



Al-Mustaqbal-College University
Chemical Engineering and Petroleum
Industry Department
Analytical chemistry
First class / first term
Lecture Four
Buffer Solution
By
Asst. lect. Ban Ali Hassan

Lecture four

Example 2 \ find the pH of a 0.012 M sodium hydroxide (NaOH) solution.

Sol. :

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

$$= -\log [0.012] = 1.92$$

$$\text{pH} + \text{p(OH)} = 14$$

$\text{pH} = 14 - 1.92 = 12.08$, the solution has **pH of 12.08** and is a **strong base**.

Lecture four

pH and pOH are related to one another; **THEY ARE NOT INDEPENDENT OF EACH OTHER.** As pH increases, pOH decreases. As pH decreases, pOH increases. By knowing what ion you are measuring on which scale, this will tell you whether or not the solution is acidic or basic.

Ion Concentration	Solution Type	pH	pOH
$[H^+] > [OH^-]$	Acidic	$pH < 7$	$pOH > 7$
$[H^+] < [OH^-]$	Basic	$pH > 7$	$pOH < 7$
$[H^+] = [OH^-]$	Neutral	$pH = 7$	$pOH = 7$

Because these scales are related, an equation can be used to explain their correlation.

$$pH + pOH = 14$$

$$pH = 14 - pOH$$

$$pOH = 14 - pH$$

The bottom equations are manipulations of the top equation

The **KEYS** to calculations are knowing:

1. Knowing what type of solution you are working with
2. What equation to use first

Example Calculations:

1. Calculate the pH and pOH of a 0.33 M H₂SO₄ solution.

- Are you working with an acid or a base? _____
- Are you given the concentration, or *Molarity* of that solution? _____
- Next, plug the concentration or Molarity into the correct equation.

Because this is an acidic solution, we have to calculate pH first.

$$pH = -\log[0.33] = \underline{\hspace{2cm}}$$

Now we can calculate the pOH, because we have calculated the pH.

$$pOH = 14 - pH = \underline{\hspace{2cm}} = \underline{\hspace{2cm}}$$

2. Calculate the pH and pOH of a 0.25 M NaOH solution.

- Are you working with an acid or a base? _____
- Are you given the concentration, or *Molarity* of that solution? _____
- Next, plug the concentration or Molarity into the correct equation.

Because this is a basic solution, we have to calculate pOH first.

$$pOH = -\log[0.25] = \underline{\hspace{2cm}}$$

Now we can calculate the pH, because we have calculated the pOH.

$$pH = 14 - pOH = \underline{\hspace{2cm}} = \underline{\hspace{2cm}}$$

Lecture four

Ex/ What is the pH of a 0.0005 M solution of NaOH at 25 °C ?



$$[\text{OH}^-] = 0.0005 \text{ M} = 5 \times 10^{-4} \text{ M}$$

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

$$= -\log 5 \times 10^{-4}$$

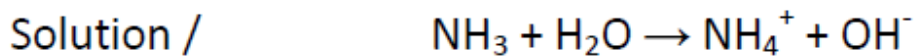
$$= -\log 5 + 4 \log 10$$

$$= -0.699 + 4$$

$$= 3.301$$

$$\text{pH} = 14 - 3.301 = 10.7$$

Ex/ What is the pH of a 0.1 M NH_3 solution ? K_b 1.8×10^{-5}



$$0.1 \qquad \qquad 0 \qquad 0$$

$$0.1 - X \qquad \qquad X \qquad X$$

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$

$$1.8 \times 10^{-5} = \frac{(X)(X)}{0.1 - X}$$

$$1.8 \times 10^{-5} = \frac{X^2}{0.1}$$

$$X^2 = 1.8 \times 10^{-6}$$

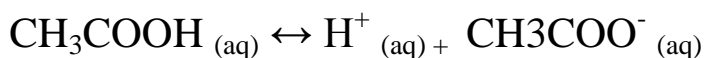
$$X = 1.34 \times 10^{-3} = [\text{OH}^-]$$

Lecture four

In this case, if the solution contained equal molar concentrations of both the acid and its salt, it would have a pH of 4.76. It wouldn't matter what the concentrations were, as long as ethanoate in solution. They were the same.

It can change the pH of the buffer solution by changing the ratio of acid to salt, or by choosing a different acid and one of its salts.

Ethanoic acid is a weak acid, and the position of this equilibrium will be to the left:



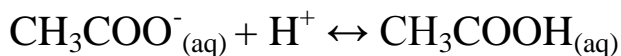
a- Adding sodium ethanoate to this adds lots of extra ethanoate ions.

According to Le Chatelier's Principle, that will tip the position of the equilibrium even further to the left. The solution will therefore contain these important things:

- Lots of un-ionized ethanoic acid.
- Lots of ethanoate ions from sodium ethanoate.
- Enough hydrogen ions to make the solution acidic.

b- Adding an acid to this buffer solution

The buffer solution must remove most of the new hydrogen ions otherwise the pH would drop markedly. Hydrogen ions combine with the ethanoate ions to make ethanoic acid. Although the reaction is reversible, since ethanoic acid is a weak acid, most of the new hydrogen ions are removed in this way so the pH won't change very much but because of the equilibrium involved, it will fall a little bit.



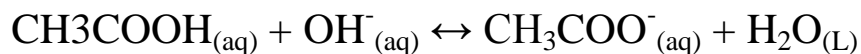
c- Adding an alkali to this buffer solution

Alkaline solutions contain hydroxide ions and the buffer solution removes most of these. This time the situation is a bit more complicated because there are two processes which can remove hydroxide ions.

Lecture four

- **Removal by reacting with ethanoic acid**

The most likely acidic substance which a hydroxide ion is going to collide with is an ethanoic acid molecule. They will react to form ethanoate ions and water.



Because most of the new hydroxide ions are removed, the pH doesn't increase very much.

- **Removal of the hydroxide ions by reacting with hydrogen ions**

Remember that there are some hydrogen ions present from the ionization of the ethanoic acid.



Hydroxide ions can combine with these to make water. As soon as this happens, the equilibrium tips to replace them. This keeps on happening until most of the hydroxide ions are removed.

Equilibrium moves to replace the removed hydrogen ions.



Hydroxide ions combine with these to make water

Again, because you have eq[ui]librium, hydroxide ions are removed just most of them. The water formed re-ionizes to a very small extent to give a few hydrogen ions and hydroxide ions.