ALMUSTAQBAL UNIVERSITY

College of Health and Medical Techniques

Medical Laboratories Techniques Department

Stage : First year students

Subject : Lecture 8A

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Calculations of $[H_3O^+]$, pH , $[OH^-]$ and pOH for strong Acids and Bases

A solution is acidic if $[H_3O^+] > [OH^-]$. and is basic if $[H_3O^+] < [OH^-]$.

Strong acids are acids that completely dissociate in water.

Strong acids, such as HNO_3, almost completely dissociated

$$HNO_3\,(aq) + H_2O\,(l) \rightarrow H_3O^+(aq) + NO_3^-(aq)$$

0.1 M   0.1 M

The hydronium ion $[H_3O^+]$ is the acidic species in solution, and its concentration determines the acidity of the resulting solution

pH of a strong acids:

When a solution of 0.1 M HNO_3 dissolves in water it dissociates completely to its ions (i.e : 0.1M $[H_3O^+]$) .

$[H_3O^+] = C$ where C is the initial concentration of the strong acid

Example :

Calculate the pH of a 0.1 M solution of HCl.

$HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$
0.1 M 0.1 M

$[H_3O^+]$ = C = The original concentration of the strong acid [HCl] = 0.1 M

pH = - log $[H_3O^+]$ = - log (0.1) = 1

Example:

Calculate the pH of the following strong acid solutions:

(a) $1.3x10^{-2}$ M $HClO_4$, (b) 1.3×10^{-3} M HCl, (c) $1.3x10^{-4}$ M HNO_3 .

Solution:

a) $HClO_4 + H_2O \rightarrow H_3O^+ + ClO_4^-$

$1.3x10^{-2}$ M $1.3x10^{-2}$ M

pH = - log [H_3O^+] = - log 1. 3 × 10^{-2} = 1.89

(b) $HCl + H_2O \rightarrow H_3O^+ + Cl^-$

$1.3x10^{-3}$ M $1.3x\ 10^{-3}$ M

pH = - log [H_3O^+] = - log 1.3 x 10^{-3} = 2.89

c) $HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$

1.3×10^{-4} M 1.3×10^{-4} M

pH = - log [H_3O^+] = - log [1.3× 10^{-4}] = 3.89

Example:

Calculate the pOH and pH of the following strong base solutions:

(a) 0.05 M NaOH, (b) 0.05 M $Ca(OH)_2$, (c) 0.05 M $La(OH)_3$.

solution:

a) NaOH $\rightarrow$ Na^+ + OH^-

0.05 M 0.05 M

pOH = - log [OH^-] = - log 5× 10^{-2} = 1.3

As pH + pOH = 14

pH = 14 – 1.3 = 12.7

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

0.05 M 2(0.05) = 0.1 M

pOH = - log [OH^-] = - log 0.1 = 1.0

pH = 14 – 1.0 = 13.0

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

0.05 M 3(0.05) = 0.15 M

pOH = - log [OH^-] = - log 0.15 = 0.82

pH = 14 – 0.82 = 13.18

Example:

Calculate the hydrogen ion concentration [H_3O^+] for the solutions with the following pH values: (a) 3.47, (b) 0.20, (c) 8.60 .

solution:

pH = - log [H_3O^+]

log[H_3O^+] = - pH

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

(a) [H_3O^+] = 10^{-pH} = $10^{-3.47}$ = 3.4 × 10^{-4} M .

(b) [H_3O^+] = 10^{-pH} = $10^{-0.2}$ = 6.3× 10^{-1} M .

(c) $[H_3O^+] = 10^{-pH} = 10^{-8.6} = 2.5 \times 10^{-9}$ M .

Example:

Calculate the change in pH for 0.01 M HCl solution on 10 times dilution.

