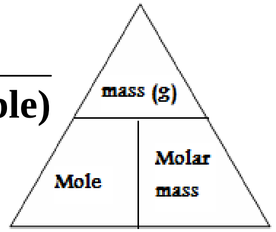
Number of hydrogen atoms =

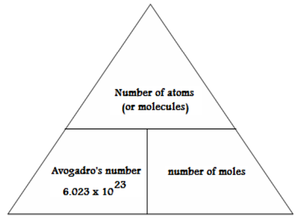
2 moles of H x 6.02214 x $10^{23}$ H atom / mole

Number of hydrogen atoms = 1.20 x $10^{24}$ H atom

**Calculate the number of atoms in 6.46 grams of helium (He)?**

$$\text{number of moles} = \frac{\text{mass (g)}}{\text{molar mass (g / mole)}} = \frac{6.46\,\text{g}}{4.003\,\text{g / mole (from the periodic table)}}$$



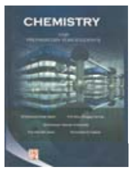




number of moles = 1.61 mole.

Number of He atoms = number of moles × Avogadro's number

= 1.61 moles of He x 6.02214 x $10^{23}$ He atom / mole

= 9.66 x $10^{23}$ He atom



- **This could be done in one step:**

$$\text{number of atoms} = \frac{\text{mass (g)}}{\text{molar mass (g/mole)}} \text{ x } N_A$$

$$\textbf{number of } \boldsymbol{He} \textbf{ atoms} = \frac{\textbf{6.46 (g)}}{\textbf{4.0 (g/mole)}} \textbf{ x 6.02214 x } \mathbf{10^{23}} \textbf{ He atom / mole}$$

**= 9.66 x $10^{23}$ He atom**

## Calculate the mass of one atom of sodium?

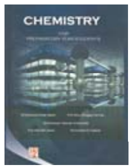
**Use the following formula:**

$$\text{number of atoms} = \frac{\text{mass (g)}}{\text{molar mass (g/mole)}} \text{ x } N_A$$

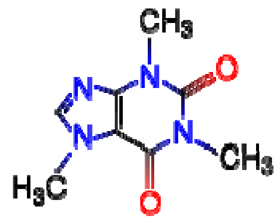
**Or**

$$\text{mass of 1 atom} = \frac{\text{number of atoms x molar mass (g/mole)}}{N_A}$$

$$\text{mass of 1 atom Na} = \frac{\text{1 atom x 23.0 (g/mole)}}{6.022 \text{ x } 10^{23} \text{ (atom/mole)}} = 3.82 \text{ x } 10^{-23} \text{ g}$$



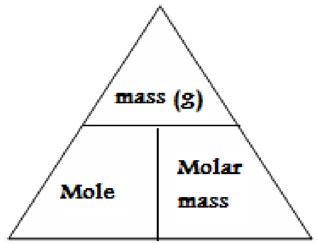
**Caffeine is a stimulant drug and it is found in coffee, tea and beans. Its molecular formula is $C_8H_{10}N_4O_2$. Calculate the number of oxygen atoms in 19.40 grams of caffeine.**




**Molar mass of caffeine = 8 x C + 10 x H + 4 x N + 2 x O**

**= 8 x 12 + 10 x 1 + 4 x 14 + 2 x 16 = 194 g / mole**

$$\text{number of moles} = \frac{\text{mass (g)}}{\text{molar mass (g / mole)}} = \frac{19.40\ \text{g}}{194\ \text{g / mole (from the periodic table)}}$$

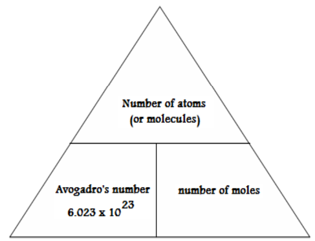



**number of moles = 0.10 mole**

**Total number of $C_8H_{10}N_4O_2$ molecules= number of moles x $N_A$**

**= 0.10 moles x 6.022 x $10^{23}$ molecules / mole**

**Total number of $C_8H_{10}N_4O_2$ molecules = 6.022 x $10^{22}$ molecules**




$$\text{Number of oxygen atoms} = \frac{\text{number of oxygen atoms}}{\text{molecules}} \text{ x total number of molecules}$$

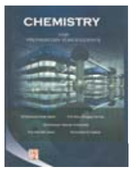
$$\text{Number of oxygen atoms} = \frac{\text{2 oxygen atoms}}{\text{molecules}} \text{ x } 6.022 \text{ x } 10^{22} \text{ molecules}$$

**Number of oxygen atoms = 1.20 x $10^{23}$ oxygen atoms**

Number of carbon atoms = 4.8 x $10^{23}$ carbon atoms

Number of hydrogen atoms = 6.022 x $10^{23}$ hydrogen atoms

Number of nitrogen atoms = 2.40 x $10^{23}$ nitrogen atoms



**Caffeine is a stimulant drug and it is found in coffee, tea and beans. Its molecular formula is $C_8H_{10}N_4O_2$.**
**Calculate the number of oxygen atoms in 19.40 grams of caffeine.**

