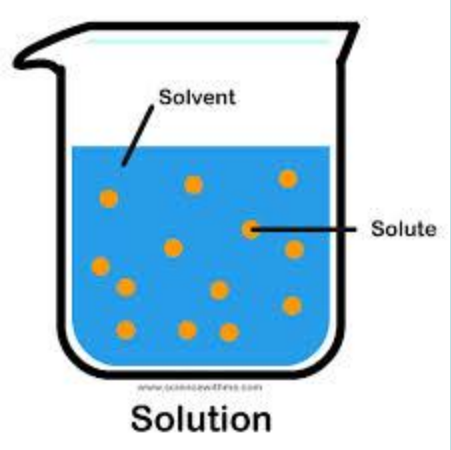
Solutions

A solution is a homogeneous mixture of more than two substances (solute(s) and solvent(s))



Molarity = number of moles / Volume in Liter

Number of moles = mass (gm)/ molar mass

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

eg: for 1M HCl solution the normality of this solution is equal to 1
For 1M solution of H_2SO_4 the normality of this solution is equal to 2

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

For 3N solution of H_2SO_4, 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

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.

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

$14.8 \text{ M} \times V_1 = 0.10 \text{ M} \times 0.25 \text{ L}$

and solving for V_1 gives 1.69×10^{-3} liters, or 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

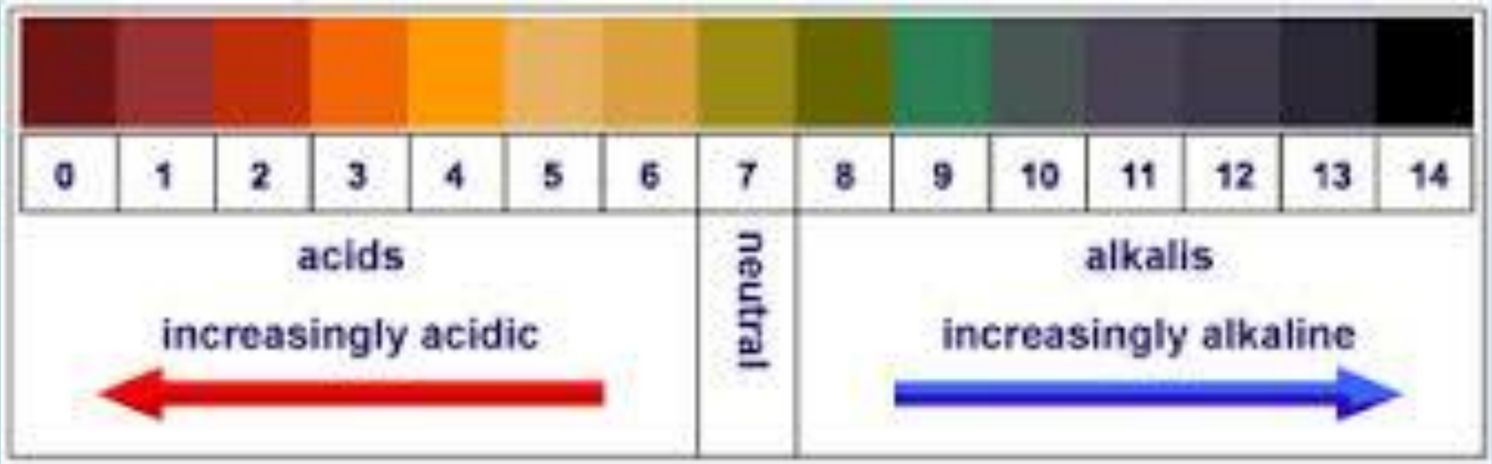
$$CH_3COONa \rightarrow CH_3COO^- + Na^+$$

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 buffer solution can be determined using the Henderson-Hasselbalch equation.

$pH = pK_a + \log [A^-]/[HA]$

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.

$$H_2CO_3 \leftrightarrow HCO_3^- + H^+$$

Carbonic acid   bicarbonate ion

Phosphate buffer

The phosphate buffer consists of dihydrogen phosphate ion ($H2PO_4^-$) in equilibrium with monohydrogen phosphate ion (HPO_4^{2-}) and H+

$$H_2PO_4^- \leftrightarrow H^+ + HPO_4^{2-}$$

(weak acid)   (conjugate base)

The weak acid, $H2PO_4^-$, and its conjugate base, HPO_4^{2-} , are in equilibrium