Solution:

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

a. Before dilution (original solution)

$[H_3O^+] = 0.01 = 10^{-2}$ M   **$pH = -\log (0.01) = 2$**

b. After dilution for 10 times

$M_1V_1 = M_2 V_2$

$1 \times 0.01 = 10 \times M_2$

$M_2 = 0.001 = 10^{-3}$ M   **$pH = -\log(10^{-3}) = 3$**

$\Delta pH = 3 - 2 = 1$

Then Changing the concentration for 10 times changes the pH by 1 unit

Exercise:

Calculate the change in pH for 0.1 M HNO_3 solution on dilution of 100 times.

Example :

Calculate the pH of a solution obtained by mixing the following volumes of the two solutions of the strong acid HCl :

a) 100 mL of (pH= 2)

b) 500 mL of (pH= 4) .

solution:

$HCl + H_2O \rightarrow H_3O^+ + Cl^-$

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

$[H_3O^+]_a = 10^{-2}$ M the concentration of the pH =2 acid

$[H_3O^+]_b = 10^{-4}$ M the concentration of the pH =4 acid

No. of moles of $[H_3O^+]_{total}$ = No. of moles of $[H_3O^+]_a$ + No. of moles of $[H_3O^+]_b$

No. of moles = Molarity (M) x Volume (liter)

$V_a(\text{liter}) = \frac{100\ (mL)}{1000} = 0.1$ L $\qquad V_b(\text{liter}) = \frac{500\ (mL)}{1000} = 0.5$ L

No. of moles of $[H_3O^+]_a = 10^{-2}$ M x 0.1 liter $=10^{-3}$ mole

No. of moles of $[H_3O^+]_b = 10^{-4}$ M x 0.5 liter = 5 x 10^{-5} mole

No. of moles of $[H_3O^+]_{total}$ = 5 x 10^{-5} + 10^{-3} = 1.05 x 10^{-3} moles

Molarity of the resulting solution = $\frac{\text{No.of moles of } [H_3O^+]_{total}}{(V_a+V_b)liter}$

Molarity of the resulting solution = $\frac{1.05 \text{ x}10^{-3}\ mole}{(0.1+0.5)liter}$ = 1.75 x 10^{-3} M

$[H_3O^+]_{total}$ = 1.75 x 10^{-3} M

pH = -log (1.75 x 10^{-3} M) = 2.75

Exercise 1: Calculate the pH of the acidic solution obtained by mixing 100 mL of (pH= 2) of HCl with 400 mL of (pH= 3) of HNO_3.

Exercise 2: Calculate the pH of the basic solution obtained by mixing 200 mL of (pH= 10) of KOH with 300 mL of (pH= 8) of NaOH.

Example :

Calculate the pH of a solution obtained by mixing 50 mL of the strong acid HCl solution (pH= 3.0) with 10 mL of the strong base KOH solution (pH= 12.0) .

Answer:

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

$[H_3O^+]$ for HCl solution = 1.0×10^{-3} M .

$[H_3O^+]$ for KOH solution = 1.0×10^{-12} M .

As $[OH^-] = \frac{K_w}{[H_3O^+]}$ then

$[OH^-]$ for KOH solution = $\frac{1.0x10^{-14}}{1.0x10^{-12}}$ = $1.0\ x10^{-2}$ M

mmole HCl = Molarity x volume(mL)

mmol HCl = $1.0\ x10^{-3}$ M x 50 mL = 0.05 mmol

mmol KOH = $1.0\ x10^{-2}$ M x 10mL = 0.1 mmol

HCl	+	KOH	→	KCl + H_2O
0.05 mmol		0.1 mmol		

Excess of KOH = mmole KOH – mmole HCl

Excess of KOH = (0.1 – 0.05)mmol = 0.05 mmole

$[OH^-] = \frac{0.05\ mmol}{(50+10)mL} = 8.33\ x10^{-4}$M

pOH = - log $[OH^-]$ = - log (8.33×10^{-4}) = 3.08

pH = 14 – 3.08 = 10.92

Exercise: Calculate the pH of the solution obtained by mixing 10 mL of 0.20 M H_2SO_4 and 20 mL of 0.30 M NaOH .