Ministry of Higher Education and Scientific Research

Al-Furat Al-Awsat Technical University

College of Health and Medical Technique /Kufa

Learning package In field of General Chemistry / Practical

Presented to the 1st class students of College of Health and Medical Technique /Kufa

Aesthetic and Laser Techniques Department

Designed by

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Safety

Safety Equipment: It is always essential to have all safety equipments in a lab. The most essential safety items required in a chemistry laboratory are safety glasses, laboratory coats, emergency showers, fumes hoods, eye wash stations, fire suppression equipment. All should be trained to use properly the safety equipments so that emergency situations can be controlled quickly.

General chemistry and lab. Safety: The chemistry laboratory can be a place of discovery and learning. However, by the very nature of lab. Work, it can be a place of danger if proper common-sense precautions are not taken. While every effort has been made to eliminate the use of explosive, highly toxic, and carcinogenic substances from the experiment which you will perform, there is a certain hazard associated with use of a variety of chemicals and glassware. You are expected to learn and adhere to the following general safety guidelines to ensure a safe laboratory environment for both yourself and the people you may be working near.

1. Safety goggles must be worn at all times while in the laboratory. This rule must be followed when you are actually working on an experiment or simply writing in your lab. Notebook.
2. Contact lenses are not allowed. Even when worn under safety goggles, various fumes may accumulate under the lens and cause serious injuries or blindness.
3. Closed toe shoes must be worn in the lab. Sandals and shorts are not allowed. Long hair must be tied back when using open flames.
4. Eating, drinking, and smoking are strictly prohibited in the lab.
5. Never taste anything, never directly smell the source of any vapor or gas, do not inhale these vapors but take in only enough to detect an odor if one exists.
6. Coats, backpacks, etc., should not be left on the lab benches and stools. Beware that lab chemicals can destroy personal possessions.
7. Always wash your hands before leaving laboratory.
8. Learn where the safety and first-aid equipment is located. This includes fire extinguishers, fire blankets, and eye-wash stations.
9. Notify the instructor immediately in case of an accident.

10. Know what chemicals you are using. Carefully read the label *twice* before taking anything from at bottle.

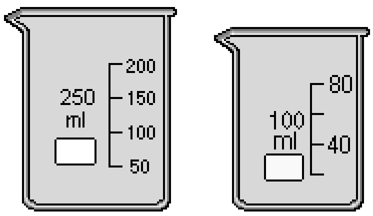
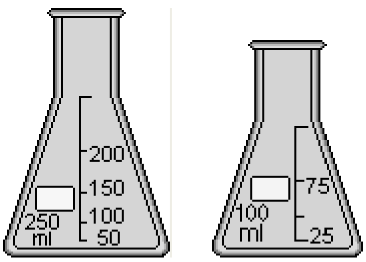
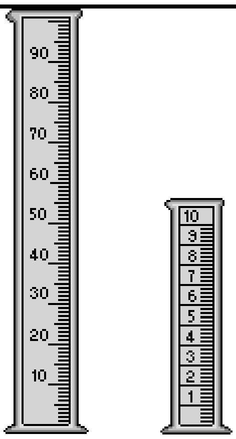
11. Excess reagents are never to be returned to stock bottles. If you take too much, dispose of the excess.
12. Many common reagents, for example, alcohols and acetone, are highly flammable. Do not use them anywhere near open flames.
13. Always pour acids into water. If you pour water into acids, the heat of reaction will cause the water to explode into steam.

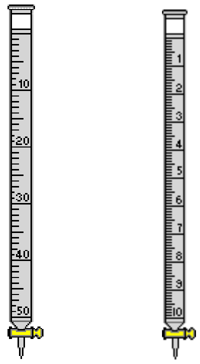
14. If chemicals come into contact with your skin or eyes, flush immediately with copious amounts of water and consult with your instructor.
15. Never point a test tube that you are heating at yourself or your neighbor.
16. Clean up all broken glassware immediately and dispose of the broken glass properly.
17. Turn off burners whenever you leave your workstation. Be sure that the gas is shut off when you leave the lab.
18. Beware of hot glass—it looks exactly like cold glass.

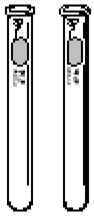
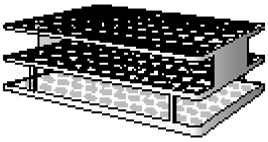
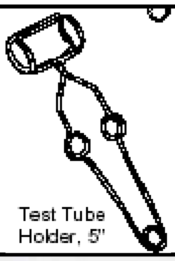
Laboratory equipment

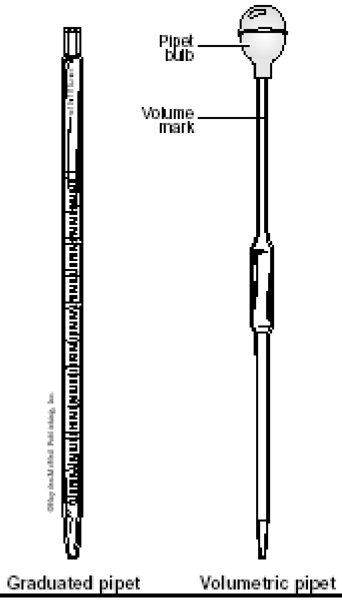
Generally, the laboratory equipment is made of glass, some are plastic and others are metallic.