## This could be done in one step:

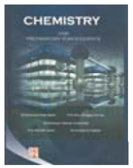
$$\text{number of atoms} = \frac{\text{mass (g)}}{\text{molar mass (g/mole)}} \text{ x } N_A \text{ } x \text{ } \textit{number of atoms per molecule}$$

$$\textit{number of oxygen atoms} = \frac{19.4(g)}{194(g/mole)} x\, 6.022\, x\, 10^{23}\, x\, 2 = 1.2\, x 10^{23}\ \textit{oxygen atoms}$$

$$\textit{number of carbon atoms} = \frac{19.4(g)}{194(g/mole)} x\, 6.022\, x\, 10^{23}\, x\, 8 = 4.8\, x 10^{23}\ \textit{carbon atoms}$$

$$\textit{number of hydrogen atoms} = \frac{19.4(g)}{194(g/mole)} x\, 6.022\, x\, 10^{23}\, x\, 10 = 6.022\, x 10^{23}\ \textit{hydrogen atoms}$$

$$\textit{number of nitrogen atoms} = \frac{19.4(g)}{194(g/mole)} x\, 6.022\, x 10^{23}\, x\, 4 = 2.4\, x 10^{23}\ \textit{nitrogen atoms}$$



# Mass Percent

The Mass Percent of an element is defined as:

$$\text{Mass Percent of an element} = \frac{\text{Mass of the element}}{\text{Total molar mass of the sample}} \times 100\%$$

**What is the mass percent of carbon, hydrogen, and oxygen in pure ethanol $C_2H_6O$?**

**-First: calculate the molar mass of $C_2H_6O$**

MW of $C_2H_6O$ = 2 x C + 6 x H + 1 x O

= 2 x 12.01 + 6 x 1.008 + 1 x 16.00

MW $C_2H_6O$ = 46.07 g/mole

**-Second: calculate the mass percents**

$$\text{Mass \% C} = 100 \times \left( \frac{\text{mass of C}}{\text{total molar mass}} \right) = 100 \times \left( \frac{2 \times 12.01}{46.07} \right) = 52.14\ \%$$

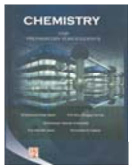
$$\text{Mass \% H} = 100 \times \left( \frac{\text{mass of H}}{\text{total molar mass}} \right) = 100 \times \left( \frac{6 \times 1.008}{46.07} \right) = 13.13\ \%$$

$$\text{Mass \% O} = 100 \times \left( \frac{\text{mass of O}}{\text{total molar mass}} \right) = 100 \times \left( \frac{1 \times 16.00}{46.07} \right) = 34.72\ \%$$

Note that the mass percentages should add up to 100%.

Mass % = Mass % C + Mass % H + Mass % O

= 52.14 % + 13.13 % + 34.72 % = 99.99 %



- **Ascorbic acid (vitamin C) contains only C, H, and O. Combustion of 1.000 g of Ascorbic acid produced 40.9% C and 4.5% H. What is the empirical formula for Ascorbic Acid?**

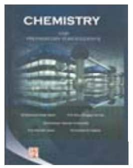
**First: calculate the mass percent of Oxygen.**

**Since the sample contains C, H, and O, then the remaining**

**100% - 40.9% - 4.5% = 54.6% is Oxygen**

**Second: Suppose 100 g of this substance**

| Steps | | C | H | O |
|---|---|---|---|---|
| 1 | Mass /g | 40.9 | 4.5 | 54.6 |
| 2 | No. of moles = $\frac{\text{mass}}{\text{molar mass}}$ | $\frac{40.9}{12}$ =3.4 | $\frac{4.5}{1}$ =4.5 | $\frac{54.6}{16}$ =3.4 |
| 3 | ÷ smallest number (3.4) | 1 | 1.3 | 1 |
| 4 | x by a number to make step 3 integer numbers (x 3) | 1 x 3 = 3 | 1.3 x 3 = 4 | 1 x 3 = 3 |
| 5 | Empirical formula $C_3H_4O_3$ | 3 C | 4 H | 3 O |



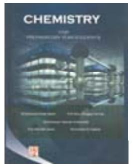
Empirical formula $C_3H_4O_3$

## What is the molecular formula if the molecular mass of Ascorbic Acid was founded to be 176 g/mole?

$$\text{Molecular Formula} = \left(\frac{\text{Molecular weight of unknown (g/mole)}}{\text{mass of emperical formula}}\right) \text{x empirical Formula}$$

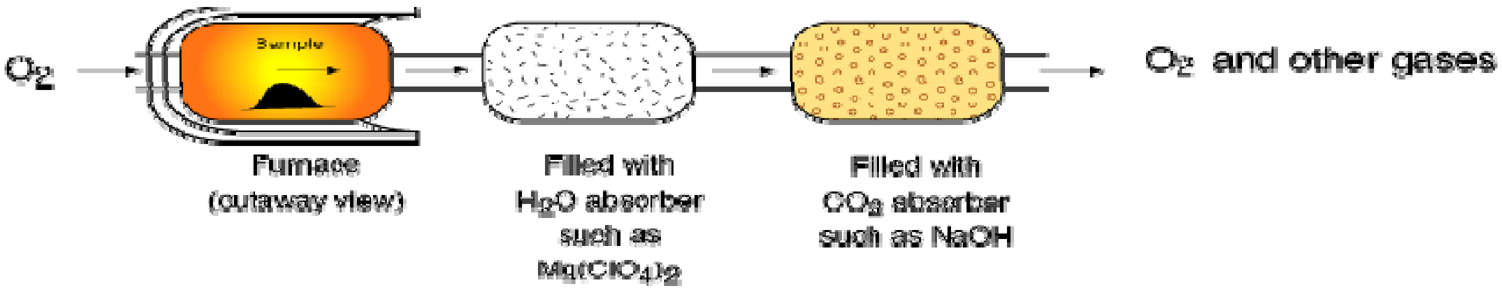
$$\text{Molecular Formula} = \left(\frac{176\ \text{(g/mole)}}{3\text{x}12 + 4\text{x}1 + 3\text{x}16}\right) \text{x } C_3H_4O_3 =$$

$$= 2 \text{ x } C_3H_4O_3 = C_6H_8O_6$$



# Combustion Analysis

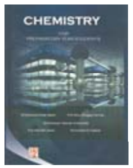
- It is used to determine the mass % for different elements in the compound.




