SOLUTIONS AND BUFFERS

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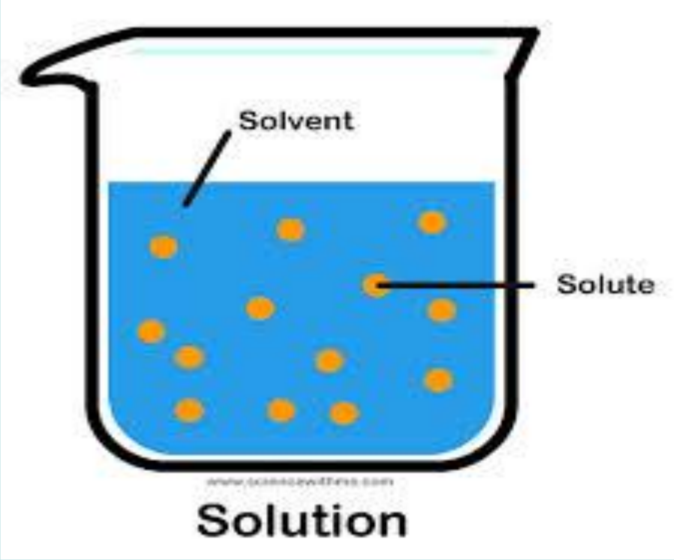
Solutions

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

1- Liquid solutions

Liquids dissolve gases, other liquids, and solids.
An example of a dissolved gas is oxygen in water, which allows fish to breathe under water.



An examples of a dissolved liquid is ethanol in water, as found in alcoholic beverages.

An example of a dissolved solid is salt in water

eg: NaCl solution

$$NaCl + H2O \rightarrow Na^{+} + Cl^{-}$$

2- Gaseous mixtures

If the solvent is a gas, gases (non-condensable) or vapors (condensable) are dissolved under a given set of conditions.
An example of a gaseous solution is air (oxygen and other gases dissolved in nitrogen).

3-Solid solutions

If the solvent is a solid, then gases, liquids, and solids can be dissolved.
Gas in solids:
Hydrogen dissolves rather well in metals, especially in palladium
Liquid in solid:
Mercury in gold, forming an amalgam
Solid in solid:
Radium sulfate dissolved in barium sulfate

Concentration Expressions:

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

$HCl \rightarrow H^+ + Cl^-$

For 1M solution of H_2SO_4 the normality of this solution is equal to 2

$H_2SO_4 \rightarrow 2H^+ + SO_4^{-2}$

For 3N solution of H_2SO_4, the Molarity of this solution is equal to ?

$H_2SO_4 \rightarrow 2H^+ + SO_4^{-2}$

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 ?

$Al(OH)_3 \rightarrow 3OH^- + Al^{+3}$

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

Normality = 6 * 3 = 18

Dilution Formula

$C_1 V_1 = C_2 V_2$

C_1 is the initial concentration
V_1 is the initial volume

C_2 is the final concentration
V_2 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

$$14.8\ \text{M} \times V_1 = 0.10\ \text{M} \times 250\ \text{mL}$$

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

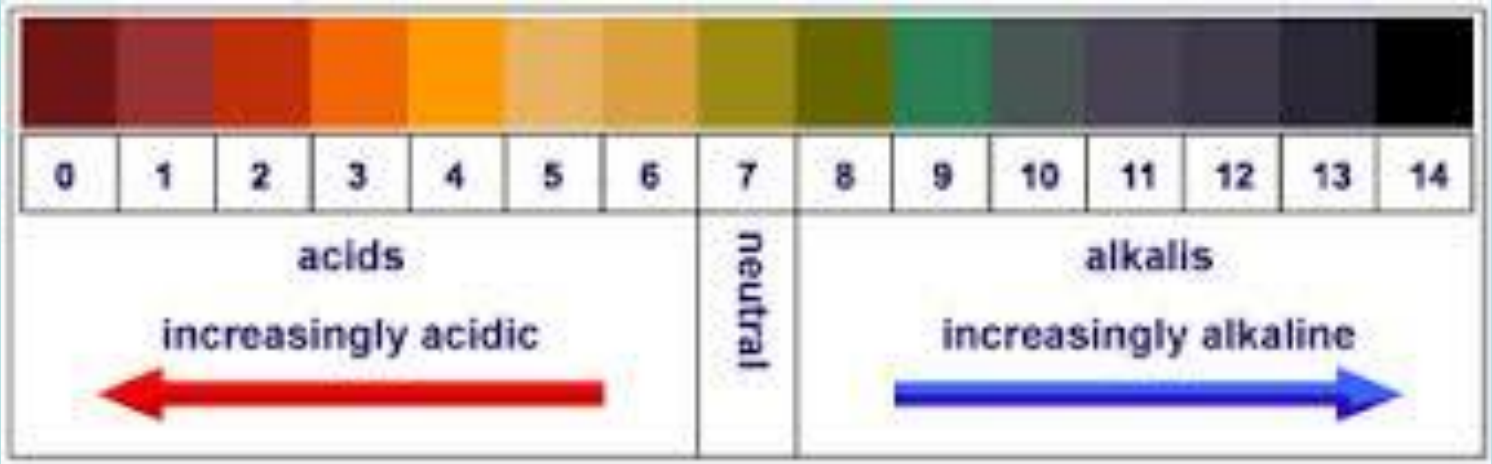
$CH_3COONa \rightarrow CH_3COO^- + Na^+$

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.

$HA \leftrightarrow H^{+} + A^{-}$

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

Types 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