(2) Percent Concentration

** Chemists frequently express concentrations in terms of percent (parts per hundred).

** Three common methods are:

$$\text{weight percent (w/w)} = \frac{\text{weight solute}}{\text{weight solution}} \times 100\%$$

$$\text{volume percent (v/v)} = \frac{\text{volume solute}}{\text{volume solution}} \times 100\%$$

$$\text{weight/volume percent (w/v)} = \frac{\text{weight solute, g}}{\text{volume solution, mL}} \times 100\%$$

PERCENT BY MASS

Example
What is the percent by mass of a solution that contains 26.5 g of glucose in 500 g of solution?

Solution

%w/w= (26.5/500) x100 = 53%

PERCENT BY VOLUME

Example
How would you prepare 250 mL of 70 % (v/v) of rubbing alcohol?
Solution

%v/v= (v/v) x100

70= (v/250)x 100
100v= 70x250

V= 70x250/100= 175 ml

PERCENT BY MASS OVER VOLUME (m/v)

Example

If the density of 500 g solution is 0.857 g/mL, what is the percent (m/v) of 275 g rubbing alcohol?

Solution:
D= w/v
0.857= 500/v
V= 500/0.857= 583.4 ml of solution
%w/v= (275/583.4)x100 = 47.13%

Normality concentration

As per the standard definition, normality is described as the number of gram or mole equivalents of solute present in one liter of a solution.

- Normality (N)= Number of gram equivalents × [volume of solution in liters]$^{-1}$

How can calculate the gram equivalents?

1- gram equivalent of acid

Eq. wt= Molecular weight of acid / no. of H^+

Ex. Eq. wt of HCl

HCl------------ $H^+ + Cl^-$ H_3PO_4------------- $3H^+ + PO_4^{-3}$

Eq. wt of HCl = 36.5/1= 36.5 CH_3COOH ------------ $H^+ + CH_3COO^-$

Ex. Eq. wt of H_2SO_4

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

Eq. wt of H_2SO_4 = 98 /2= 49

2- gram equivalent of base

Eq. wt= Molecular weight of base / no. of OH^-

Ex. Eq. wt of NaOH

NaOH ------------ $Na^+ + OH^-$

Eq. wt of NaOH = 40/1= 40

Ex. Eq. wt of $Ca(OH)_2$

$Ca(OH)_2$ ------------ $Ca^{+2} + 2OH^-$

Eq. wt of $Ca(OH)_2$ = 74/2= 37 $Al(OH)_3$-------$Al^{+3} + 3OH^-$

2- gram equivalent of salt

Eq. wt= Molecular weight of salt / no. of metal x oxidation no. of metal

Ex. Eq. wt of NaCl

Eq. wt of NaCl = 58.5/1x1= 58.5

Ex. Eq. wt of Na_2CO_3 $CaCO_3$

Eq. wt of Na_2CO_3= 106/2x1= 53

Example 1.

Calculate the normality of 0.321 g sodium carbonate when it is mixed in a 250 mL solution.

Solution

N= no. of eq.mol / v of solution in litre

= (w/eq wt) x (1000/v in ml) eq wt of Na_2CO_3 = 106/2x1=53

N= (0.321/ 53) x (1000/250)

N= 0.024 N Na_2CO_3

M=

Exampe 2. What is the normality of the following?

- 0.1381 M NaOH sol. N= 0.1381N
- 0.0521 M H_3PO_4 sol. Eq. wt. of H_3PO_4 = 1/3 Mwt
- N= 3x M
- N=3x 0.0521= 0.1563 N

0.015 M Ca CO_3 Eq. wt of $CaCO_3$= 100/1x2= 50

N= 2x 0.015= 0.03N

Acids and Bases

Concepts

A compound is classified as an acid or a base based on certain properties. At present there are several theories which define the concepts of acidity and basicity. Some important concepts are detailed below:

Arrhenius concept of acids and bases:

The Arrhenius definition of acids says that they are compounds that give off H^+ ions in water and that bases are compounds that give off OH^- ions in water. Thus, according to this theory only protic acids are allowed and only hydroxide bases are allowed to be classified as an acid or a base.

H_2SO_4 -------------- H^+ (aq) + HSO_4^- (aq)

HSO_4^- ---------- H^+ + SO_4^{-2}

Bronsted-Lowry definition:

The Bronsted-Lowry definition of acids is that acids are compounds that give off H^+ ions when they react with another compound. Likewise, this definition says that bases are compounds that accept H^+ ions from other compounds. Furthermore, it also brings a new concept of conjugate acids and conjugate bases. Since an acid after donating its proton is technically a base according to this definition and is referred to as a conjugate acid. So every acid has its conjugate base and vice versa. Also, the stronger is an acid, the weaker is its conjugate base and vice versa.

