Answer on Question #82833, Chemistry / General Chemistry

0.571 mol of a weak acid, HA, and 11.6 g of NaOH are placed in enough water to produce 1.00 L of solution. The final pH of this solution is 4.14. Calculate the ionization constant, Ka, of HA.

Solution

First of all find the amount of NaOH placed in the solution:

v (NaOH) = $\frac{m}{M}$ = $\frac{11.6}{40}$ = 0.29 (mol)

NaOH reacts with weak acid HA which is in excess:

NaOH + HA ➔NaA + H_2O

Base and acid react in ratio 1:1, hence the amount of NaA obtained is 0.29 mol.

Find the amount of HA left after the reaction:

v (HA) = 0.571 – 0.29 = 0.281 (mol)

After the reaction the solution turns into a buffer with 0.29 mol of NaA and 0.281 mol of HA.

Use the formula for ionization constants of acids in buffers:

$K_a = \frac{[H^+]\times[A^-]}{[HA]} = \frac{[H^+]\times[NaA]}{[HA]}$; where $[H^+]$, [HA], [NaA] – concentrations;

Find $[H^+]$:

pH = – lg$[H^+]$, thus $[H^+] = 10^{-4.14} = 7.24 \times 10^{-5}$;

$K_a = \frac{7.24\times10^{-5}\times0.29}{0.281}$ = **7.48×10^{-5}**

Answer

7.48×10^{-5} is the ionization constant of HA.