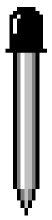
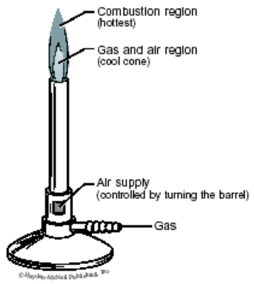
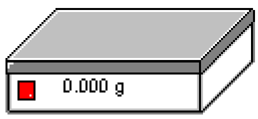
General Chemistry Equipments

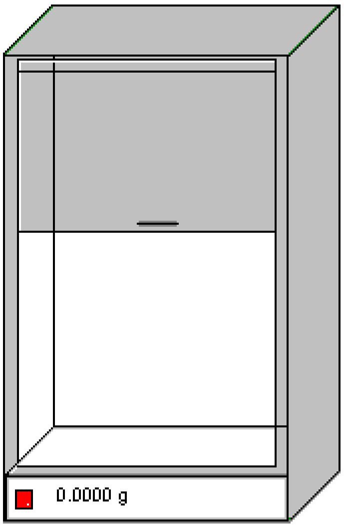
Picture	Name	Description
	**Beaker**	graduated container with a spout, it is used to Running reactions, mixing chemicals, heating chemicals
	**Conical Flask Erlenmeyer Flask**	graduated cone-shaped container that is used very frequently in laboratories, it used for running reactions, heating chemicals mixing chemicals – easier for mixing than beakers
	**Cylinder**	graduated tube with a spout that is used especially for accurately measuring the volume of liquids.

	**Burette**	Long graduated tube for measuring liquids with high precision, it is fitted with a valve for manually regulating the flow.
	**Rod**	Long metal part to which various laboratory devices can be clamped.
	**Holder**	part with a screw for attaching a clamp on to thestand's rod.
	**Centrifuge**	Is a laboratory device that is used for the separation of fluids, gas or liquid, based on density.
	**Stand**	Stand is used to hold items using clamps or rings
	**Clamp**	Used to catch items
	**Ring Stand**	Ring stands are used to hold items being heated.

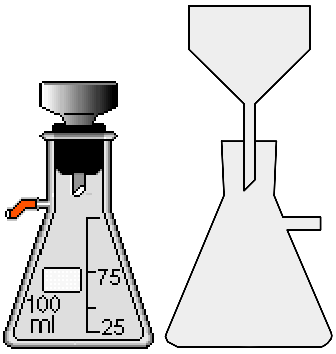
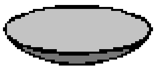
	**Test tube**	Cylindrical tube used for mixing and heating chemicals and running reactions– smaller quantities than beakers and flasks.
	**Test tube rack**	The test tube rack is used to hold test tubes while reactions happen in them or while they are not needed
	**Test tube holder**	The holder is used tohold test tubes whenthey are hot and untouchable.
	**Test tube Brush**	The test tube brush is used to easily clean the inside of a test tube.

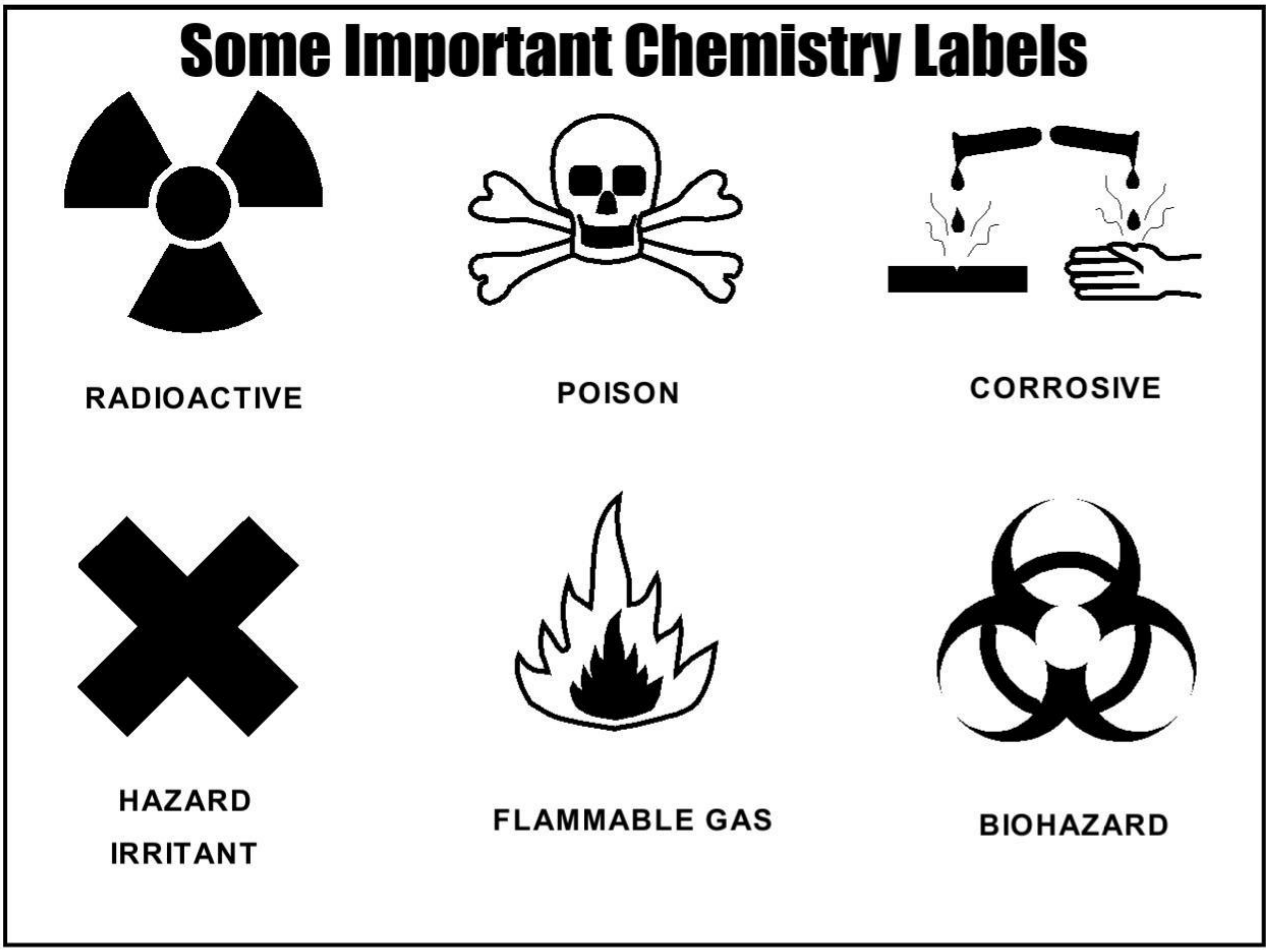
	**Pipette**	fine tube that is open at both ends, it is used to transfer very precise quantities of liquids from one container to another.