The sample is burned in the presence of excess oxygen which converts all the carbon to carbon dioxide and all the hydrogen to water.

The $CO_2$ and $H_2O$ produced are absorbed in two different stages and their masses determined by measuring the increase in weight of the absorbers.

$$\text{Mass of an element in a sample} = \frac{\text{Mass of the product}}{\text{Its molar mass}} \times \text{number of atoms} \times \text{atomic mass of the element}$$



**Ethanol, contains carbon, hydrogen and oxygen. Upon the combustion of a 1.15 g sample of ethanol gives 2.20 g of $CO_2$ and 1.35 g $H_2O$.**

**Calculate the mass percent for carbon, Oxygen, and hydrogen.**

First: determine the mass of the element (Hydrogen) in naphthalene.

$$\text{Mass of H in a sample} = \frac{\text{mass of } H_2O \text{ produced}}{\text{Molar mass of } H_2O} \times \text{number of H atoms} \times \text{atomic mass of H}$$

$$\text{Mass of Hydrogen in the sample} = \frac{1.35 \text{ g}}{18.0 \text{ g/mole}} \times 2 \text{ H atoms} \times 1.0 \text{ g/mole}$$

Mass of Hydrogen in the sample = 0.15 g

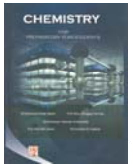
$$\text{Mass of C in a sample} = \frac{\text{mass of } CO_2 \text{ produced}}{\text{Molar mass of } CO_2} \times \text{number of C atoms} \times \text{atomic mass of C}$$

$$\text{Mass of Carbon in the sample} = \frac{2.20 \text{ g}}{44.0 \text{ g/mole}} \times 1 \text{ C atoms} \times 12.0 \text{ g/mole}$$

Mass of Carbon in the sample = 0.60 g

Mass of Oxygen in the sample = mass of the sample – mass of H – mass of C

= 1.15 – 0.15 – 0.60 = 0.40 g of oxygen



## Determine the % of each element:

$$\text{Mass \% of Hydrogen} = 100\text{ x } \frac{\text{mass of hydrogen}}{\text{Total sample mass}} = 100\text{ x } \frac{0.15\text{ g H}}{1.15\text{ g ethanol}} = 13.0\ \%$$

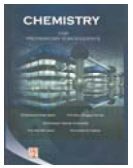
$$\text{Mass \% of Carbon} = 100\text{ x } \frac{\text{mass of Carbon}}{\text{Total sample mass}} = 100\text{ x } \frac{0.60\text{ g C}}{1.15\text{ g ethanol}} = 52.2\ \%$$

$$\text{Mass \% of Oxygen} = 100\text{ x } \frac{\text{mass of oxygen}}{\text{Total sample mass}} = 100\text{ x } \frac{0.40\text{ g O}}{1.15\text{ g ethanol}} = 34.8\ \%$$

Check: total mass % = mass % of H + mass % of C + mass % of O = 100 %

## Second: to determine the Empirical formula for ethanol:

| Steps | | C | H | O |
|---|---|---|---|---|
| 1 | Mass /g | 0.60 | 0.15 | 0.40 |
| 2 | No. of moles = $\frac{\text{mass}}{\text{molar mass}}$ | $\frac{0.60}{12}$ =0.05 | $\frac{0.15}{1}$ =0.15 | $\frac{0.40}{16}$ =0.025 |
| 3 | ÷smallest number (0.025) | 2 | 6 | 1 |
| 4 | Empirical formula $C_2H_6O$ | 2 C | 6 H | 1O |



**Empirical formula $C_2H_6O$**

# Chemical Reactions

It is process in which one or more pure substances are converted into one or more different pure substance.

All chemical reactions involve a change in substances and a change in energy.

Neither matter nor energy is created or destroyed in a chemical reaction, only changed.

## Chemical equation

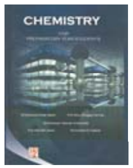
- When a chemical reaction occurs, it can be described by an equation.
- This shows the chemicals that react (***reactants***) on the left-hand side, and the chemicals that they produce (***products***) on the right-hand side.

$$\text{Reactants} \xrightarrow{\text{Reaction conditions}} \text{Products}$$

Reaction between hydrogen gas and oxygen gas to produce liquid water

$$\text{hydrogen gas} + \text{oxygen gas} \longrightarrow \text{liquid water}$$

$$2H_2(g) \quad + \quad O_2(g) \longrightarrow 2H_2O(l)$$



## Balancing chemical equations

- first write the correct formula for both reactants and products and then **balance** all of the atoms on the left side of the reaction with the atoms on the right side.

**Write the chemical equation which represents the burning of glucose in presence of oxygen gas which produces carbon dioxide and water.**

**To answer this question, follow the following steps:**

**1. *Identify the reactants and the products and put an arrow in between.***

glucose + oxygen gas $\longrightarrow$ carbon dioxide + water

***2. Try to figure out the correct formula for the reactants and products,***

Glucose is $C_6H_{12}O_6$, oxygen gas is $O_2$, carbon dioxide is $CO_2$, and water is $H_2O$.

