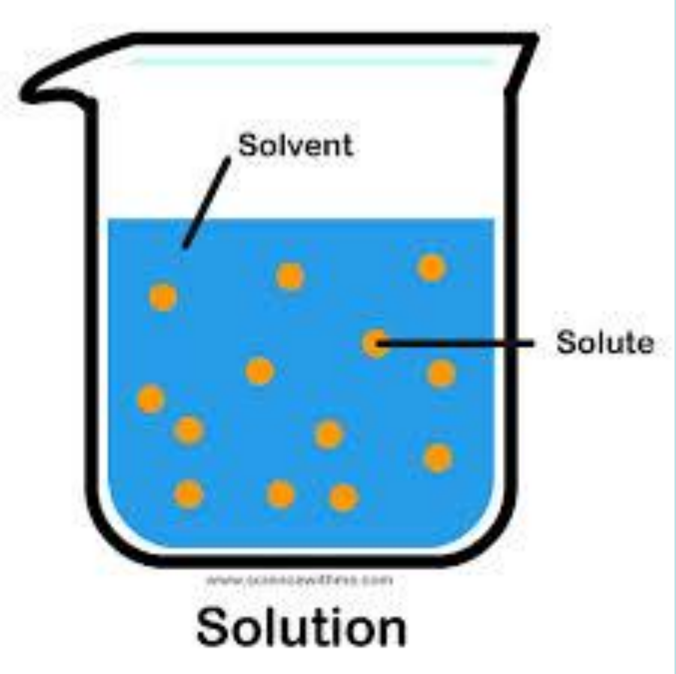
Solutions

A solution is a homogeneous mixture of at least two substances (one is called a solute and the other is called a solvent)

Solutes and solvents can be gases, liquids or solids

eg: NaCl solution , Sugar solution

$$NaCl + H_2O \rightarrow Na^+ + Cl^-$$



Molarity = number of moles / Volume in Litre

Number of moles = mass (gm) / molar mass

Example:

Prepare 1 liter of 1.00 M NaCl solution.

First calculate the molar mass of NaCl which is the mass of a mole of Na plus the mass of a mole of Cl or 22.99 + 35.45 = 58.44 g/mol

Molarity = number of moles/ volume (L)
Number of moles = molarity X Volume = 1X1 =1
Number of moles= mass/ molar mass

Mass = molar mass X number of moles
Mass = 58.44 X 1 = 58.44 gm

Weigh out 58.44 g NaCl.

Place the NaCl in a 1 liter volumetric flask

Add a small volume of distilled, deionized water to dissolve the salt.
Fill the flask upto the 1 L line.

Normality = Molarity * n
n is number of H+ or OH-

eg: for 1M HCl solution the normality of this solution is equal to 1

HCl → H+ + Cl-

For 1M solution of H2SO4 the normality of this solution is equal to 2
$H_2SO_4 \rightarrow 2H^+ + SO_4^{-2}$

For 3N solution of H2SO4, the Molarity of this solution is equal to ?

Normality = Molarity * n
in this case n is equal to 2

3 = M * 2 **M= 3/2 = 1.5**

For 6M solution of $Al(OH)_3$, the Normality of this solution is equal to ?

Normality = Molarity * n
in this case n is equal to 3

N = 6 * 3 = 18

Dilution Formula

C1 V1 = C2 V2

C1 is the initial concentration
V1 is the initial volume

C2 is the final concentration
V2 is the final volume

Example:
A laboratory procedure calls for 250 ml of an approximately 0.10 M solution of NH_3. Describe how you would prepare this solution using a stock solution of concentrated NH_3 (14.8 M).

Substituting known volumes in equation

C1 V1 = C2 V2

l 14.8 M × V_1 = 0.10 M × 250 mL

and solving for V_1 gives 1.7 mL.

Since we are making a solution that is approximately 0.10 M NH_3 we can use a micropippette to measure the 1.7 mL of concentrated NH_3, transfer the NH_3 to a volumetric flask (250 ml), and add sufficient water to give a total volume of approximately 250 mL.

BUFFERS

A buffer is a solution that resist the change in pH, a buffer can be a weak acid and its conjugate base (or a weak base with its conjugate acid). The weak acid is able to donate H^+ ions to neutralize incoming basic ions while the conjugate base is able to accept H^+.