	**Volumetric Flask**	The Volumetric flask is used to mix chemicals to accurately determine concentration.
	**Dropper**	Use to add drops of liquid
	**Bunsen burner**	Bunsen burners are used for heating and exposing items to flame. They have many more uses than a hot plate, but do not replace a hot plate.
	**Electronic balance**	An electronic balance, Directly measures the mass of an object placed on the weighing pan.

	**High sensitivity balance**	Use to measure the mass of an object in high accuracy
	**Hot plate**	Use to heat objects
	**Water bath**	This chemistry equipment used for regulating the temperature of substances subjected to heat.

	**Thermometer**	The thermometer is used For measuring temperature. Alcohol have red bulbs, mercury have a silver colored bulb
	**Spatula**	The spatula is used to transfer solid chemicals
	Stirring rod	used to stir, assist in pouring liquids
	**Wash bottle**	Flexible container that is squeezed lightly to squirt a liquid, it is used especially for dispense distilled water and to wash the glassware

	**Buchner funneland Filtering Flask**	Used to rapidly filtration
	**Funnel**	Used for filtering and for adding chemicals without spilling. (Thefunnel can be used to target liquids into any container so they will not be lost or spilled).
	**Watch Glass**	The watch glass is used to hold solids when being weighed or transported. They should never be heated.
	**Filter Paper**	Used to filter solids from liquids


Some Important Chemistry Labels
RADIOACTIVE
POISON
CORROSIVE
HAZARD
IRRITANT
FLAMMABLE GAS
BIOHAZARD

Preparation of solutions

The purpose

To understand how to prepare a solution, and the calculations related with preparation process.

Introduction

Solution: is a homogeneous mixture created by dissolving one or more solute in a solvent. The chemical present in a smaller amount, the solute, is soluble in the solvent (the chemical present in a larger amount). Aqueous solution is a solution where water is the solvent.

Solutions should be a one phase system with more than one component (or species). The phase may be solid, liquid or gas.

solution = solute + solvent

examples of solutions

Solution phase	example
Solution of gas in gas	Air
Solution of gas in liquid	NH_3 in water
Solution of gas in solid	H_2 dissolved in pd
Solution of liquid in gas	Aerosol
Solution of liquid in liquid	Acid in water
Solution of liquid in solid	Mercury in copper
Solution of solid in gas	Smoke
Solution of solid in liquid	Sugar in water
Solution of solid in solid	Copper amalgam

Hint: Amalgam is a mixture of mercury and another metal, usually silver, that is used in dentistry to make fillings.

Solutions with accurately known concentrations can be referred to as **standard (stock) solutions**. These solutions are bought directly from the manufacturer or formed by dissolving the desired amount of solute into a volumetric flask of a specific volume. Stock solutions are frequently diluted to solutions of lesser concentration for experimental use in the laboratory**.**

A. Preparation of Liquid - Liquid Solutions

General method for preparation of diluted acid solution:

The normality of concentrated acid can be calculated from the information written on the bottle (percentage % w/w , specific gravity , equivalent weight) according to the equation:-

$$N = \frac{sp. \times \% \times 1000}{eq.wt}$$

To calculate the volume of conc. acid that should be taken (diluted) to prepare a specific volume of diluted acid in the selected normality we have to use the equation below:

$$N_1 \times V_1 = N_2 \times V_2$$

of conc. acid of dilute acid

Notes:

1. Dilution is usually carried out by a factor of ten exactly.
2. Density and Specific Gravity

 Density = mass/volume (The units are usually grams/mL).

 Specific Gravity = (mass of substance)/(mass of equal volume water)

 = (Density of substance)/Density of water)

Note that the temperature must be specified for specific gravity. Specific gravity is a unit less ratio (i.e., it doesn't matter what the units are so long the same units are used for the substance and water). Specific gravity is more often used in commerce and commercial reagent labeling than density.

General Procedure:

Transfer (x) mL of concentrated acid into volumetric flask (x) mL using cylinder or pipette, complete to the mark with distilled water and mix well.

For ex. Preparation of Approximately 0.1 N Hydrochloric acid

For HCl

%= 37% *Sp.gr=1.18* *eq.wt=36.5*

So that

Normality = 1.18 × 0.37 × 1000 / 36.5 = 11.96 N ~ 12N

To calculate the volume of (conc. HCl) that should be taken to prepare (250) mL of (0.1)N HCl

$$12N \times V_1 = 0.1N \times 250\ mL$$

So that

V_1 = 2.08 mL should be diluted with distilled water in 250 mL volumetric flask to obtain 0.1N HCl

B. Preparation of Solid - Liquid Solutions

We can express the concentration of solution prepared by dissolving a specific amount of solid substance in a specific volume of solution, in two ways formal concentration, and normal concentration.

$$formality = \frac{no.of\ fw}{1 liter\ of\ solution} \qquad normality = \frac{no.of\ equivalents}{1 liter\ of\ solution}$$

To calculate the weight need to prepare any one of these concentration in specific volume of solvent we use the equations:

For formal concentration $Weight = \frac{fw \times F \times v(mL)}{1000}$

For normal concentration $Weight = \frac{eq.wt \times N \times v(mL)}{1000}$

The normality of a solution is the gram equivalent weight of a solute per liter of solution.It may also be called the equivalent concentration. It is indicated using the symbol N.

$$N = \frac{Gram\ eq.of\ Solute}{Volume\ of\ sol.in\ litre} = \frac{Weight}{Equivalent\ weight} \times \frac{1000}{V\ ml}$$

$$\text{Equivalent Weight} = \frac{Molar\ Mass}{n}$$

Equivalent Weight of Acid, base and salts:

Equivalent weight = molecular weight / X

In the above formula X represents the valency factor.

