CHAPTER

2 Basic Terminology and Conversion-Formulas & Equations

2.1 GRAM EQUIVALENT CONCEPT

Equivalent Weight (E)

Equivalent weight

It is defined as the mass of an element/compound/ion which combines or displaces 1 part of hydrogen or 8 parts of oxygen or 35.5 parts of chlorine by mass.

$$\text{Equivalent weight,} \quad E = \frac{\text{Molecular weight}}{\text{Valence factor}} = \frac{M_o}{x}$$

Determination of equivalent weight amounts to determining valence factor, x.

In the case of an element, the equivalent weight is defined as:

$$\text{Equivalent weight,} \quad E = \frac{\text{Atomic weight}}{\text{Valency}} = \frac{A}{x}$$

Note that atomic weight substitutes molecular weight and valency substitutes valence factor in the definition. Valencies of hydrogen, calcium and oxygen are 1, 2 and 2 respectively. Hence, their equivalent weights are 1/1 = 1, 40/2 = 20 and 16/2 = 8 respectively.

Equivalent Weight of an Acid

The valence factor of an acid is equal to its basicity. The basicity of an acid is equal to the furnishable hydrogen ion (proton) in its aqueous solution. Importantly, basicity is not same as the number of hydrogen atoms in acid molecule. Consider acetic acid (CH_3COOH). It contains 4 hydrogen atoms in it, but only 1 furnishable hydrogen ion. As such, basicity of acetic acid is 1. With this background, equivalent weight of an acid is defined as:

$$\text{Equivalent weight,} \quad E = \frac{\text{Molecular weight of acid}}{\text{Basicity}}$$

The number of electronic charge on carbonate ion (CO_3^{2-}) is 2. Hence, equivalent weight of carbonate ion is (12 + 3 × 16)/1 = 60/2 = 30. Similarly, equivalent weight of aluminium ion (Al^{3+}) is 27/3 = 9.

Equivalent Weight of an Oxidizing or Reducing Agent

In a redox reaction, one of the reacting entities is oxidizing agent (OA). The other entity is reducing agent (RA). The oxidizer is recipient of electrons, whereas the reducer is the releaser of electrons. The valence factor for either an oxidizing or reducing agent is equal to the number of electrons transferred from one entity to another.

$$\text{Equivalent weight,} \quad E = \frac{\text{Molecular weight of compound}}{\text{Number of electrons transferred in redox reaction}}$$

Alternatively,

$$\text{Equivalent weight,} \quad E = \frac{\text{Molecular weight of compound}}{\text{Change of oxidation number in redox reaction}}$$

Potassium dichromate in acidic medium is a strong oxidizer. It means it gains electrons during redox reaction. Potassium dichromate in acidic solution results in:

$$K_2Cr_2O_7 + 14H^+ + 6e^- \rightarrow 2K^+ + 2Cr^{3+} + 7H_2O$$

$$\text{Equivalent weight of } K_2Cr_2O_7 = \frac{294.2}{6} = 29$$

Study of redox reaction is in itself an exclusive and extensive topic.

Gram Equivalent (geq)

It is equal to the mass in grams numerically equal to equivalent weight. If the mass of a chemical entity is "*g*" grams, then the given mass contains gram equivalents given by:

$$\text{Gram equivalent (geq)} = \frac{g}{E}$$

This formula is widely used to express grams of substance in terms of gram equivalent and vice versa.

Relation Between Moles and Gram Equivalents (geq)

Gram equivalents is given by: $\text{geq} = \frac{g}{E}$

Substituting for equivalent weight, we have:

$$\text{geq} = \frac{g}{E} = \frac{xg}{M_o}$$

Moles is given by:

$$n = \frac{g}{M_o}$$

Preparation of Acids and Alkalis of Particular Normality

When a stock solution of acid/alkali with known normality is available, a solution of particular normality can be prepared by using the relationship given below:

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

where N_1 = Normality of stock solution

V_1 = Volume of stock solution

N_2 = Normality of particular solution

V_2 = Volume of particular solution

Example If it is desired to make 100 mL of N/5 acid solution from a stock solution of 1 N acid strength, the calculation will be as follows:

$$1 \text{ N} \times \text{V1} - 1/5 \text{ N} \times 100$$

$$V_1 = 20 \text{ mL}$$

Twenty mL of 1N acid should be diluted to 100 mL with DDW and thoroughly mixed to obtain N/5 strength of acid. Thus, solution should be standardized against a weighted sample of a primary standard.

