

# ALMUSTAQBAL UNIVERSITY

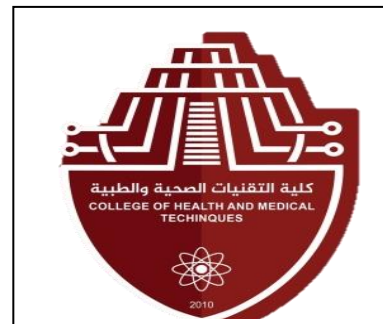
College of Health and Medical Techniques

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



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

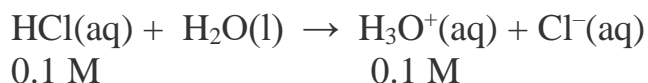
### pH of a strong acids:

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.



$[\text{H}_3\text{O}^+] = C =$  The original concentration of the strong acid  $[\text{HCl}] = 0.1 \text{ M}$

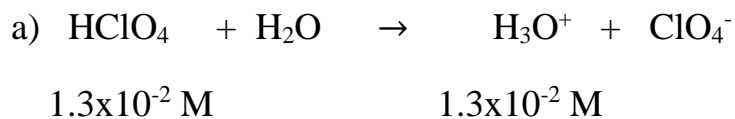
$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log (0.1) = 1$$

**Example:**

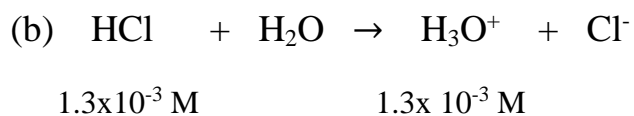
Calculate the pH of the following strong acid solutions:

(a)  $1.3 \times 10^{-2} \text{ M HClO}_4$ , (b)  $1.3 \times 10^{-3} \text{ M HCl}$ , (c)  $1.3 \times 10^{-4} \text{ M HNO}_3$ .

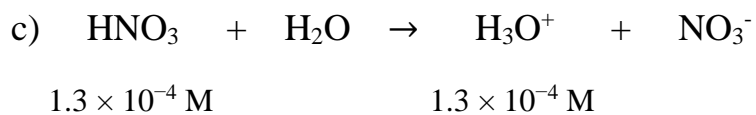
Solution:



$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log 1.3 \times 10^{-2} = 1.89$$



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$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log [1.3 \times 10^{-4}] = 3.89$$

**Example:**

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

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

solution:



$$0.05 \text{ M} \qquad \qquad \qquad 0.05 \text{ M}$$

$$\text{pOH} = -\log [\text{OH}^-] = -\log 5 \times 10^{-2} = 1.3$$

**As  $\text{pH} + \text{pOH} = 14$**

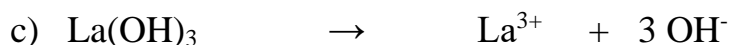
$$\text{pH} = 14 - 1.3 = 12.7$$



$$0.05 \text{ M} \qquad \qquad \qquad 2(0.05) = 0.1 \text{ M}$$

$$\text{pOH} = -\log [\text{OH}^-] = -\log 0.1 = 1.0$$

$$\text{pH} = 14 - 1.0 = 13.0$$



$$0.05 \text{ M} \qquad \qquad \qquad 3(0.05) = 0.15 \text{ M}$$

$$\text{pOH} = -\log [\text{OH}^-] = -\log 0.15 = 0.82$$

$$\text{pH} = 14 - 0.82 = 13.18$$

**Example:**

Calculate the hydrogen ion concentration  $[\text{H}_3\text{O}^+]$  for the solutions with the following pH values: (a) 3.47, (b) 0.20, (c) 8.60 .

solution:

$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

$$\log[\text{H}_3\text{O}^+] = -\text{pH}$$

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$$

$$\text{(a) } [\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-3.47} = 3.4 \times 10^{-4} \text{ M} .$$

$$\text{(b) } [\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-0.2} = 6.3 \times 10^{-1} \text{ M} .$$

$$(c) [\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-8.6} = 2.5 \times 10^{-9} \text{ M} .$$