1. Acids:

Taking an example of sulfuric acid as follows:-

$H_2SO_4 + 2OH^- \rightarrow 2H_2O + SO_4^{2-}$

The equivalent weight of the acid can be determined by determining the individual molecular weight of each of the elements from the periodic table and firstly adding them together.

This will give us the molecular weight of the acid. 2(1) + (32) + 4(16) = 98.0.

The acid is seen to be donating two protons as the sulfate ion is seen to acquire negative charges.

Therefore the equivalent weight of the acid would be 98.0/2 = 49.0.

2. Bases:

The reasoning for the base is the same for example in the case of calcium hydroxide base $Ca(OH)_2$

$Ca(OH)_2 \rightarrow Ca+ + 2OH^-$

The number of hydroxyl ions released by the calcium hydroxide base is 2.

Therefore, its valency factor value will be two.

The molecular weight of the calcium hydroxide base is 74.

As we know, Equivalent weight = molecular weight / X

The equivalent weight of calcium hydroxide base= 74 / 2 = 37

3. Salts:

For the salts, X (valency factor) is the total positive charge on the positive ion (cation).

In the case of silver carbonate salt (Ag_2CO_3)

$Ag_2CO_3 \rightarrow 2Ag^+ + CO^{3-}$

The total positive charge of the silver cation is two.

Therefore, its valency factor or X value will be two.

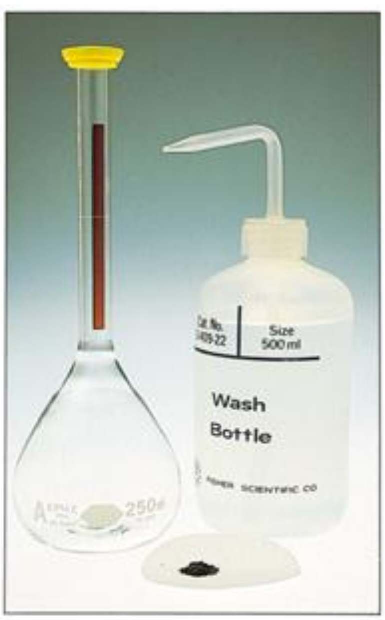
The molecular weight of the silver carbonate salt is 275.75 g/mol.

As we know, Equivalent weight = molecular weight / X

The equivalent weight of silver carbonate salt= 275.75 / 2 = 137.87.

General Procedure:

Dissolve (x) gm of solid substance into a small beaker with distilled water and transfer the solution after dissolution into the volumetric flask (x) mL , washing the beaker many times and adding the washing into the volumetric flask for quantitative transferring of the solution, complete to the mark and mix well.




***For ex. Preparation of 0.1 F Sodium Chloride* NaCl**

fw= 58.4 *F* = 0.1F *v (mL)*= 250 mL

The weight need is $Weight = \dfrac{58.4 \times 0.1 \times 250}{1000}$ =1.46 g

So that to prepare 250 mL of 0.1 F NaCl we have to dissolve of 1.46 g of NaCl in 250 mL distilled water.

Methods of Expressing the Concentration of Solution:

The concentration of a solution can be expresses in a number of ways. The important methods are:

1- Molarity: is the number of moles of a solute dissolved in a liter of solution.

2- Normality : is the number of equivalents of a solute dissolved in a liter of solution.

3- Formality : is the number of formula mass of a solute dissolved in a liter of solution.

4- Molality: the number of moles of solute dissolved in one kilogram of solvent.

(Molality = No. of moles / weight of solvent (kg))

5- Weight – Volume Percentage (% w/v)

$$\% \, w / v = \frac{\text{Weight of solute (g)}}{\text{Volume of solution (ml)}} \times 100$$

6- Weight – Weight Percentage (% w/w)

$$\% \, W / W = \frac{\text{Weight of solute (g)}}{\text{Weight of solution (g)}} \times 100$$

7- Volume – Volume Percentage (% v / v)

$$\% \, V / V = \frac{\text{Volume of solute (ml)}}{\text{Volume of solution (ml)}} \times 100$$

Q/ what are the molarity of (125 ml) solution containing of 0.050 moles of HCl ?

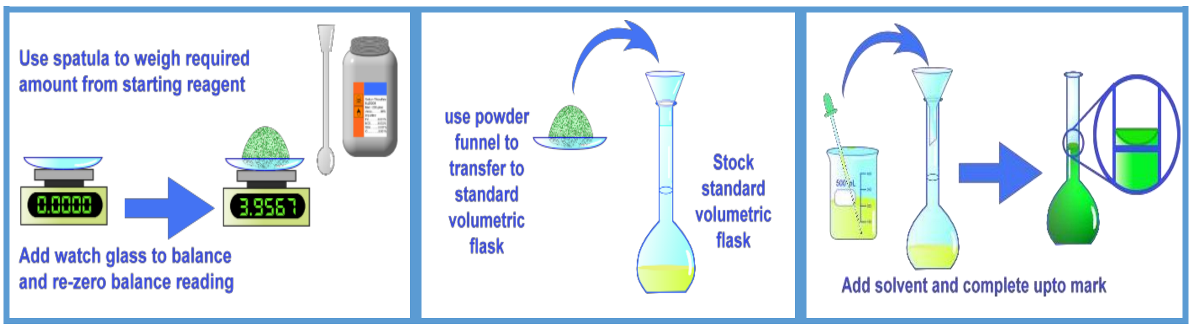
A/ before anything : convert ml to L

125ml/1000 = 0.125 L

M= n (mol) / V(L)

M= 0.050 / 0.125= 0.40 mol/L

Q/ you dissolve 152.5 gm of $CuCl_2$ in water to make solution with final volume of 2.25L what is the molarity ?