Preparation of Acid and Alkali Solutions

1. ***Redistilled water:*** Prepare by redistilling single-distilled water from an all purex distillation apparatus.
2. ***Detonized distilled water:*** Slowly pass distilled water through a 25 cm column of glass rubbing (1 to 2.5 cm in diameter) that has been charged with 2 parts of volume of a strongly basic anion exchange resin in the hydroxyl form and 1 part by volume of a strongly acidic cation-exchange resin in the hydrogen form. Use ion-exchange resins of a quality suitable for analytical work.

Chemical Reagents

Prepare the following reagents by cautiously adding the required amount of concentrated acid or alkali to the designated volume of the proper type of distilled water : mix thoroughly, and dilute to 1000 mL.

Acid solutions

See Table 1.1 for preparation of HCl. H_2SO_4 and HNO_2 solutions.

Combining expressions, we have:

$$\text{geq} = xn$$

gram equivalent = valence factor *X*moles

2.2 NORMAL SOLUTION

The solution that contains one gram equivalent weight of a substance per liter of solution is known as normal solution.

Example 1 N solution of NaOH – gram, molecular weight of NaOH = 40 g

Acidity of NaOH = 1

∴ Equivalent weight of NaOH = 40

When 40 g of NaOH is dissolved in DDW and the final volume of the solution is made up to 1L, this solution is known as 1 N solution of NaOH.

Normality: Normality of a solution is designed as the number of gram equivalents of a substance dissolved per liter of the solution and is represented by the symbol *N*.

$$\text{Normality} \qquad N = \frac{\text{Gram equivalent of substance}}{\text{Volume of solution, } L}$$

Molar solution: The solution that contains gram molecular weight of a substance dissolved in 1 L of solution is called molar solution.

Example 1 M solution of glucose –

Gm. Molecular weight of glucose, $C_2H_{12}O_6$ – 180 g

When 180 g glucose is dissolved in DDW and the final volume of the solution is made up to 1 L, this solution is known as 1 M solution of glucose.

Molarity: Molarity of a solution is defined as the number of gram moles of the substance dissolved per liter of the solution and is represented by the symbol.

$$\text{Molarity,} \qquad M = \frac{\text{Gram moles of substance}}{\text{Volume of solution, } L}$$

Relation between normality and molarity of a solution

$$\text{Normality} = \text{Molarity} \times \frac{\text{Molecular mass of substance}}{\text{Equivalent mass of substance}}$$

For acids: normality = molarity × basicity of acid

For bases: normality = molarity × acidity of base

Mass percentage (w/v): Mass percentage is defined as the number of parts by the mass of are substance per hundred parts by volume of solution.

Example 1 g NaOH is dissolved in DDW. The final volume of the solution is brought up to 100 mL. This solution is said to be 1-% (*w*/*v*) NaOH solution.

Gram Molecular Mass

Molecular mass expressed in grams is numerically equal to gram molecular mass of the substance. Molecular mass of O_2 = 32 gram molecular mass of O_2 is the relative molecular mass expressed in grams. Note that relative atomic mass is a ratio and has no unit while gram molecular mass and gram atomic mass are expressed in grams.

Mole Concept

Since it is not possible to calculate the weight of particles individually, a collection of such particles called mole is taken for all practical purposes. It was discovered that the number of atoms present in 12 g of carbon of ^{12}C isotope is 6.023×10^{23} atoms. This is referred to as Avogadro's number after the discoverer, Avogadro. A mole of a gas is the amount of a substance containing 6.023×10^{23} particles. It is a basic unit of the amount or quantity of a substance. The substance may be atoms, molecules, ions or group of ions.

Avogadro discovered that under standard conditions of temperature and pressure, (1 atm and 273 K) a sample of gas occupies a volume of 22.4 L.

