

Analytical Chemistry

1. Electrolytes:

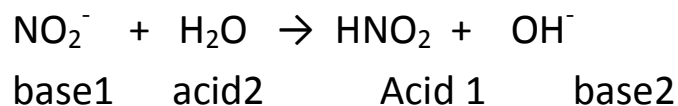
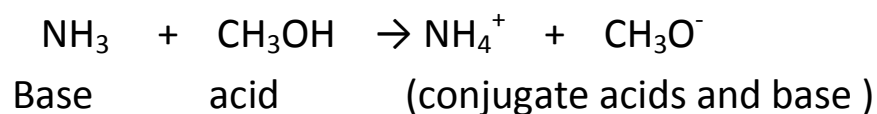
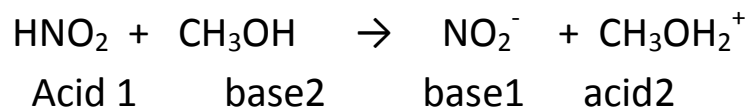
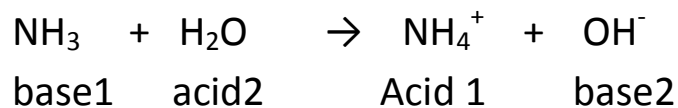
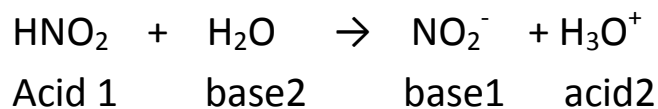
Electrolytes are solutes which ionize in a solvent to produce an electrically conducting medium. Strong electrolytes ionize completely whereas weak electrolytes are only partially in the solvent.

Table 1:

Strong electrolytes	Weak electrolytes
1.the inorganic acids $\text{HNO}_3, \text{HClO}_4, \text{H}_2\text{SO}_4, \text{HCl}, \text{HI},$ $\text{HBr}, \text{HClO}_3, \text{HBrO}_3.$	1.Many inorganic acid such as $\text{H}_2\text{CO}_3, \text{H}_3\text{PO}_4, \text{H}_3\text{BO}_3, \text{H}_2\text{S}$ $\text{H}_2\text{SO}_4.$
2.Alkali and alkaline-earth hydroxides	2.Most organic acid
3.Most Salts	3.Ammonia and most organic bases
	4.Halides, cyanides, and thiocyanates of Hg, Zn and Cd.

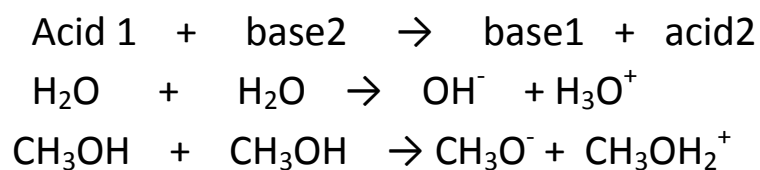
2. Acid and Bases:

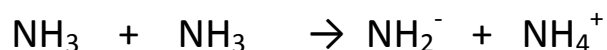
For analytical chemists, the most useful acid-base concept is that proposed independently by Bronsted and Lowry in 1923. According to the Bronsted-Lowry view, an acid is any substance that is capable of donating a proton; a base is any substance that can accept a proton.



3. Autoprotolysis:

Amphiprotic solvents undergo self-ionization or autoprotolysis to form a pair of ionic species. Autoprotolysis is an acid-base reaction, as illustrated by the following equation .

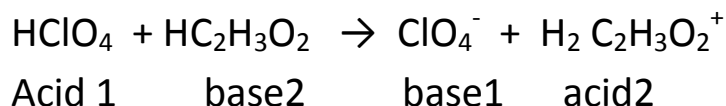




4. Strengths of acids and bases:

The extent of reaction between a solute acid (or base) and a solvent is critically dependent upon the tendency of the latter to donate or accept protons. Thus, for example, perchloric, hydrochloric, and hydrobromic acid are all classified as strong acids in water.