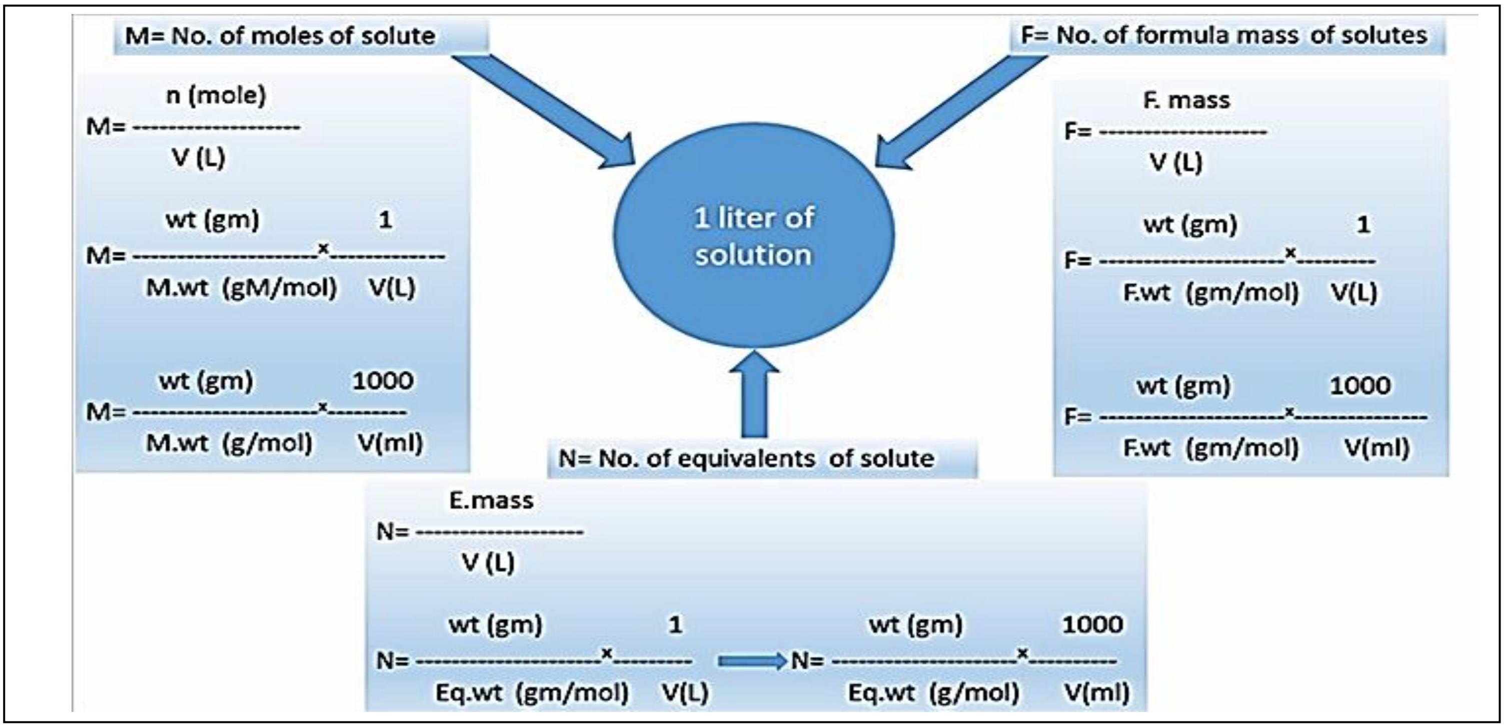
A/ M.wt ($CuCl_2$)= 63.55 + (2*35.44)= 134.45 gm/mol

n= wt / M.wt

n=152.5 / 134.45= 1.134 mol

M= n (mol) /v (L) = 1.134 /2.25

M= 0.504 mol/L


M= No. of moles of solute
M= n (mole) / V (L)
M= wt (gm) / M.wt (gM/mol) × 1 / V(L)
M= wt (gm) / M.wt (g/mol) × 1000 / V(ml)
1 liter of solution
F= No. of formula mass of solutes
F= F. mass / V (L)
F= wt (gm) / F.wt (gm/mol) × 1 / V(L)
F= wt (gm) / F.wt (gm/mol) × 1000 / V(ml)
N= No. of equivalents of solute
N= E.mass / V (L)
N= wt (gm) / Eq.wt (gm/mol) × 1 / V(L)
N= wt (gm) / Eq.wt (g/mol) × 1000 / V(ml)

Lab 3
General Chemistry

Preparation of percent solutions and Dilution of Solution

Objectives

1. To learn how to prepare percent solutions.

2. To get familiar with solution dilutions.

A. Percent solutions

- A percentage solution is a type of solution that expresses the amount of solute dissolved in a given amount of solvent as a percentage of the total solution.
- Percentage solutions are commonly used in a wide range of applications, including in the preparation of medical solutions, cleaning solutions, and industrial solutions. For example in medical applications, percentage solutions are used to prepare medications and intravenous fluids.
- It is important to note that the percentage solution is not a measure of the absolute amount of solute or solvent in the solution, but rather a measure of the relative amount of solute to solvent.

Preparation of percent solutions:

- One of the way to describe the concentration of a solution is by the percent of a solute in the solvent. The percent can further be determined in one of three ways:

1. Weight per Volume (w/v %) = $\frac{g\ of\ solute}{ml\ of\ Solution} \times 100\%$

2. Volume per Volume (v/v %)=$\frac{ml\ of\ solute}{ml\ of\ Solution} \times 100\%$

3. Weight per Weight (w/w %)=$\frac{g\ of\ solute}{g\ of\ Solution} \times 100\%$

- Where Per means "for every one" and Cent means 100

w/v %

It is the number of grams of solute dissolved in 100 mL of solution.

For example, 3% of NaOH, means 3 grams of NaOH is dissolved in 100 ml of the solution.

Example :Prepare 50 ml of 4%(w/v) NaOH solution

- 4g → 100 ml
- ? → 50 ml
- Weight = $\frac{4 \times 50}{100} = 2g$
- To prepare the solution, 2 grams of NaOH is dissolved in little water and the volume made up to 50 ml

Practically,

1. Place a beaker in a balance and zero the balance.

2. Weight 2 grams of NaOH , in the beaker and dissolve in a very small volume of water ,once the solid is dissolved, the volume is transferred to 50 ml volumetric flask.