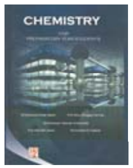
$$C_6H_{12}O_6 + O_2 \longrightarrow CO_2 + H_2O$$

***3.Count the number of each atom at both sides of the equation:***

$$C_6H_{12}O_6 + O_2 \longrightarrow CO_2 + H_2O$$

(6 C + 12 H + 6 O) + (2 O) $\longrightarrow$ (1C + 2 O) + (2H + 1 O)

Total: (6 C + 12 H + 8 O) $\longrightarrow$ (1C + 2H + 3 O )



***Balance C first, then H, and finally O:***

At the left side there are 6 C atoms and at the right side there are 1 C atom, so multiply $CO_2$ by 6 (x 6)

$$C_6H_{12}O_6 + O_2 \longrightarrow \mathbf{6}\ CO_2 + H_2O$$

At the left side there are 12 H atoms and at the right side there are 2 H atom, So multiply $H_2O$ by 6 (x 6)

$$C_6H_{12}O_6 + O_2 \longrightarrow \mathbf{6}\ CO_2 + \mathbf{6}H_2O$$

At the left side there are 8 O atoms and at the right side there are 18 O atom, So multiply $O_2$ by 6 (x 6)

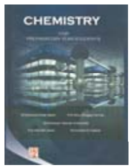
$$C_6H_{12}O_6 + \mathbf{6}O_2 \longrightarrow \mathbf{6}\ CO_2 + \mathbf{6}H_2O$$

*Recount all atoms again,*

$$C_6H_{12}O_6 + \mathbf{6}O_2 \longrightarrow \mathbf{6}\ CO_2 + \mathbf{6}H_2O$$

(6 C + 12 H + 6 O) + (12 O) (6C + 12 O) + (12H + 6 O)

Total: (6 C + 12 H + 18 O) (6 C + 12 H + 18 O)



# AMOUNT OF REACTANTS & PRODUCTS PROBLEMS

$$aA \longrightarrow bB$$

**Type1: calculate the number of moles of unknown (B) and number of moles of A is given:**

Use the following formula to calculate the # moles of B:

$$\textbf{number of moles of (B)} = \textbf{number of moles of (A)x}\left(\frac{\mathbf{b}}{\mathbf{a}}\right)$$

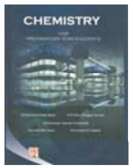
Calculate the number of moles of $CO_2$ resulted from the reaction of 3.5 moles of $C_2H_6$ with excess oxygen according to the equation

$$2\ C_2H_6 + 7\ O_2 \rightarrow 4\ CO_2 + 6\ H_2O$$

- **Use the formula:**

$$\textbf{number of moles of } (CO_2) = \textbf{number of moles of } (C_2H_6)\textbf{x}\left(\frac{4(C_2H_6)}{2(CO_2)}\right)$$

$$\textbf{number of moles of } (CO_2) = \textbf{3.5 moles of } (C_2H_6)\textbf{x}\left(\frac{4(C_2H_6)}{2(CO_2)}\right)$$




Number of moles of $CO_2$ = 7.0 moles

# AMOUNT OF REACTANTS & PRODUCTS PROBLEMS

$$aA \longrightarrow bB$$

Type 2: **calculate the mass of unknown (B) and mass of A is given:**

Use the following formula to calculate the mass of B:

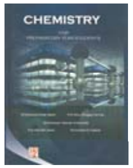
$$\text{mass of (B)} = \left(\frac{\text{mass of (A)}}{\text{Molar mass of (A)}}\right) \text{x} \left(\frac{b}{a}\right) \text{x Molar mass of (B)}$$

How many grams of water are produced when 7.00 grams of oxygen react with an excess of hydrogen according to the reaction shown below?

**$2H_2(g) + O_2(g) \text{----}> 2H_2O(g)$**

✓The "excess" reactant has nothing to do with the problem.
✓Identify which is the "given" and which is the unknown.

$$\begin{array}{ccc} 2H_2(g) + O_2(g) & \text{----}> & 2H_2O(g) \\ \quad\quad\quad 10\ g & & ? \end{array}$$



- **Use the formula:**

$$\text{mass of }(H_2O) = \left(\frac{\text{mass of } O_2}{\text{Molar mass of } O_2}\right) \text{x} \left(\frac{2(H_2O)}{1(O_2)}\right) \text{x Molar mass of }(H_2O)$$

$$\text{mass of }(H_2O) = \left(\frac{7.0\,g}{32\,g/\text{mole}}\right) \text{x} \left(\frac{2(H_2O)}{1(O_2)}\right) \text{x } 18\,g/\text{mole}$$

Mass of $H_2O$ = 7.89 g

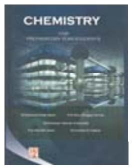
Calculate the mass of chlorine that reacts with 4.770 g of hydrogen to form hydrogen chloride according the following equation:

$$H_2 + Cl_2 \rightarrow 2\,HCl$$

- **Use the formula:**

$$\text{mass of }(Cl_2) = \left(\frac{\text{mass of } H_2}{\text{Molar mass of } H_2}\right) \text{x} \left(\frac{1(H_2)}{1(Cl_2)}\right) \text{x Molar mass of }(Cl_2)$$

$$\text{mass of }(Cl_2) = \left(\frac{4.770\,g \text{ of } H_2}{2.0\,g/\text{mole}}\right) \text{x} \left(\frac{1(H_2)}{1(Cl_2)}\right) \text{x } 71.0\,g/\text{mole}$$




Mass of $Cl_2$ = 169.3 g

# The limiting Reactant

$$aA + bB \longrightarrow dD$$

**When two substances A and B are present in random quantities and react with each other to produce D, the <u>first consumed</u> one is *the limiting reagent* and the <u>second one</u> is remained in *excess*.**

**The amount of product should be calculated from the amount of the limiting reactant**

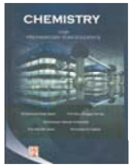
**It is easy to identify this type of problem, because the mass of two reactants are given.**

To determine the limiting reactant from given moles of substance, do the followings:

1- Calculate the number of moles for each reactant.