If anhydrous acetic acid, a poorer proton acceptor than water, is used as the solvent, none of the acids undergoes complete dissociation and remains a strong acid, instead, equilibria such as the following develop:

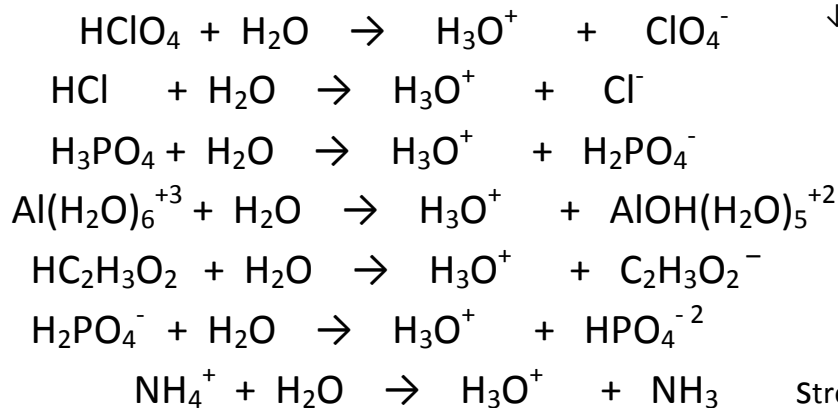


Strongest
Acid

weakest
base

↑

↓



weakest
acid

Strongest
base

The molecular formula may additionally provide structural information thus, C_2H_6O is both the empirical and the chemical formula for the chemically ferent ethanol , C_2H_5OH , and dimethyl ether CH_3OCH_3 .

5. The mole:

It is gram formula weight per formula weight (M.wt., which is the summation of the atomic weights in grams for all of the atoms in the atoms in the chemical formula for species.

$$\text{Mole} = \frac{\text{weight g}}{\text{M.Wt.g/mol}}$$

For example:

$$\text{M.wt. } CH_2O = 12*1 + 1*2 + 16*1 = 30 \text{ g/mol}$$

$$\text{M.wt. } C_6H_{12}O_6 = 12*6 + 1*12 + 16*6 = 180 \text{ G/mol}$$

$$\text{M.wt. } NaCl = 23*1 + 35*1 = 58 \text{ g/mol}$$

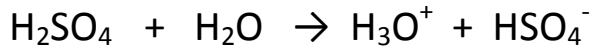
6. Molarity or molar concentration :

The molarity M, expresses the moles of a solute contained in one liter of solution.

$$M = \frac{\text{mole}}{\text{Liter}} \quad \text{or} \quad \frac{\text{mmole}}{\text{mL}}$$

Example:

Concentration of H_2SO_4 1M solution is 1 $\text{H}_2\text{SO}_4/\text{L}$.
determination the weight of H_2SO_4 .



$$\text{M.wt. H}_2\text{SO}_4 = 1*2 + 32*1 + 16*4 = 98 \text{ g/mol}$$

$$M = \frac{\text{mol}}{\text{L}} = \frac{\text{g/mol}}{\text{L}}$$

$$1\text{M} = \frac{\text{wt.}}{\text{L}} \\ \frac{98\text{g/mol}}{\text{L}} \rightarrow \text{wt.} = 98 \text{ g.}$$

$$\text{Not: } M = 0.012 = 1.2*10^{-2} \text{ M}$$

Example 1 :

Calculate the analytical and equilibrium molar concentration of the solute species in

a) An aqueous solution that contains 2.3g of ethanol, in 3.5L

$$M = \frac{\text{mol}}{\text{L}} = \frac{\text{g/mol}}{\text{L}}$$

$$\text{M.wt. C}_2\text{H}_5\text{OH} = 12*2 + 1*6 + 16*1 = 46\text{g/mol}$$

$$\text{Mole} = \frac{\text{wt. C}_2\text{H}_5\text{OH}}{\text{M.wt. C}_2\text{H}_5\text{OH}} = \frac{2.3\text{g}}{46\text{g/mol}} = 0.05 \text{ mol}$$

$$\text{M.wt. C}_2\text{H}_5\text{OH} \quad 46\text{g/mol}$$

$$M = 0.05 \text{ mol} = 0.0143 \text{ mol/L}$$

- b) An aqueous solution that contains 285mg of trichloroacetic acid Cl_3CCOOH in 10ml. (the acid is 73% ionized in water).

$$M_{\text{HA}} = \frac{\text{mmole}}{\text{mL}}$$

$$\text{mmol} = \frac{\text{mg}}{\text{mg/mmol}} = \frac{285\text{mg.}}{163\text{mg/mmol}} = 1.75\text{mmol.}$$

$$\text{M.wt.} = 35 \cdot 3 + 12 \cdot 2 + 16 \cdot 2 + 1 \cdot 1 = 163\text{g/mol or mg/mmol}$$

$$[\text{HA}] \text{ M} = \frac{1.75\text{mmol.}}{10 \text{ mL}} = 0.175 \text{ mmol./mL}$$