Practically a buffer is prepared from mixing weak acid and its strong salt, example:

$$CH_3COOH \rightleftarrows CH_3COO^- + H^+$$

Acetic acid (weak acid) — Acetate ion — Hydrogen ion

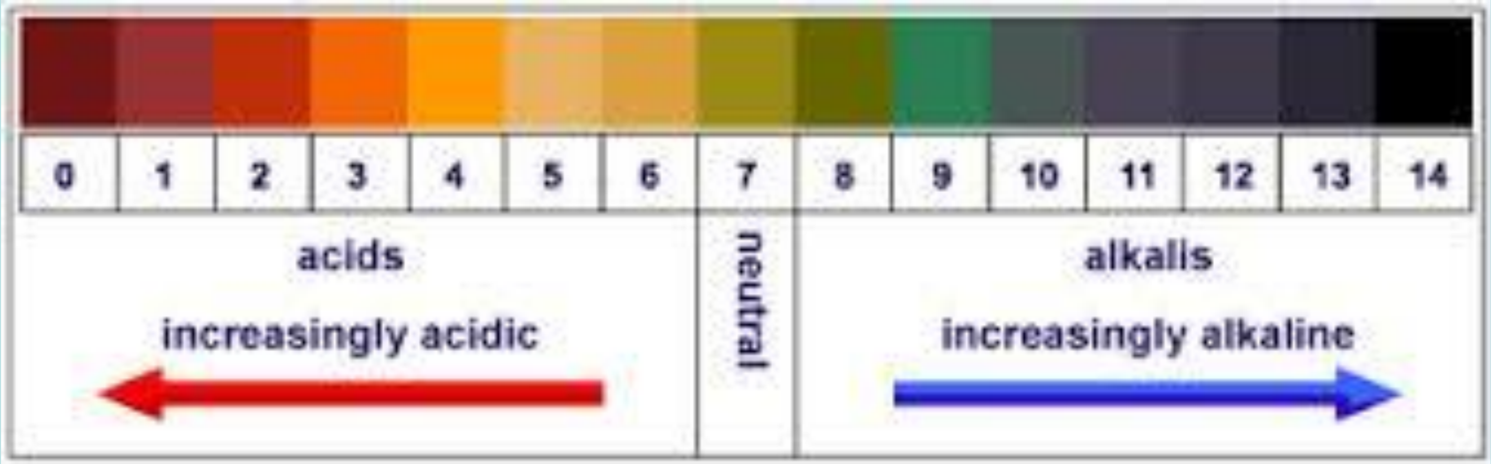
$$CH_3COONa \rightarrow CH_3COO^- + Na^+$$

Sodium acetate (salt) — Acetate ion — Sodium ion

The pH scale is a logarithmic scale representing the concentration of H^+ ions in a solution which equal the negative log of the H^+ ions concentration.

$pH = -\log [H^+]$

$[H^+] = 10^{-7}$ mol/L neutral solution pH = 7
$[H^+] > 10^{-7}$ mol/L acidic solution pH < 7
$[H^+] < 10^{-7}$ mol/L basic solution pH > 7



The pH of a buffer solution can be determined using the Henderson-Hasselbalch equation.

pH = p*Ka* + log [A-]/[HA]

Ka : dissociation constant for the weak acid

pKa (- log Ka) for weak acid.

[HA] = concentration of the buffer weak acid

[A-] = concentration of conjugate base for the buffer weak acid.

pH test strips

pH meter

Type of buffers

1-Synthetic: Can be prepared in the lab.

2.Physiological buffers (natural) :

The main physiological buffers are the bicarbonate, proteins (example haemoglobin), and the phosphate buffers .

Bicarbonate buffer

This is the most important buffer in blood . It is made from equilibrium between carbonic acid and its conjugate base bicarbonate.

$$\underset{\text{Carbonic acid}}{H_2CO_3} \leftrightarrow \underset{\text{bicarbonate ion}}{HCO_3^-} + H^+$$

Phosphate buffer

The phosphate buffer consists of dihydrogen phosphate ion (H2PO$_4^-$) in equilibrium with monohydrogen phosphate ion (HPO$_4^{2-}$) and H+

$$\underset{\text{(weak acid)}}{H_2PO_4^-} \leftrightarrow H^+ + \underset{\text{(conjugate base)}}{HPO_4^{2-}}$$

The weak acid, H2PO$_4^-$, and its conjugate base, HPO$_4^{2-}$, are in equilibrium