H_2SO_4 + H_2O --------------- H_3O^+ + HSO_4^-

Acid **Base** **Conjugate acid** **Conjugate base**

Lewis concept of acids and bases:

Acids are electron pair acceptors while bases are electron pair donors. Thus, electron deficient species like BF_3 are Lewis acids while electron rich species such as tertiary amines are Lewis bases. Lewis acids may combine with Lewis bases to generate a salt.

$(CH_3)_3N$: + BF_3 ------------ $(CH_3)_3N\ BF_3$

Lewis base Lewis acid

pH and pOH

Hydronium and hydroxide ions are present both in pure water and in all aqueous solutions, and their concentrations are inversely proportional as determined by the ion product of water (Kw). The concentrations of these ions in a solution are often critical determinants of the solution's properties and the chemical behaviors of its other solutes . A solution is neutral if it contains equal concentrations of hydronium and hydroxide ions; acidic if it contains a greater concentration of hydronium ions than hydroxide ions; and basic if it contains a lesser concentration of hydronium.

A common means of expressing quantities, the values of which may span many orders of magnitude, is to use a logarithmic scale. One such scale that is very popular for chemical concentrations and equilibrium constants is based on the p-function, defined as shown where “X” is the quantity of interest and “log” is the base-10 logarithm.

$$pX = -\log X$$

The pH of a solution is therefore defined as shown here, where $[H_3O^+]$ is the molar concentration of hydronium ion in the solution:

$$pH = -\log[H_3O^+]$$

Rearranging this equation to isolate the hydronium ion molarity yields the equivalent expression:

$$[H_3O^+] = 10^{-pH}$$

Likewise, the hydroxide ion molarity may be expressed as a p-function, or **pOH**:

$$pOH = -\log[OH^-]$$

or

$$[OH^-] = 10^{-pOH}$$

Finally, the relation between these two ion concentration expressed as p-functions is easily derived from the K_w expression:

$$K_w = [H_3O^+][OH^-]$$

$$-\log K_w = -\log([H_3O^+][OH^-]) = -\log[H_3O^+] + -\log[OH^-]$$

$$pK_w = pH + pOH$$

At 25 °C, the value of K_w is 1.0 × 10−14, and so:

$$14.00 = pH + pOH$$

As we learned earlier, the hydronium ion molarity in pure water (or any neutral solution) is $1.0 \times 10^{-7} M$ at 25 °C. The pH and pOH of a neutral solution at this temperature are therefore:

$$pH = -\log[H_3O^+] = -\log(1.0 \times 10^{-7}) = 7.00$$

$$pOH = -\log[OH^-] = -\log(1.0 \times 10^{-7}) = 7.00$$

Summary of Relations for Acidic, Basic and Neutral Solutions

Classification	Relative Ion Concentrations
acidic	$[H_3O^+] > [OH^-]$
neutral	$[H_3O^+] = [OH^-]$
basic	$[H_3O^+] < [OH^-]$

$[H_3O^+]$ (M)	$[OH^-]$ (M)	pH	pOH	Sample Solution
10^{1}	10^{-15}	–1	15	
10^{0} or 1	10^{-14}	0	14	1 M HCl; acidic
10^{-1}	10^{-13}	1	13	gastric juice
10^{-2}	10^{-12}	2	12	lime juice; 1 M CH_3CO_2H (vinegar)
10^{-3}	10^{-11}	3	11	stomach acid
10^{-4}	10^{-10}	4	10	wine; orange juice
10^{-5}	10^{-9}	5	9	coffee
10^{-6}	10^{-8}	6	8	rain water
10^{-7}	10^{-7}	7	7	pure water; neutral
10^{-8}	10^{-6}	8	6	blood; ocean water
10^{-9}	10^{-5}	9	5	baking soda
10^{-10}	10^{-4}	10	4	
10^{-11}	10^{-3}	11	3	Milk of Magnesia
10^{-12}	10^{-2}	12	2	household ammonia, NH_3
10^{-13}	10^{-1}	13	1	bleach
10^{-14}	10^{0} or 1	14	0	1 M NaOH; basic
10^{-15}	10^{1}	15	–1	

EXAMPLE 1: CALCULATION OF PH

1- What is the pH of stomach acid, a solution of HCl with a hydronium ion concentration of 0.1 *M*?

