What are the equilibrium concentrations of H2SO3, H+, HSO3−, and SO32−in a 0.050 M solution of sulfurous acid H2SO3 at 25 oC? For H2SO3 at 25 oC, Ka1 = 1.5×10−2 and Ka2 = 1.0×10−7.

Solution:

$H_2SO_3 \leftrightarrow H^+ + HSO_3^-$

$K_{a1} = \frac{[H^+][HSO_3^-]}{[H_2SO_3]}$

From this equation $[H^+][HSO_3^-] = K_{a1} \cdot [H_2SO_3] = 1.5 \cdot 10^{-2} \cdot 0.050 = 7.5 \cdot 10^{-4}$.

The concentration of $H^+ = HSO_3^- = \sqrt{7.5 \cdot 10^{-4}} = 2.738 \cdot 10^{-2}$ M.

$HSO_3^- \leftrightarrow H^+ + SO_3^{2-}$

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

So, $[H^+][SO_3^{2-}] = K_{a2} \cdot [HSO_3^-] = 1.0 \cdot 10^{-2} \cdot 2.738 \cdot 10^{-2} = 2.738 \cdot 10^{-9}$

The concentration of $H^+ = SO_3^{2-} = \sqrt{2.738 \cdot 10^{-9}} = 5.23 \cdot 10^{-5}$ M.

Answer:

The equilibrium concentrations in 0.050 M solution are:

$[H_2SO_3] = 5 \cdot 10^{-2} - 2.738 \cdot 10^{-2} =$ **$2.262 \cdot 10^{-2}$ M;**

$[H^+] = 2.738 \cdot 10^{-2} + 5.23 \cdot 10^{-5} =$ **$2.743 \cdot 10^{-2}$ M**,

$[HSO_3^-] = 2.738 \cdot 10^{-2} - 5.23 \cdot 10^{-5} =$ **$2.733 \cdot 10^{-2}$ M**, and

$[SO_3^{2-}] =$ **$5.23 \cdot 10^{-5}$ M**.