3. Wash the beaker at least 2 times with small amount of distilled water and transfer it to the volumetric flask, to make sure all the solute is dissolved and there is no left overs.

4. Bring up to a final volume 50 ml.

W/W%

It is the number of grams of solute dissolved in 100 gram of solution

- The concentrations of many commercial acids are giving in terms of w/w%.

- In order to calculate the volume of the stock solution required for a given preparation the density (specific gravity) of stock solution should be provided.

Example: Prepare 500 g of a 5% (w/w) solution of NaCl

Percent strength is 5% w/w, total weight desired is 500g

- 5g → 100 g
- ? → 500g
- Weight = $\frac{5 \times 500}{100}$ =25 g NaCl needed
- 500 g – 25 g = 475 g amount of solvent needed
- Dissolve 25 g of NaCl in 475 g of water
- Then the total weight of solution will be equal to 500g

+

v/v %

It is the number of milliliters of solute dissolved in 100 mL of solution.

This is used when both chemicals in a solution are liquid , For example, the 10% ethanol means that there is 10 mL of ethanol dissolved in 100 mL of the solution.

Example: prepare 20 ml of 0.2 %acetic acid?

- 0.2ml → 100 ml
- ? → 20ml

- $\text{volume} = \frac{0.2 \times 20}{100} = 0.04\text{ml}$
- That mean to prepare the solution, 0.04 ml of acetic acid should be completed to 20 ml with H2O (D.W)

Ex(1): Calculate the wt. percent (w/w)% of Na2SO4 solution which prepared by dissolving 25g in 200g of solution.

A/

$$\text{wt. percent (w/w)} = \frac{25\text{g}}{200\text{g}} \times 100 = 12.5\% \text{ (w/w)}$$

Ex(2): Calculate the volume percent (v/v)% of solution prepared by dissolved 50ml of methanol in 250ml of water.

A/

$$\text{volume percent(v/v)} = \frac{50}{50+250} \times 100 = 16.7\% \text{ (v/v)}$$

B. Solution dilution

1)Volume to volume dilutions :

- This type of dilutions describes the ratio of the solute to the final volume of the dilute solution

Example: Prepare 30 ml of a 1:50 dilution of the 0.07M NaOH solution

- 1ml---------→50 ml
- ?------------→30 ml
- (1x 30)/50ml= 0.6ml

0.6 ml of the starting solution (0.07M NaOH) is needed and volume made up to a final volume of 30 ml.

2)Preparing dilutions by using the C1v1=C2 v2 formula

V_1= Volume of starting solution needed to make the diluted solution.
C_1= Concentration of starting solution.
V_2= Final volume of diluted solution.
C_2= Final concentration of diluted solution.

Example: Prepare 50 ml of a 2.5x10-3M from prepared 0.4M HCl.

- C1 x V1 = C2 x V2
- 0.4 x V1 = 2.5x10-3 x 50
- V1=0.337 ml

0.337 ml of the starting solution is taken and final volume made up to 50 ml by the addition of water.

Volumetric Methods
Titration

An Overview of titration

Volumetric or titrimetric analyses are quantitative analytical techniques which employ a titration in comparing an unknown with a standard. In a titration, a volume of a standardized solution containing a known concentration of reactant "A" is added incrementally to a sample containing an unknown concentration of reactant "B". The titration proceeds until reactant "B" is just consumed (stoichiometric completion). This is known as the *equivalence point*. At this point the number of equivalents of "A" added to the unknown equals the number of equivalents of "B" originally present in the unknown.

To be successful, a titration must involve a chemical reaction between the two solutions which are being mixed. This chemical reaction should be simple, rapid and complete. Although a titration may involve an acid base reaction, a precipitation reaction, a complexation reaction, or an oxidation reduction reaction.

Titration : The gradual addition of one solution to another until the reaction is complete (as indicated by an “indicator ”).

Analyte: substance to be analyzed or to be determined.

Titrant: reagent of know concentration.

Equivalence point : Is the point at which the chemical reaction is complete (theoretical end of titration)

Endpoint : Is the point at which the indicator dictates that the reaction is complete by practically observed.

Unfortunately, the endpoint and the equivalence point are not exactly the same. The difference between the two is called the titration error.

Indicators : Indicators are organic compounds that have different colors in solutions of different pH.

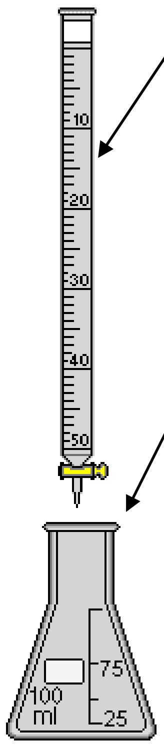
Standard solutions : A standard solution is a solution of known and exactly defined concentration. A primary standard can be used to produce a standard solution

Primary Standards :

A highly purified compound that serves as a reference material in volumetric methods. The basic requirements of primary standard substance are:

1. Obtainable in high purity.
2. Stable in air, can be stored, nonhygroscopic.
3. Absence of hydrated water.
4. Low cost.
5. Moderate solubility.
6. Large molecular weight.
7. Reacts stoichiometrically and rapidly.

General Procedure of Titration

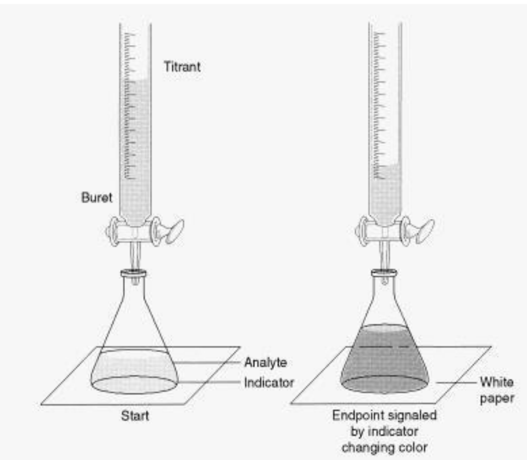


Burette

A graduated cylinder with a stopcock at the bottom. The graduations start at 0 at the top, so it measures how much volume is *added* to the conical (Erlenmeyer) flask. It fill with either acid or base (usually of known concentration) called the *titrant*.