2- Determine the ratio, by divide the number of moles for each reactant by its coefficient.

3- The reactant with the *smallest* ratio is the *limiting reactant*.



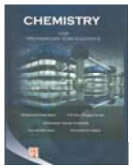
If 2.0 moles of NO were mixed with 2.0 moles of $O_2$ to react as:

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

Determine the limiting reactant.

To determine the limiting reactant, use the following table:

| **Steps** | | **2NO** | **1 $O_2$** |
|---|---|---|---|
| **1** | No. of moles | 2.0 | 2.0 |
| **2** | Coefficient | **2** | **1** |
| **3** | Ratio | $\frac{2.0}{2} = 1.0$ | $\frac{2.0}{1} = 2.0$ |
| **4** | Look for the smallest no. | Smallest | largest |



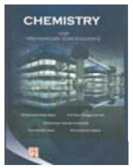
## If 400.0 g Fe were mixed with 300.0 g $O_2$ to react as:

$$4\,Fe(s) + 3\,O_2\,(g) \rightarrow 2\,Fe_2O_3\,(s)$$

## Determine the limiting reactant?

To determine the limiting reactant, use the following table:

| Steps | | 4 Fe | 3 $O_2$ |
|---|---|---|---|
| 1 | Mass /g | 400.0 | 300.0 |
| 2 | No. of moles = $\frac{\text{mass}}{\text{molar mass}}$ | $\frac{400.0}{56.0}$ =7.1 | $\frac{300.0}{32.0}$ =9.4 |
| 3 | Coefficient | 4 | 3 |
| 4 | Ratio | $\frac{7.1}{4}$ = 1.8 | $\frac{9.4}{3}$ = 3.1 |
| 5 | Look for the smallest no. | Smallest | largest |



# Chemical Reaction Yield

- For any chemical reaction there are theoretical and actual (practical) yield.

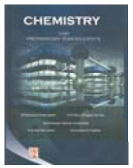
- *Theoretical yield (T.Y.)* is the amount of product that would result if all the limiting reactant reacted.

- *Actual yield (A.Y.)* is the amount of product actually obtained from a reaction.

- Due to many factors can affected on the reaction, A.Y. is always less than T.Y.

- *Percent yield* is the efficient for a given reaction:

$$\%\ \text{yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100$$




The actual yield most of the time, is the number before any of the following words: collected, obtained, isolated, produced, separated, formed…..etc.

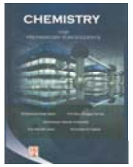
How many grams of $SF_4$ (g) can theoretically be prepared from 6.0 g of $SCl_2$ (g) and 35.0 g of NaF(s)?

$$3\ SCl_2\ (g) + 4\ NaF\ (s) \rightarrow SF_4\ (g) + S_2Cl_2\ (l) + 4\ NaCl\ (s)$$

**- Step 1: determine the limiting reactant**

| Steps | | 3 $SCl_2$ | 4 NaF |
|---|---|---|---|
| **1** | Mass /g | 6.0 | 35.0 |
| **2** | Molar mass (g/mole) | 103.0 | 42.0 |
| **3** | $\text{No. of moles} = \dfrac{\text{mass}}{\text{molar mass}}$ | $\dfrac{6.0}{103.0}$=0.06 | $\dfrac{35.0}{42.0}$=0.83 |
| **4** | Coefficient | **3** | **4** |
| **5** | Ratio | $\dfrac{0.06}{2} = 0.03$ | $\dfrac{0.83}{4} = 0.21$ |
| **6** | Look for the smallest no. | **smallest** | largest |

$SCl_2$ is the limiting reactant



**- Step 2: compare between the limiting reactant (known) and the unknown product, to determine the theoretical yield.**

$$mass\ of\ C = \frac{mass\ of\ A}{molar\ mass\ of\ A}\ x\ molar\ mass\ of\ C\ x\ \frac{c}{a}$$

In this case, the unknown C is the $SF_4$, the known A is $SCl_2$ (Limiting Reactant), a equal 3 and c equal 1.

Use the formula: $$mass\ of\ SF_4 = \frac{mass\ of\ SCl_2}{molar\ mass\ of\ SCl_2}\ x\ molar\ mass\ of\ SF_4\ x\ \frac{1}{3}$$

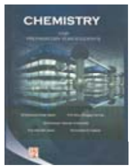
$$mass\ of\ CO(NH_2)_2 = \frac{6.0\ g\ of\ SCl_2}{103.0\ g/mole\ of\ SCl_2}\ x\ 108.0\ g/mole\ of\ SF_4\ x\ \frac{1}{3} = 2.45\,g$$

**So, the theoretical yield for the $SF_4$ equals 2.45 g.**

**In an experiment, the actual yield was 6.35 g and the theoretical yield was 8.3 g what is the percentage yield of the product?**

**Use the formula:** $$\%\ \text{yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100$$

$$\%\ \text{yield} = \frac{6.35}{8.3} \times 100 = 76.5\%$$



Many tons of urea ($CO(NH_2)_2$) are produced every year in fertilizer industries. When 119.0 g ammonia react with 80.0 g $CO_2$ as the equation:

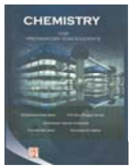
$$2\,NH_3\,(g) + CO_2\,(g) \rightarrow CO(NH_2)_2\,(s) + H_2O$$