Weak acid base

$$0.175\text{M} \qquad 0 \qquad 0$$

$$0.175\text{M}-x \qquad x \qquad x$$

$$M_{[\text{HA}]_{\text{remain}}} = \frac{0.27 \cdot 285\text{mg.}/163\text{mg/mmol.}}{10 \text{ mL.}} \\ = 0.047\text{mmol./mL.}$$

$$M_{\text{H}_3\text{O}^+} = 0.175\text{M} - 0.047\text{M} = 0.128 \text{ mmol./mL.}$$

Example 2 :

Describe the preparation of 2L of 0.1M Na₂CO₃ from the solid

$$M = \frac{\text{mole}}{L} \rightarrow \text{mole} = M * L$$

$$\text{Mol. Na}_2\text{CO}_3 = 0.1 \frac{\text{mol}}{L} \text{Na}_2\text{CO}_3 * 2L = 0.2 \text{ mol.}$$

$$\text{Mol.} = \frac{\text{wt.}}{\text{M.wt.}} \rightarrow \text{wt.} = \text{mol.} * \text{M.wt.}$$

$$\text{M.wt. Na}_2\text{CO}_3 = 23*2 + 12*1 + 16*3 = 106 \text{ g/mol.}$$

$$\text{Wt. of Na}_2\text{CO}_3 = 0.2 \text{ mol.} * 106 \text{ g/mol} = 21.2\text{g}$$

Therefore, dissolve 21.2 g of the Na₂CO₃ in water and dilute to exactly 2.00 L.

Example 3 :

Describe the preparation of 2L of 0.1M Na⁺ from pure Na₂CO₃ .



$$\text{Mol. Na}_2\text{CO}_3 = 0.1 \text{ mol./L} * 2L * \frac{1 \text{ Na}_2\text{CO}_3}{2 \text{ mol.}} = 0.1 \text{ mol.}$$

Wt. $\text{Na}_2\text{CO}_3 = 0.1 \text{ mol. Na}_2\text{CO}_3 * 106\text{g/mol. Na}_2\text{CO}_3 = 10.6 \text{ g.}$

Dissolve 10.6 g. in water and dilute to 2.00L.

6. Normality or normal concentration. N

The specialized method for expressing concentration is based upon the number of equivalents of solute that are contained in a liter of solution .

$$N = \text{eq./L} \quad \text{or} \quad \text{meq./mL.}$$

$$\text{Equivalent (eq.)} = \text{g./eqw.} \quad \text{Or} \quad \text{meq.} = \text{g./meqw.}$$

7. Titer :

Titer defines concentration in terms of the weight of some species with which a unit volume of the solution reacts .

8. P- function :

$$PX = - \log [x]$$

Example 4 :

Calculate the p-value for each ion in a solution that is 2.00×10^{-3} M in NaCl and 5.4×10^{-4} M in HCl .



$$5.4 \times 10^{-4} \quad 0 \quad 0$$

$$0 \quad 5.4 \times 10^{-4} \quad 5.4 \times 10^{-4}$$



$$2 \times 10^{-3} \quad 0 \quad 0$$

$$0 \quad 2 \times 10^{-3} \quad 2 \times 10^{-3}$$

$$\begin{aligned} \text{PH} &= -\log [\text{H}^+] = -\log[5.4 \times 10^{-4}] = -\log 5.4 - \log 10^{-4} \\ &= -0.73 - (-4) = 3.27 \end{aligned}$$

To obtain pNa , we write

$$\begin{aligned} \text{pNa} &= -\log[2 \times 10^{-3}] = -\log 2 - \log 10^{-3} \\ &= -0.301 - (-3) = 2.699 \end{aligned}$$

The chloride ion concentration is given by the sum of the two concentration that is :

$$[\text{Cl}] = 2 \times 10^{-3} + 5.4 \times 10^{-4} = 2 \times 10^{-3} + 0.54 \times 10^{-3} = 2.54 \times 10^{-3}$$

$$\begin{aligned} \text{pCl} &= -\log 2.54 \times 10^{-3} = -\log 2.54 - \log 10^{-3} = -0.4048 - (-3) \\ &= 2.595 \end{aligned}$$

Example 5 :

Calculate the molar concentration of Ag^+ in a solution having a pAg of 6.372 .

$$\text{pAg} = -\log[\text{Ag}^+] = 6.372$$

$$\log[\text{Ag}^+] = -6.372 = -7.00 + 0.628$$

$$\begin{aligned} [\text{Ag}^+] &= \text{antilog}(-7) * \text{antilog}(0.628) = 10^{-7} * 4.246 \\ &= 4.25 * 10^{-7} \end{aligned}$$

Note: it is noteworthy that the p-value for a species becomes negative when its concentration is greater than unity. For example, in a 2M solution of HCl
 $[\text{H}^+] = 2 \rightarrow \text{pH} = -\log 2 = -0.30$

9. Density and specific gravity :

The density of a substance measures its mass per unit volume g/mL.