Conical (Erlenmeyer) flask

- Fill with the substance being titrated (usually unknown concentration), 10 to 25 mL of the substance measured with a graduated cylinder or pipette.
- Add the indicator to the Erlenmeyer flask – 2 drops of indicator, a pH meter, thermometer, etc.
- Swirl the flask as you add the titrant (or use a magnetic stirrer).
- Add titrant until the endpoint is reached (color change, a steep increase in pH).



Procedure:

The analyte is prepared by dissolving the substance being studied into a solution. The solution is usually placed in a flask for titration. A small amount of indicator is then added into the flask along with the analyte. The reagent(titrant) is usually placed in a burette and slowly added to the analyte and indicator mixture. The amount of reagent used is recorded when the indicator causes a change in the color of the solution. The endpoint is the point where all of the analyte has be reacted with the reagent.

Volumetric (titration) calculations:

By using this following relation:

$$\left(\frac{MaVa}{na} = \frac{MbVb}{nb}\right)$$

Where (na) and (nb) are (stoichiometric mole ratios from the chemical equation).

Example: Calculate the concentration of 100.0 mL of NaOH solution titrated to the end point with 75.8 mL of a 0.100 M standard solution of H_2SO_4.

$H_2SO4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$

Sol/

n (H_2SO4)=1 ,n (NaOH)= 2

$$\frac{MV(H2SO4)}{n(H2SO4)} = \frac{MV(NaOH)}{n(NaOH)}$$

$$\frac{o.1M \times 75.8ml}{1} = \frac{M \times 100ml}{2}$$

M(NaOH)=0.1516

Acid-base titration: -

this method depends on the neutralization between an acid and a base when mixed in solution by using an appropriate indicator. The acid-base indicator indicates the endpoint of the titration by changing color. The endpoint and the equivalence point are an important because the equivalence point is determined by the stoichiometry of the reaction while the endpoint is just the color change from the indicator. the pH of the endpoint can be estimated using the following rules:

A strong acid will react with a strong base to form a neutral solution (pH = 7). Titrate 25 ml portions of the approximately 0.1 N solution of sodium hydroxide with the approximately 0.1 N hydrochloric acid. Use methyl orange, methyl red, bromothymol blue and phenolphthalein as indicators. **A strong acid will react with a weak base to form an acidic solution (pH < 7).** Titrate 25 ml portions of the approximately 0.1 N solution of ammonia with the standard 0.1 N hydrochloric acid. Use methyl red and phenolphthalein as indicators.

A weak acid will react with a strong base to form a basic solution (pH > 7). Titrate 25 ml portions of the approximately 0.1 N solution of acetic acid with the standard solution of 0.1 N sodium hydroxide. Use methyl red and phenolphthalein as indicators.

Indicator

A useful indicator has a strong color that changes quickly near its pKa. These traits are desirable so only a small amount of an indicator is needed. If a large amount of indicator is used, the indicator will effect the final pH, lowering the accuracy of the experiment. The indicator should also have a pKa value near the pH of the titration's endpoint. Choosing an indicator with a pKa near the endpoint's pH will also reduce error because the color change occurs sharply during the endpoint where the pH spikes, giving a more precise endpoint.

Acid-Base Titrations

Standardization is the process of determining the exact concentration of a solution. This process is used when the (titrant solution) it is not pure and its purity is accurately unknown in this case we weight grams of the substance and dissolve it in the desired volume and then standardize this solution with another standard reagent solution to find its real concentration. This is what is called standardization.
Note/ that if the substance is pure or its purity is well known there will be no need for standardization.
As mentioned earlier the solution, which is usually placed in the burette is called standard or reagent or titrant solution. Its concentration must be exactly known because any error in the preparation of this solution will be reflected on the result of the analysis and therefore needs to be a 100 % true.

Standardization of Approximately 0.1N HCl Solution

Introduction:

HCl stock solution can be standardized using a primary standard solution. Two important methods are usually used; against sodium carbonate, or sodium tetraborate decahydrate (borax). The second substance is preferred due to its relatively large molecular mass, superior purity, non-hygroscopic characters, and sharpness of the end point at room temperature.

a. Against Na_2CO_3

Sodium carbonate reacts with HCl according to the equation:

$$Na_2CO_3 + 2\,HCl \longrightarrow 2\,NaCl + CO_2 + H_2O$$

This reaction occurs if enough HCl is used which can be visualized using methyl orange indicator. Usually, an appropriate weight of Na_2CO_3 is dissolved in water and titrated against HCl to determine the normality of the acid.

b. Against Sodium Tetraborate

Sodium tetraborate (borax) can be used for neutralization of acids as mentioned previously. Usually, methyl red is the preferred indicator where boric acid is the result of the reaction as the following equation shows:

$$Na_2B_4O_7\cdot 10H_2O + 2HCl \rightarrow 2NaCl + 5H_2O + 4H_3BO_3$$

Principle:

In this experiment, we will use solid Na_2CO_3 (sodium carbonate), a primary standard, to standardize an HCl solution by titration. The reaction is an acid-base neutralization in which sodium chloride and carbonic acid are produced. The carbonic acid decomposes to give carbon dioxide and water. The equation for the reaction is:

$$2HCl_{(aq)} + Na_2CO_{3(aq)} \dashrightarrow 2NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$$

Titration consists of the careful addition of one solution from a burette (HCl) to another substance in a conical flask (Na_2CO_3) until all of the substance in the flask has reacted. The solution added from the burette is called the titrant. For every titration, there must be a way to determine when the titration reaction is complete. In acid-base titrations, this is accomplished by adding a small amount of an organic dye, called an indicator, to the solution to be titrated. If the indicator is chosen correctly, a color change, called the end point, occurs when the moles of acid equal the moles of base. the indicator methyl orange will be used. It changes from yellow to red.