**and produce 100.0 g urea, calculate % yield?**

**Step 1: determine the limiting reactant**

| Steps | | 2 $NH_3$ | 1 $CO_2$ |
|---|---|---|---|
| 1 | Mass /g | 119.0 | 80.0 |
| 2 | Molar mass | 17.0 | 44.0 |
| 3 | $\text{No. of moles} = \dfrac{\text{mass}}{\text{molar mass}}$ | $\dfrac{119.0}{17.0} = 7.0$ | $\dfrac{80.0}{44.0} = 1.8$ |
| 4 | Coefficient | **2** | **1** |
| 5 | Ratio | $\dfrac{7.0}{2} = 3.5$ | $\dfrac{1.8}{1} = 1.8$ |
| 6 | Look for the smallest no. | largest | **smallest** |

$CO_2$ is the limiting reactant



**Step 2: compare between the limiting reactant (known) and the unknown product, to determine the theoretical yield.**

$$mass\ of\ C = \frac{mass\ of\ A}{molar\ mass\ of\ A}\ x\ molar\ mass\ of\ C\ x\ \frac{c}{a}$$

In this case, the unknown C is the urea [$CO(NH_2)_2$], the known A is CO2 (Limiting Reactant), a equal 1 and c equal 1.

$$mass\ of\ \mathbf{CO(NH_2)_2} = \frac{mass\ of\ CO_2}{molar\ mass\ of\ CO_2}\ x\ molar\ mass\ of\ \mathbf{CO(NH_2)_2}\ x\ \frac{1}{1}$$

$$mass\ of\ CO(NH_2)_2 = \frac{80.0\ g\ of\ CO_2}{44.0\ g/mole\ of\ CO_2}\ x\ 60.0\ g/mole\ of\ CO(NH_2)_2\ x\ \frac{1}{1} = 109.0\,g$$

So, the theoretical yield for the urea equal 109.0 g

.

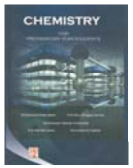
**Step 3: determine the % yield**

Use the formula:

$$\%\ \text{yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100$$

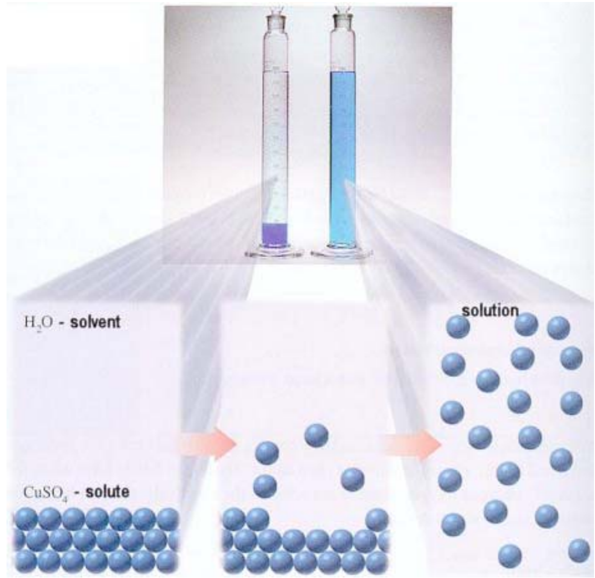
According to the problem, 100.0 g of the urea were produced (actual yield).

$$\%\ \text{yield} = \frac{100.0}{109.0} \times 100 = 91.7\,\%$$



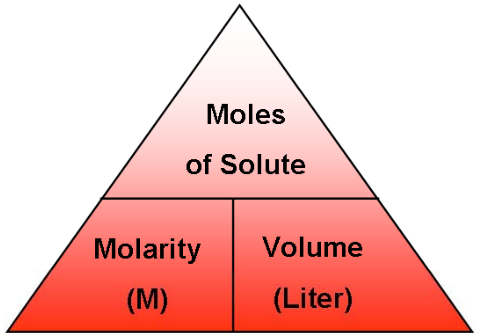
# SOLUTIONS & CONCENTRATION

- **A solution is a homogeneous mixture of 2 or more substances (gas, liquid, or solid) in a single phase and it contains a solute (the substance that is dissolved in a solvent) and a solvent (a liquid in which a substance is dissolved).**
- **When the solvent is water, the solution is said to be aqueous (aq).**

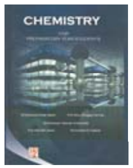



**Concentration of solution can be expressed in different ways:**

$$\text{Molarity (M)} = \frac{\text{moles of Solute}}{\text{volume of solution (liter)}}$$




$$\text{Weight\%} = \frac{\text{weight of solute}}{\text{weight of solution}} \text{x100}$$

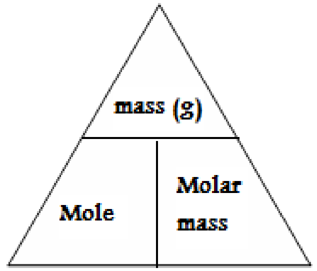


**Calculate the mass required to prepare a 250 mL 0.01 M solution of $KMnO_4$?**


**Convert 250 ml to L (250/1000 = 0.250 L)**
**Using the formula:**

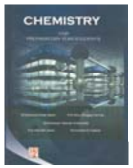
**# moles = molarity x volume**
**= 0.01 mol/L x 0.250 L**
**=0.0025 mol**

**Mass = # moles x molar mass**
**Molar mass of KMnO4 = 158.0 g/mole**
**Mass of $KMnO_4$ needed = 0.0025 mol x 158.0 g/mole**
**= 0.395 g of $KMnO_4$**
**So, weigh 0.395 g of $KMnO_4$ and dissolve them in 250 ml volumetric flask.**