Molar Volume

One mole of any gas at STP will have a volume of 22.4 L. This is called molar volume.

The molar volume [22.4 L at STP] plays a vital role in stoichiometric calculations because it is the link between volume and mass in reactions involving gases.

Relationship Between Gram Molecular Weight and Gram Molecular Volume

Gram molecular weight (GMW) or mole is the relative molecular mass of a substance expressed in grams. It is also called gram molecular weight of that element.

Gram molecular volume (GMV) or molar volume is the volume occupied by one gram molecular weight of a gas at STP (standard temperature and pressure).

All gases of equal volumes contain same number of molecules under the same conditions of temperature and pressure. Equal number of molecules of different gases will occupy equal volumes under the same conditions of temperature and pressure.

One mole of gas = 6.023×10^{23} molecules

$$\text{Gram molecular volume} = \frac{\text{Gram molecular weight}}{\text{Weight/volume of gas at STP}}$$

$$\therefore \text{ Molar volume of } O_2 = \frac{32}{1.429 \text{ g/L}} = 22.4 \text{ L}$$

$$\therefore \text{ Molar volume of } H_2 = \frac{2.016}{0.09 \text{ g/L}} = 22.4 \text{ L}$$ 1 mole of a gas = 22.4 L at STP

Table 2.1 *Preparation of uniform acid solutions**

Desired Component	*Hydrochloric Acid (HCl)*	*Sulfuric Acid (H_2SO_4)*	*Nitric Acid (HNO_3)*
Specific gravity (20/4C) if ACS grade conc. acid	1.174 – 1.189	1.834 – 1.816	1.409 – 1.418
Percent of active ingredient in conc. reagent	16 – 17	96-98	69-70
Normality of conc. reagent	11 -12	16	15 – 16
Volume (mL) of conc. reagent to prepare 1 L of			
18 N solution	–	500 (1 + 1) †	–
6 N solution	500 (1 + 1) †	167 (1 + 5) †	180
1 N solution	83 (1 + 1) †	28	64
0.1 N solution	8.1	2.8	6.4
Volume (mL) of 6 N reagent to prepare 11 of 0.1 N solution	17	17	17
Volume (mL) of 1 N reagent to prepare 11 of 0.02 N solution	20	20	20

* All Values Approximate.

† *a* + *b* system of specifying preparatory volumes appears frequently throughout this manual and means that if the volume of the concentrated reagents is diluted with k volumes of diluted water to form the required solution.

(i) Stock sodium hydroxide, NaOH, 15 N (for preparation of 6 N, 1 N, and 0.1 N solutions.) Cautiously dissolve 625 g solid NaOH in 800 mL distilled water to form 1 liter of solution. Remove sodium carbonate precipitate from the solution by keeping it at the boiling point for a few hours in a hot water bath or by letting the particles settle for at least 48 hours in an alkali-resistant container (waxlined or polyethylene) protected from atmospheric carbon dioxide with a atmospheric carbon dioxide with a soda lime tube. Use the supernate for the preparation of the dilute solutions listed in Table 1.2.

Alternatively, prepare the dilute solutions by dissolving the weight of solid NaOH indicated in Table 1.2 in carbon dioxide free distilled water and diluting to 1000 mL.

Table 2.2 *Preparation of uniform sodium hydroxide solutions*

Normality of NaOH Solution	*Required Weight of NaOH to Prepare 1000 mL of solution g*	*Required Volume of 15 N NaOH to Prepare 1000 mL of Solution mL*
6	240	400
1	40	67
0.1	4	6.7

1,000,000 lb of water. It makes no difference what units are used as long as both weights are expressed in the same unit.

When elements are present in minute or trace quantities, the use of parts per million results in small decimal values. Therefore, it is more convenient to use parts per billion (ppb) in these cases. One part per billion is equal to one-thousandth of one part per million (0.001 ppm). For example, in studies of steam purity using a specific ion electrode to measure sodium content, values as low as 0.001 ppm are not uncommon. This is more conveniently reported as 1.0 ppb.