10. Part per million; parts per billion:

For very dilute solutions, it is convenient to express concentrations in terms of parts per million:

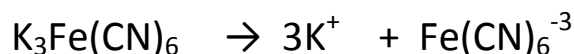
$$\text{Ppm} = \frac{\text{weight of solute}}{\text{Weight of solution}} * 10^6$$

If the solvent is water and the quantity of solute is so small that the density of the solution is still essentially 1g/ML

$$\text{Ppm} = \frac{\text{mg. solute}}{10^6 \text{mg. water}} = \frac{\text{mg. solute}}{\text{L solution}}$$

Example 6:

What is the molarity of K^+ in a solution that contains 63.3 ppm of $\text{K}_3\text{Fe}(\text{CN})_6$?



Because the solution is so dilute, it is safe to assume that the density of the solution is 1g/ML therefore.

$$63.3 \text{ ppm } \text{K}_3\text{Fe}(\text{CN})_6 = 63.3 \text{ mg } \text{K}_3\text{Fe}(\text{CN})_6 / \text{L}$$

$$[\text{K}^+] = \frac{63.3 \text{ mg } \text{K}_3\text{Fe}(\text{CN})_6}{\text{L}} \cdot \frac{3 \text{ mmol } \text{K}^+}{\text{mmol } \text{K}_3\text{Fe}(\text{CN})_6} \cdot \frac{\text{mmol } \text{K}_3\text{Fe}(\text{CN})_6 \cdot \text{mol}}{329 \text{ mg} \cdot 10^3 \text{ mmol}}$$
$$= 5.77 \cdot 10^{-4} \text{ M}$$

11. Percentage concentration :

It should be noted that the denominator in each of these expressions refers to the solution rather than to the solvent.

$$\text{Weight percent (w/w)} = \frac{\text{wt. of solute}}{\text{Wt. of soln}} \cdot 100$$

$$\text{Weight percent (v/v)} = \frac{\text{volume of solute}}{\text{volume of soln}} \cdot 100$$

Volume of soln

$$\text{Weight-volume percent (w/w)} = \frac{\text{wt. of solute, g}}{\text{Volume of soln, ml}} * 100$$

Example 7 :

Describe the preparation of 100mL of 6.0M HCl from the concentration reagent . the label on the bottle states that the reagent is 37 % HCl and has a specific gravity of 1.18

$$d = \text{g/ml}$$

$$\frac{\text{g HCL}}{\text{ml concd}} = \frac{1.18 \text{g concd soln} * 37 \text{ HCL}}{\text{ml concd} * 100 \text{g concd}}$$

$$= 0.437 \text{ g/ml}$$

$$\text{G HCL required} = 100 \text{ml} * \frac{6.0 \text{mmol HCL}}{\text{ml}} * \frac{0.0365 \text{g}}{\text{mmol HCL}}$$

$$= 21.9 \text{g}$$

$$\text{Vol.} = 21.9 \text{ g HCL} * \frac{\text{ml concd HCL}}{0.437 \text{g}} = 50.1 \text{ ml concd HCL}$$

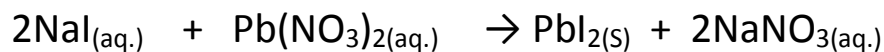
Dilute 50 ml of the concentrated reagent to a volume of about 100 mL.

12. Solution-diluent volume ratio:

the composition of dilute solution is sometimes specified in term of the volume of a more concentrated reagent and the volume of solvent to be used in diluting it . the volume of the former is separated from that of the latter by a colon . thus, a 1:4 HCL solution contains four volumes of water for each volume of concentrated hydrochloric acid.

13. Stoichiometric relationships:

A balanced chemical equation is a statement of the combining rations (in moles) that exist between reacting substances and their products . thus, the equation



Indicates that 2 mole of sodium iodide in aqueous solution combine with 1mole of lead nitrate to produce 1mole of solid lead iodide and 2 mole of aqueous sodium nitrate.