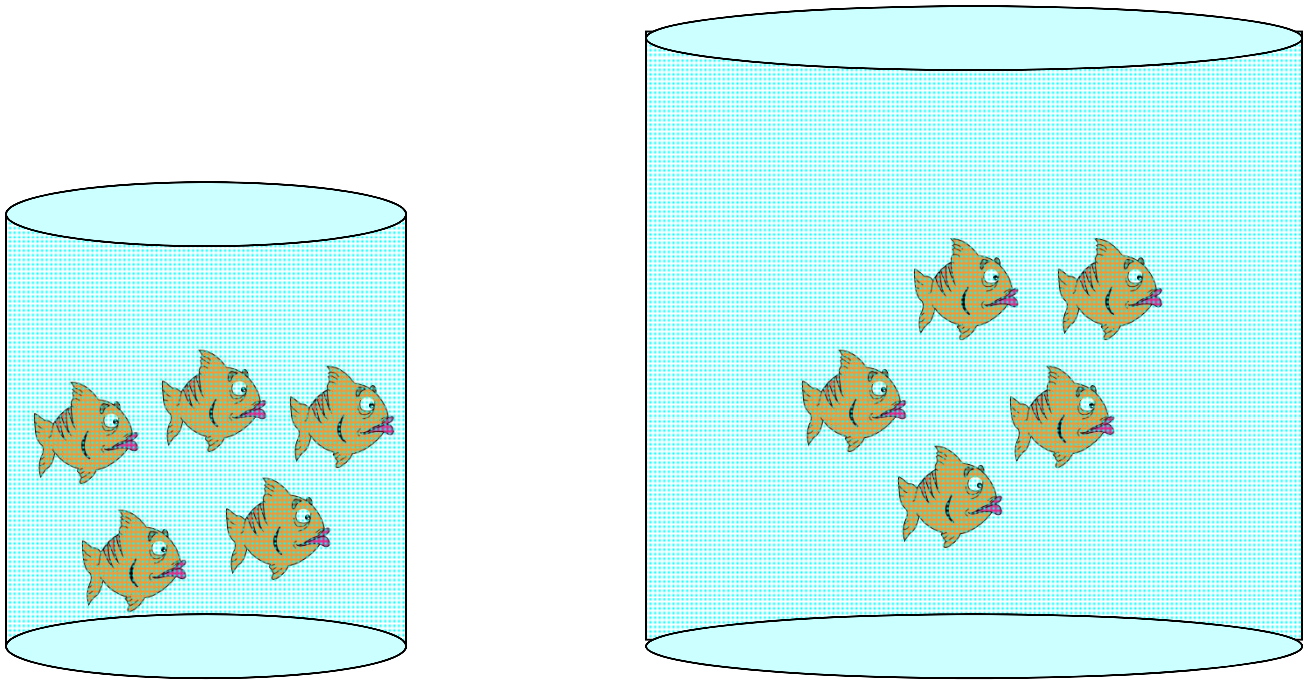



**If a solution contains 0.035 moles solute in 2.0 L of water, what is the molarity?**

Molarity (M) = moles of solute / volume of solution (liter)
= 0.035 moles / 2.0 L = $1.8 \times 10^{-2}$ M



# Dilution of concentrated solutions



**# fish = 5**
**Volume = 1 L**
**Concentration = 5 fishes/1 L**

**# fish = 5**
**Volume = 2 L**
**Concentration = 5 fishes/2 L**

- If you have 5 fishes in a 1 L tank and you moved them in another 2 L tank, what will happened?
- The number of the fishes remain the same (5 fishes), but their concentrations changes.
- The fishes are the moles, when you put same number of moles in different volumes, the number of moles stay the same, but the concentrations changed.

# DILUTION OF CONCENTRATED SOLUTIONS

- **When we dilute a solution by mixing it with more solvent, the amount of solute present does not change, but the total volume and the concentration of the solution do change.**
- **To calculate the molarity after dilution, we can use the following formula:**

## # moles before dilution = # moles after dilution

$$\textbf{(Molarity x Volume)}_{\text{before dilution}} = \textbf{(Molarity x Volume)}_{\text{after dilutio}}$$

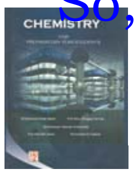
$$M_1 \times V_1 = M_2 \times V_2$$

How many milliliters of 18.0 M $H_2SO_4$ are required to prepare 1.00 L of a 0.900 M solution of $H_2SO_4$?

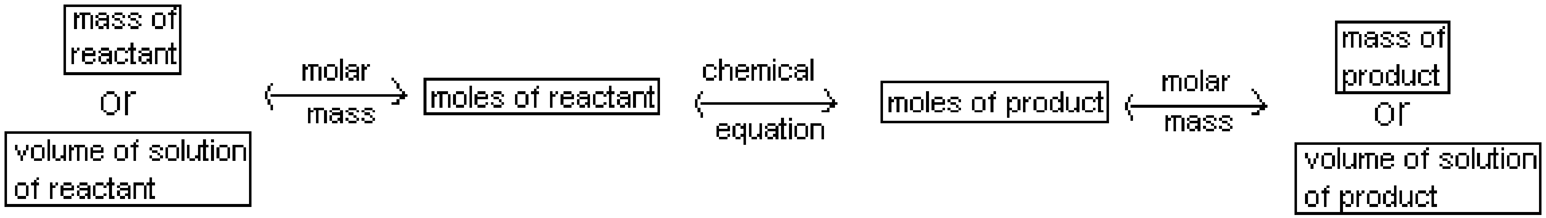
Using the formula: $M_1 \times V_1 = M_2 \times V_2$