Procedure:

A. Preparation of 0.1 N HCl

The normality of concentrated hydrochloric acid can be calculated from the information written on the bottle (percentage %, specific gravity(density), equivalent weight) according to the equation: -

Preparation of approximately 0.1 N Hydrochloric acid:

% HCl = 37%   Sp.gr=1.18   eq.wt=36.5

So that

$$\text{Normality} = (\text{sp. gr} \times \% \times 1000)/(\text{eq. wt})$$
$$= 1.18 \times 0.37 \times 1000 / 36.5 = 11.96\ N \sim 12N$$
$$(12N \times V_1 = 0.1N \times 250\ mL)$$

Transfer (2.08) mL of concentrated hydrochloric acid into volumetric flask (250) mL using cylinder, complete to the mark with distilled water and mix well.

CAUTION: HYDROCHLORIC ACID CAN CAUSE CHEMICAL BURNS IN ADDITION TO RUINING YOUR CLOTHING. IF YOU SPILL ANY OF THE ACID ON YOU, WASH THE CONTAMINATED AREA THOROUGHLY AND REPORT THE INCIDENT TO YOUR INSTRUCTOR.

B. Preparation of the Sodium Carbonate Solution

1- Dissolve (2.64) gm of Na_2CO_3 into a small beaker with distilled water and transfer the solution after dissolution into the volumetric flask (250) mL, washing the beaker many times and adding the washing into the volumetric flask for quantitative transferring of the solution, complete to the mark and mix well.
2- Transfer (10) mL of Na_2CO_3 solution into a clean conical flask using pipette.
3- Add two or three drops of methyl orange indicator into the conical flask (the color is yellow).

C. Doing the Titration

1. Rinse a burette with at least two 3 mL portions of your HCl solution. Discard each portion into a waste beaker. Completely fill the burette with the HCl solution and remove any air in the tip by running out some of the liquid into the waste beaker. Make sure that the lower part of the meniscus is between zero and 1.00 mL on the burette. Allow the burette to stand for at least 30 seconds before reading the exact position of the meniscus. Remove any hanging drop from the burette tip by touching it to the side of the waste beaker. Record the initial burette reading.

2. Place the conical flask containing the sodium carbonate solution under the burette with the capillary tip inside the mouth of the conical flask. Place a piece of white paper under the flask. Add the HCl carefully until the solution just starts to turn red and then record the burette reading.

3. Repeat the procedure for three trials.

Calculations:

Conical flask Burette

(Na_2CO_3 solution) (HCl solution)

$$N_{base} \times V_{base} = N_{acid} \times V_{acid}$$

N_{base}= 0.1 N

V_{base}= 10 mL

N_{acid}= unknown

V_{acid}= (x)mL from burette

So that

$$Nacid=Nbase \times Vbase/ Vacid$$

Calculate the mean normality of your three trials.

Mean X =X1+X2+X3/n

n = number of trials

Volume of HCl	Normality of HCl
X =	

Determination the accuracy

Accuracy is the closeness of the results to the real value or theoretical value. The percentage of error or relative error (E_{rel}%) is the measure of the accuracy.

$$E_{rel}\% = \frac{O - A}{A} \times 100$$

E = Absolute error

O = Observed value

A = accepted value (Theoretical value)

Preparation and Standardization of Approximately 0.1 N Sodium Hydroxide

Introduction:

Sodium hydroxide NaOH is not a primary standard substance because of:

1- Absorb CO_2 from air
2- Highly hygroscopic (which mean absorb moisture from the atmosphere)
3- Low purity

The presence of carbon dioxide CO_2 is one reason why we cannot weight out pure NaOH and use it as a primary standard. The reaction of NaOH with atmospheric CO_2 is as follows:

$$CO_2 + 2OH \rightarrow CO_3^{2-} + H_2O$$

so that a layer of carbonate will coat NaOH.

Therefore, NaOH solutions are prepared approximately and then standardized using a primary standard such as potassium hydrogen phthalate ($KHC_8H_4O_4$) or by using a secondary standard such as hydrochloric acid HCl.

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

Procedure:

1- Dissolve (1) gm of NaOH into a small beaker with distilled water and transfer the solution after dissolution into the volumetric flask (250) mL, washing the beaker many times and adding the washing into the volumetric flask for quantitative transferring of the solution, complete to the mark and mix well.
2- Transfer (10) mL of NaOH solution into a clean conical flask using pipette.
3- Add two or three drops of Phenolphthalein indicator into the conical flask (the color is pink).
4- Washing the burette with water and finally wash it with D.W and then with small portions of HCl (0.1) N itself. Then fill the burette with HCl and avoid the bubbles. Adjust the level of HCl in the burette at zero or any other position.

5- Start titration against the HCl solution very carefully until the color change from pink to colorless, then record the burette reading.

Calculations:

Conical flask		Burette
(NaOH solution)		(HCl solution)

$$N_{base} \times V_{base} = N_{acid} \times V_{acid}$$

N_{base}= unknown

V_{base}= 10 mL

N_{acid}= 0.1 N

V_{acid}= (x)mL from burette

So that

$$N_{base} = N_{acid} \times V_{acid} \ / V_{base}$$

Examples of some indictors:

Indicator	Color on acidic side	Color on basic side
Methyl violet	Yellow	Violet
Bromophenol blue	Yellow	Blue
Methyl orange	Red	Yellow
Methyl red	Red	Yellow
Litmus	Red	Blue
Bromothymol blue	Yellow	Blue
Phenolphthalein	Colorless	Pink

Importance of volumetric analysis:

1. High precision is obtained.
2. Simple apparatus is required.
3. Easy process and fast result.
4. Different methods for different types of substance