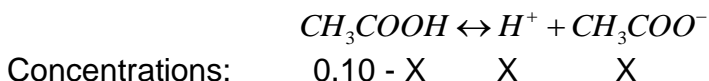


Problem :

- A) A beaker contains 10 ml of 0.10 M acetic acid.
What is the pH of the solution?
- B) To the beaker in part (A), 10 ml more of 0.10M acetic acid is added.
What is the pH of the solution now?
- C) To the beaker in part (A) is added 10 ml of pure water.
What is the pH of the new solution?
- D) To the beaker in part (A) is added 10 ml of 0.10M sodium acetate.
What is the pH of this new solution?

Solution :

(A) : Acetic acid is a weak acid. It has an ionization constant of 1.8×10^{-5} .



$$K_{eq} = \frac{[H^+][CH_3COO^-]}{CH_3COOH} = 1.8 \times 10^{-5}$$

$$1.8 \times 10^{-5} = \frac{(X)(X)}{(0.10 - X)} \rightarrow \frac{X^2}{0.10}; \text{ assuming } X \ll 0.10 \text{ (to}$$

be verified)

$$\text{Solve for } X \text{ (i.e., } H^+) \rightarrow [H^+] = 1.34 \times 10^{-3} \text{ moles/liter}$$

Note: this result is much smaller than .10, so the assumption is valid.

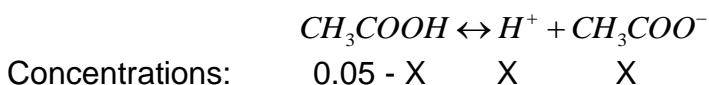
$$pH = -\log[H^+] \rightarrow -\log(1.34 \times 10^{-3})$$

OK, you solve this, to get... **pH = 2.9**

B) You've added both acid and water in the same proportion as originally in the beaker.

The new solution is the same as the original, except there's more of it.
 Therefore, the
 pH of the new solution is the same as that of the old one... **pH = 2.9**

C) Here, you've doubled the volume of the solution, but kept the number of moles of acetic acid, the same. Therefore, the concentration of acid is reduced by half. So, instead of 0.10M acetic acid, we now have 0.05M acetic acid :



$$K_{eq} = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]} = 1.8 \times 10^{-5}$$

$$1.8 \times 10^{-5} = \frac{(X)(X)}{(0.05 - X)} \rightarrow \frac{X^2}{0.05}; \text{ assuming } X \ll 0.05 \text{ (to$$

be verified)

$$\text{Solve for } X \text{ (i.e., } H^+) \rightarrow [H^+] = 9.49 \times 10^{-4} \text{ moles/liter}$$

Note: this result is much smaller than .05, so the assumption is valid.

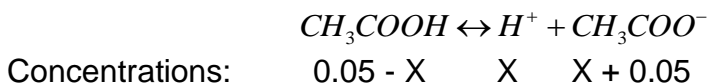
$$pH = -\log[H^+] \rightarrow -\log(9.49 \times 10^{-4})$$

Again, you solve this, to get... **pH = 3.0**

D) Once again, the volume of the solution is doubled when the sodium acetate is added.

Therefore the original molarities of acetic acid and sodium acetate are halved, since the Amounts of acid and of salt are unchanged.

Now, we have the acetate ion common to the ionized acid and to the ionized salt...



$$K_{eq} = \frac{[H^+][CH_3COO^-]}{CH_3COOH} = 1.8 \times 10^{-5}$$

$$1.8 \times 10^{-5} = \frac{(X)(X + 0.05)}{(0.05 - X)}$$

simplifies a lot... Assuming that $x \ll 0.05$, show that the above equation

$$1.8 \times 10^{-5} = (X) \rightarrow [H^+] = 1.8 \times 10^{-5} \text{ moles/liter}$$

is valid. Note: this result is much smaller than .05, so the assumption

$$pH = -\log[H^+] \rightarrow -\log(1.8 \times 10^{-5})$$

Again, you solve this, to get... **pH = 4.7**