Solution:

pH= -log $[H^+]$ = -log 0.1 = -log 10^{-1} = 1

pOH = 14-1 = 13

2- What is the pH of a solution of H_2SO_4 with a hydronium ion concentration of 1.2×10^{-3} *M*?

Solution:

$pH = -\log[H_3O^+]$
$= -\log(1.2 \times 10^{-3})$
$= -(-2.92) = 2.92$

$pOH = 14-2.29= 11.71$

3- Calculate the hydronium ion concentration of blood, the pH of which is 7.3 (slightly alkaline).

Solution:

$pH = -\log[H_3O^+]$

$[H_3O^+] = 10^{-pH}$

$[H_3O^+]= 5.01x\ 10^{-8}$

CALCULATION OF POH

1- What are the pOH and the pH of a 0.0125-*M* solution of potassium hydroxide, KOH?

Solution:

KOH --------- K^+ + OH^-

$pOH = -\log [OH^-]$

$pOH= -\log 0.0125 = 1.9$

$pH= 14- 1.9 = 12.1$

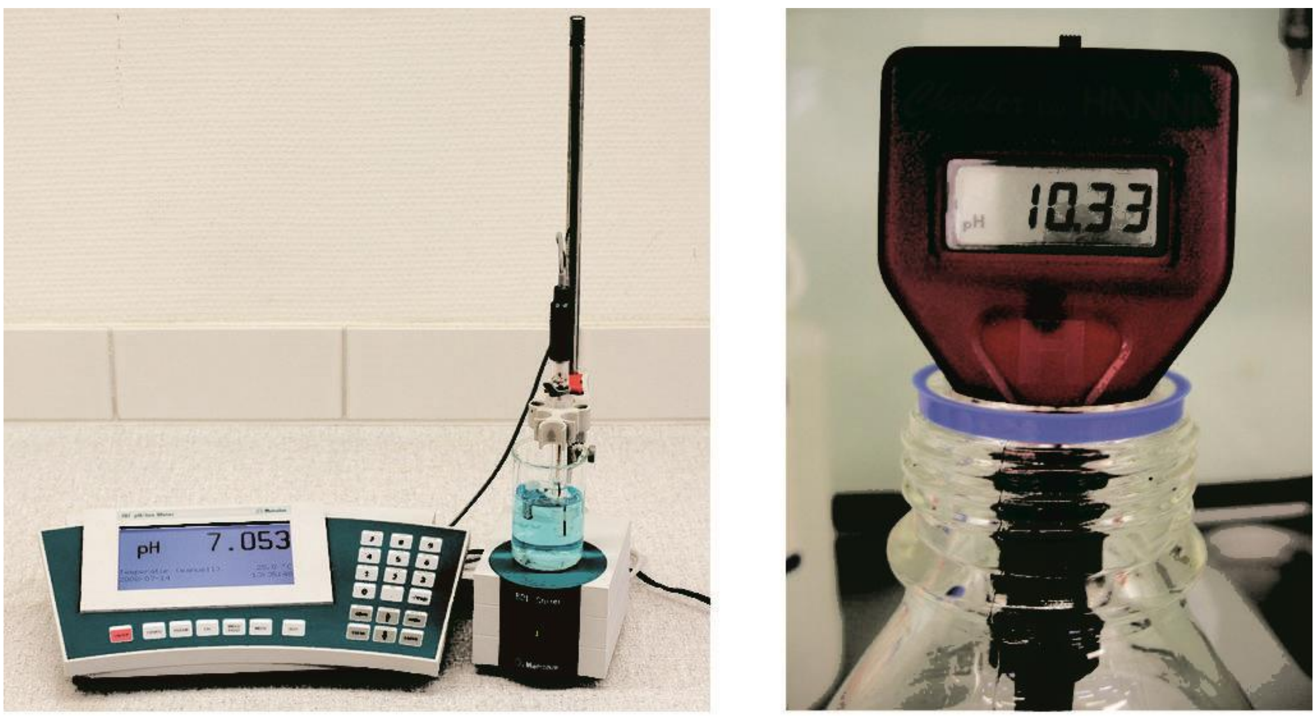
2- The hydronium ion concentration of vinegar is approximately 4 × 10^{-3}*M*.

What are the corresponding values of pOH and pH?

Solution:

CH_3COOH -------------- CH3COO$^-$ + H^+

pH = −log[H_3O^+] = -log **4 × 10^{-3} = 2.39**
pOH= 14- 2.39 = 11.61



(a) (b)