**Example:**

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

**Solution:**

$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

**a. Before dilution (original solution)**

$$[\text{H}_3\text{O}^+] = 0.01 = 10^{-2} \text{ M} \quad \text{pH} = -\log (0.01) = 2$$

**b. After dilution for 10 times**

$$M_1 V_1 = M_2 V_2$$

$$1 \times 0.01 = 10 \times M_2$$

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

$$\Delta \text{pH} = 3 - 2 = 1$$

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

**Exercise:**

**Calculate the change in pH for 0.1 M HNO<sub>3</sub> solution on dilution of 100 times.**

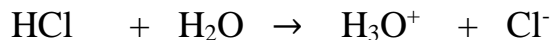
**Example :**

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:



$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$$

$[\text{H}_3\text{O}^+]_a = 10^{-2} \text{ M}$  the concentration of the pH =2 acid

$[\text{H}_3\text{O}^+]_b = 10^{-4} \text{ M}$  the concentration of the pH =4 acid

No. of moles of  $[\text{H}_3\text{O}^+]_{\text{total}} = \text{No. of moles of } [\text{H}_3\text{O}^+]_a + \text{No. of moles of } [\text{H}_3\text{O}^+]_b$

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

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

No. of moles of  $[\text{H}_3\text{O}^+]_a = 10^{-2} \text{ M} \times 0.1 \text{ liter} = 10^{-3} \text{ mole}$

No. of moles of  $[\text{H}_3\text{O}^+]_b = 10^{-4} \text{ M} \times 0.5 \text{ liter} = 5 \times 10^{-5} \text{ mole}$

No. of moles of  $[\text{H}_3\text{O}^+]_{\text{total}} = 5 \times 10^{-5} + 10^{-3} = 1.05 \times 10^{-3} \text{ moles}$

$$\text{Molarity of the resulting solution} = \frac{\text{No. of moles of } [\text{H}_3\text{O}^+]_{\text{total}}}{(V_a + V_b) \text{ liter}}$$

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

$$[\text{H}_3\text{O}^+]_{\text{total}} = 1.75 \times 10^{-3} \text{ M}$$

$$\text{pH} = -\log ( 1.75 \times 10^{-3} \text{ M} ) = 2.75$$

**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<sub>3</sub>.

**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:

$$[ \text{H}_3\text{O}^+ ] = 10^{-\text{pH}}$$

$$[ \text{H}_3\text{O}^+ ] \text{ for HCl solution} = 1.0 \times 10^{-3} \text{ M .}$$

$$[ \text{H}_3\text{O}^+ ] \text{ for KOH solution} = 1.0 \times 10^{-12} \text{ M .}$$

$$\text{As } [\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]} \text{ then}$$

$$[\text{OH}^-] \text{ for KOH solution} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-12}} = 1.0 \times 10^{-2} \text{ M}$$

$$\text{mmole HCl} = \text{Molarity} \times \text{volume(mL)}$$

$$\text{mmol HCl} = 1.0 \times 10^{-3} \text{ M} \times 50 \text{ mL} = 0.05 \text{ mmol}$$

$$\text{mmol KOH} = 1.0 \times 10^{-2} \text{ M} \times 10 \text{ mL} = 0.1 \text{ mmol}$$



$$0.05 \text{ mmol} \quad \quad \quad 0.1 \text{ mmol}$$

$$\text{Excess of KOH} = \text{mmole KOH} - \text{mmole HCl}$$

$$\text{Excess of KOH} = (0.1 - 0.05) \text{ mmol} = 0.05 \text{ mmole}$$

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

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

$$\text{pH} = 14 - 3.08 = 10.92$$

**Exercise:** Calculate the pH of the solution obtained by mixing 10 mL of 0.20 M  $\text{H}_2\text{SO}_4$  and 20 mL of 0.30 M NaOH .