

Problems on pH & Buffers

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Q1- Calculate the pH of the following :

A) 0.005 M HNO₃ ?



$$[\text{H}^+] = 0.005 \longrightarrow \text{pH} = -\log [\text{H}^+] = -\log [0.005] = \boxed{2.3}$$

B) 0.02 H₂SO₄ ?

$$[\text{H}^+] = 0.04$$

$$\text{pH} = -\log [\text{H}^+] = -\log [0.04] = \boxed{1.4}$$

C) 0.1 M HCOOH (K_a = 1.8 x 10⁻⁴) ?
weak acid

$$K_a = \frac{[\text{H}^+][\text{HCOO}^-]}{[\text{HCOOH}]} = \frac{[\text{H}^+]^2}{[\text{HCOOH}]} \longrightarrow 1.8 \times 10^{-4} = \frac{[\text{H}^+]^2}{0.1}, \quad [\text{H}^+] = 4.24 \times 10^{-3}$$
$$\text{pH} = -\log [\text{H}^+] = \boxed{2.37}$$

Q2- Determine the pH of a solution made by dissolving 0.01 moles of benzoic acid (C₆H₅COOH; pK_a = 4.2) in water to form 1 liter of solution. What is the pH?

weak acid , M = 0.01 mol/L , K_a = 10^{-4.2} , pH = ??



$$K_a = \frac{[\text{H}^+]^2}{[\text{C}_6\text{H}_5\text{COOH}]} \rightarrow 10^{-4.2} = \frac{[\text{H}^+]^2}{[0.01]} \rightarrow [\text{H}^+] = 7.94 \times 10^{-4}$$

$$\text{pH} = -\log [\text{H}^+] = \boxed{3.1}$$

Q3- The pK_b of a base is 4.5. What is the pH of a 0.01 M solution of the base?

$$K_b = 10^{-4.5} \rightarrow K_b = \frac{[\text{OH}^-]^2}{[x\text{OH}]} \rightarrow 10^{-4.5} = \frac{[\text{OH}^-]^2}{0.01} \rightarrow [\text{OH}^-] = 5.63 \times 10^{-4}$$

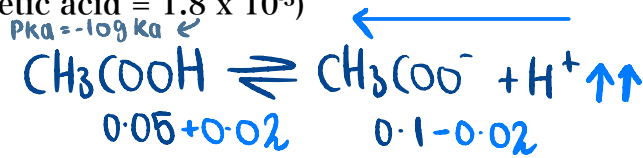
$$\text{pOH} = -\log [\text{OH}^-] = 3.25$$

$$\text{pH} = \boxed{10.75}$$

Q4- Calculate the pH of a phosphate buffer that contains 60% monosodium phosphate (NaH₂PO₄) and 40% disodium phosphate (Na₂HPO₄). (pK_a = 7.2)

$$\text{pH} = \text{pK}_a + \log \frac{[\text{x}^-]}{[\text{Hx}]} \rightarrow \text{pH} = 7.2 + \log \frac{[40]}{[60]} = \boxed{7.02}$$

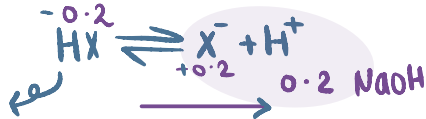
Q5- Predict and calculate the pH of a buffer containing 0.05 M acetic acid (CH_3COOH) and 0.1 M sodium acetate (CH_3COONa), when 0.02 M HCl is added to the solution. (K_a for acetic acid = 1.8×10^{-5})



$$\text{pH} = \text{pK}_a + \log \frac{[\text{X}^-]}{[\text{HX}]} \rightarrow \text{pH} = 4.7 + \log \frac{[0.1 - 0.02]}{[0.05 + 0.02]}$$

$$\boxed{\text{pH} = 4.75}$$

Q6- Calculate the pH of a solution prepared by dissolving 500 mg of a monoprotic acid (with molecular weight 120 g/mol) in 20 mL of 0.2 M NaOH. The pK_a of the acid is 6.8.



$$\text{MW} = \frac{g}{\text{mol}} \rightarrow \frac{g}{\text{MW}} = \text{mol} \rightarrow \frac{500 \text{ mg}}{120 \frac{g}{\text{mol}}} = 4.16 \text{ mmol} \rightarrow M = \frac{4.16 \text{ mmol}}{20 \text{ mL}} = 0.21 \text{ M}$$

$$\text{pH} = 6.8 + \log \frac{[0.2]}{[0.21 - 0.2]} = \boxed{8.1}$$

a) Calculate the pH of a solution prepared by dissolving 0.03 moles of formic acid (HCOOH ; $\text{pK}_a = 3.75$) in water to make 1 liter of solution.

$$M = 0.03 \text{ mol/L}$$

$$K_a = \frac{[\text{H}^+]^2}{[\text{HX}]} \Rightarrow 10^{-3.75} = \frac{[\text{H}^+]^2}{0.03} \rightarrow [\text{H}^+] = 2.3 \times 10^{-3}$$

$$-\log [2.3 \times 10^{-3}] = \boxed{2.64}$$

b) After preparing the solution in part (a), 0.01 moles of concentrated sodium hydroxide (NaOH) were added. What is the new pH of the solution? (Assume no volume change due to the addition of NaOH.)



$$\text{pH} = 3.75 + \log \frac{[2.3 \times 10^{-3} + 0.01]}{[0.03 - 0.01]} = \boxed{3.54}$$