In recent times, the convention for reporting analytical results has been shifting toward the use of milligrams per liter (mg/L) as a replacement for parts per million and micrograms per liter (µg/L) as a replacement for parts per billion.

Test procedures and calculations of results are based on the milliliter (mL) rather than the more common cubic centimeter (cc or cm^3). The distinction between the two terms is very slight. By definition, a milliliter is the volume occupied by 1 g of water at 4°C, whereas a cubic centimeter is the volume enclosed within a cube 1 cm on each edge (1 mL = 1.000028 cm^3).

Miligram Per Liter (mg/L)

The milligrams per liter (mg/L) convention is closely related to parts per million (ppm). This relationship is given by:

$$\text{ppm} \times \text{solution density} = \text{mg/L}$$

Thus, if the solution density is close or equal to 1, then ppm = mg/L. This is normally the case in dilute, aqueous solutions of the type typically found in industrial water systems. Control testing is usually conducted without measurement of a solution's density. For common water samples, this poses no great inaccuracy, because the density of the sample is approximately 1 milligram per liter (mg/L) and parts per million (ppm) begin to diverge as the solution density varies from 1. Examples of this are a dense sludge from a clarifier underflow (density greater than 1) or closed cooling system water with high concentrations of organic compounds (density less than 1).

In water and wastewater analysis, the physical and chemical constituents of the sample can be expressed in the following units.

Milligrams Per Liter: mg/L

Milligrams per liter (mg/L) is a weight-volume relationship, i.e., the weight of physical or chemical constituent in unit volume of water or wastewater sample. This offers a convenient basis for a calculation of suspended and dissolved matter in the sample. Furthermore, mg/L is directly applicable to the metric system.

$$\text{mg/L} = \text{g/m}^3 \text{ and g/L } \times 10^3 = \text{kg/m}^3$$

Store the NaOH solutions in polyethylene (rigid, heavy type) bottles with polyethylene screw caps, paraffin-coated bottles with rubber or neoprene stoppers, or pyres bottles with rubber or neoprene stoppers. Check the solutions periodically. Protect them by attaching a tube of carbon dioxide absorbing granular material such as soda lime. Ascarite,* Caroxite, t or equivalent. Use at least 70 cm of rubber tubing to minimize vapor diffusion from the bottle. Replace the absorption tube before it becomes exhausted. Withdraw the solution by a siphon to avoid opening the bottle.

(ii) *Ammonium hydroxide solutions:* NH_4OH: Prepare 5 N, 3 N, and 0.2 N ammonium hydroxide solutions by diluting 333 mL, 200 mL, and 13 mL, respectively, of the concentrated reagent (sp gr 0.90, 20.0%, 15 N) to 1000 mL with distilled water.

2.3 BUFFER SOLUTIONS

A buffer solution is one which resists change in pH when small quantities of an acid or an alkali are added to it.

Acidic Buffer Solutions

An acidic buffer solution is simply one which has a pH less than 7. Acidic buffer solutions are commonly made from a weak acid and one of its salts – often a sodium salt.

A common example would be a mixture of ethanoic acid and sodium ethanoate in solution. In this case, if the solution contained equal molar concentrations of both the acid and the salt, it would have a pH of 4.76. It wouldn't matter what the concentrations were, as long as they were the same.

You can change the pH of the buffer solution by changing the ratio of the acid to salt, or by choosing a different acid and one of its salts.

Alkaline Buffer Solutions

An alkaline buffer solution has a pH greater than 7. Alkaline buffer solutions are commonly made from a weak base and one of its salts.

A frequently used example is a mixture of ammonia solution and ammonium chloride solution. If these were mixed in equal molar proportions, the solution would have a pH of 9.25. Again, it doesn't matter what concentrations you choose as long as they are the same.

