

Answer on Question #82833, Chemistry / General Chemistry

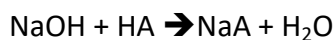
0.571 mol of a weak acid, HA, and 11.6 g of NaOH are placed in enough water to produce 1.00 L of solution. The final pH of this solution is 4.14. Calculate the ionization constant, K_a , of HA.

Solution

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

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

NaOH reacts with weak acid HA which is in excess:



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:

$$v(\text{HA}) = 0.571 - 0.29 = 0.281 \text{ (mol)}$$

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:

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

Find $[\text{H}^+]$:

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

$$K_a = \frac{7.24 \times 10^{-5} \times 0.29}{0.281} = \mathbf{7.48 \times 10^{-5}}$$

Answer

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