

Answer on the question #63364, Chemistry / General Chemistry

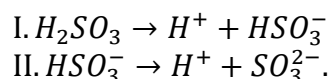
Question :

Determine the concentrations of the ionic species present in a 0.00222 M solution of the H_2SO_3 . ($pK_{a1} = 1.85$, $pK_{a2} = 7.20$).

$[H_2SO_3]$ M, $[HSO_3^-]$ M, $[SO_3^{2-}]$ M, $[H_3O^+]$ M, $[OH^-]$ M

Solution:

The dissociation steps are:



The relation for equilibrium constants:

$$\begin{aligned} K_{a1} &= \frac{[HSO_3^-][H^+]}{[H_2SO_3]} \\ K_{a2} &= \frac{[SO_3^{2-}][H^+]}{[HSO_3^-]} \end{aligned}$$

As pK_{a2} is much bigger than pK_{a1} , we can assume that the dissociation by second stage is merely marginal. Then:

$$\begin{aligned} [HSO_3^-] &= [H^+] = x \\ [H_2SO_3] &= 0.00222 - x \end{aligned}$$

$$K_{a1} = \frac{x^2}{0.00222 - x} = 10^{-1.85}$$

$$x = 0.001953 \text{ M}$$

$$[HSO_3^-] = [H^+] = 0.001953 \text{ M}; [H_2SO_3] = 0.00222 - x = 0.000267 \text{ M}.$$

Then, taking the concentration of HSO_3^- , H^+ we calculate small dissociation for the second stage:

$$K_{a2} = \frac{[SO_3^{2-}][H^+]}{[HSO_3^-]}$$

$$[SO_3^{2-}] = K_{a2} \cdot \frac{[HSO_3^-]}{[H^+]}$$

$$[SO_3^{2-}] = 10^{-7.20} = 6.3 \cdot 10^{-8} \text{ M}.$$

Now we see, that our assumption was correct, sulfite anion concentration is really much lower.

Let's calculate the concentration of hydroxide anion:

$$K_w = [H^+][OH^-] = 10^{-14}$$

$$[OH^-] = \frac{10^{-14}}{[H^+]} = 5.12 \cdot 10^{-12}$$

Answer: $[HSO_3^-] = [H^+] = 0.001953 \text{ M}$, $[H_2SO_3] = 0.000267 \text{ M}$, $[SO_3^{2-}] = 6.3 \cdot 10^{-8} \text{ M}$, $[OH^-] = 5.12 \cdot 10^{-12} \text{ M}$.

Remark: of course, H_3O^+ and H^+ cation is the same (not exactly, but in the meaning of concentration), we just used H^+ notation to make clearer calculations. You can take the concentration of H^+ , obtained here and use it as a concentration of H_3O^+ .

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