$M_1$ = 18.0 M, $V_1$=?? And $M_2$ = 0.900 M, $V_2$ = 1.00 L

So,




$$V_1 = \frac{M_2 \times V_2}{M_1} = \frac{0.900\,\text{M} \times 1.00\,\text{L}}{18.0\,\text{M}} = 0.0500\text{L} = 50.0\text{mL}$$




**What volume of 1.5 M HCl is required to react with 34.6 mL of 2.44 M NaOH?**

$$NaOH_{(aq)} + HCl_{(ag)} \rightarrow NaCl_{(aq)} + H2O_{(l)}$$

**First calculate the number of moles of NaOH:**

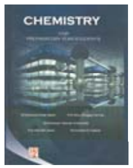
2.44 M X (34.6/1000)L = 0.0844 mole NaOH

**From the chemical equation:**

$$NaOH_{(aq)} + HCl_{(ag)} \rightarrow NaCl_{(aq)} + H2O_{(l)}$$

**One mole** of HCl reacts with **one mole** of NaOH
0.0844 mole HCl reacts with 0.0844 mole NaOH

Number of moles of HCl = molarity of HCl X volume of solution
0.0844moles HCl = 1.5 M X V
The volume of HCl= 0.056 L = 56 mL



**According to the reaction:**

$$Ba(OH)_{2\,(aq)} + 2\,HNO_{3\,(aq)} \rightarrow Ba(NO_3)_{2\,(aq)} + 2\,H_2O_{(l)}$$

**What volume of 0.5M $HNO_3$ is required to react with 41.77 mL of 0.1603 M $Ba(OH)_2$?**

From the chemical equation:

$$Ba(OH)_{2\,(aq)} + 2\,HNO_{3\,(aq)} \rightarrow Ba(NO_3)_{2\,(aq)} + 2\,H_2O_{(l)}$$

**2 Moles** of $HNO_3$ react with **one mole** of $Ba(OH)_2$

# moles of $Ba(OH)_2$ = molarity X volume of solution

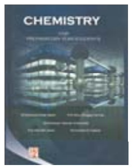
= 0.1603 M X (41.77/1000) L = $6.696 \times 10^{-3}$ mol

The moles of $HNO_3$ which reacted = 2 x $6.696 \times 10^{-3}$ = $13.39 \times 10^{-3}$ mol

# moles of $HNO_3$ = molarity X volume of solution

$13.39 \times 10^{-3}$ mol = 0.5 M X V

V= 0.0417 L = 41.7 mL



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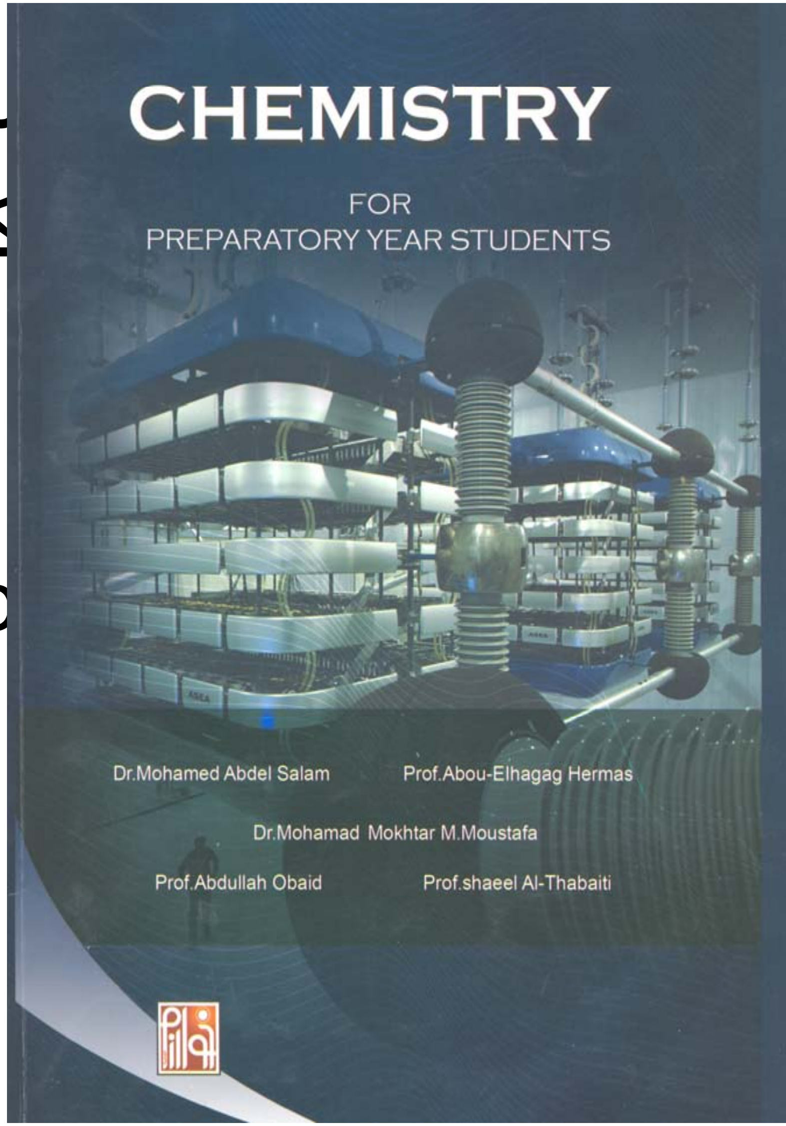


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