2.4 EXPRESSION OF ANALYTICAL RESULTS

Parts Per Million (PPM)

Water analysis involves the detection of minute amounts of a variety of substances. The expression of results in percentage would require the use of cumbersome figures. For this reason, the results of a water analysis are usually expressed in parts per million (ppm) instead of percentage. One part per million equals one ten-thousandth of one percent (0.0001%), or one part (by weight) in a million parts, for example, 1 oz in 1,000,000 oz of water, or 1 lb in

1 cubic meter (m^3)	=	1,000,000 cubic centimeter (cc) or cm^3
1 cc	=	1 mL
1 m^3	=	1000 L

Conversions

- Concentration of a solution as weight percent (w/w) = $\frac{\text{Mass of solute (g)}}{\text{100 g solution}} \times 100\%$
- Concentration of a solution as volume percent (v/v) = $\frac{\text{mL substance}}{\text{100 mL solution}} \times 100\%$
- Concentration of a solution as weight per volume percent (v/v)
  $= \frac{\text{Mass of substance (g)}}{\text{100 mL solution}} \times 100\%$
- Concentration as part million (ppm) = mg substance/L solution: that is, ppm = mg/L or µg/mL or ng/µL; part per billion (ppb) = µg/L or ng/mL
- 1 Atmosphere (atm) = 760 torr
  = 760 mm mercury
  = 14.6 pounds per square inch (psi)
  = 101.306 kPa
- $^oF = (1.8 \times {}^oC) + 32$
- $^oC = \frac{(^oF - 32)}{1.8}$
- K (Kelvin) = °C + 273
- Density = mass (The units commonly used for density are g/mL, for volume liquids, gm/cm^3 for solids and g/L for gases).
- 1 Mole (mol) = molecular or formula weight in grams (i.e., 1 mol NaOH = 40 g NaOH)
- Molarity (M) = $\frac{\text{Mol of substances}}{\text{Liter solution}}$
- Normality (N) = $\frac{\text{Gram equivalent weight of substance}}{\text{Liter solution}}$ or Milligram equivalent/mL, solution
- STP = Standard Temperature and Pressure, which is 0°C and 1 atm
- NTP = Normal Temperature and Pressure, which is 25°C and in atm (At STP volume of 1 mol of any gas = 22.4 L, while at NTP 1 mol of any gas would occupy 24.45 L)
- Ideal gas equation: $PV = nRT$

Sometimes the concentration of pollutants is expressed in lb/million gallon. It is related to mg/L by the following expression:

$$\text{mg/L} \times 8.34 = \text{lb/million gallon}$$

Parts Per Million: ppm

Parts per million is a weight to weight ratio. This is mainly used to measure the gaseous pollutants in air and water.

Relation Between mg/L and ppm

The units mg/L and ppm are related by the following:

$$\text{ppm} = \frac{\text{mg/L}}{\text{Specific gravity of sample}}$$

In the case of water, specific gravity is 1, therefore

$$\text{ppm} = \text{mg/L (in water)}$$

The wastewater collected from domestic resources has specific gravity 1, therefore

$$\text{ppm} = \text{mg/L, (in domestic wastewater)}$$

This relation is not true for industrial wastewater because it differs from domestic wastewater in specific gravity depending on the process and the material produced.

2.5 UNIT CONVERSION AND ABBREVIATIONS

Metric and English units

1 liter (L)	=	1000 milliliter (mL)
1 gallon (gal)	=	3.784 L
1 L	=	0.264 gal
1 quart (qrt)	=	0.9464 L = 946.4 mL
1 L	=	1.057 qt
1 fluid ounce	=	29.6 mL
1 mL	=	1000 microliters (μL)
1 kg	=	1000 grams (g)
1 g	=	1000 milligrams (mg)
1 mg	=	1000 micrograms (μg)
1 μg	=	1000 nanograms (ng)
1 ng	=	1000 picogram (pg)
1 kg	=	2.205 pounds (lb)
1 lb	=	453.6 g
1 ounce (oz)	=	28.35 g

4.7	22.09
5.0	25.00
30.5	155.45

$$\Sigma\, x = 30.5;\ \Sigma\, x^2 = 155.45$$

$$(\Sigma\, x)^2 = (30.5)^5 = 930.25;\ n = 6$$

$$x = \sqrt{\frac{155.45 - \dfrac{930.25}{6}}{6-1}} = \sqrt{\frac{0.41}{5}} = 0.29 \text{ mg}$$