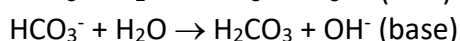
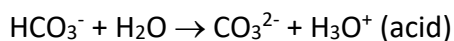


Answer on Question #82694 – Chemistry – Physical Chemistry

The bicarbonate ion, HCO_3^- can potentially act as an acid or as a base in water. Write the chemical equations for these two processes, and from the information given in Manahan, determine the corresponding acid and base dissociation constants. Given the relative magnitudes of the dissociation constants, decide whether the dominant reaction of bicarbonate in water will be as an acid or as a base. Calculate the pH of an aqueous 0.010 M solution of sodium bicarbonate in water using the dominant reaction alone, and assuming that the amounts of carbonate ion and carbonic acid from other sources are negligible in this case.

Solution:



The equilibrium in which the HCO_3^- ion acts as a Brnsted acid is described by K_{a2} for carbonic acid.

$$K_{a2} = \frac{[\text{H}_3\text{O}^+][\text{CO}_3^{2-}]}{[\text{HCO}_3^-]} = 4.7 \times 10^{-11}$$

The equilibrium in which the HCO_3^- acts as a Brnsted base is described by K_{b2} for the carbonate ion.

$$K_{b2} = \frac{[\text{H}_2\text{CO}_3][\text{OH}^-]}{[\text{HCO}_3^-]} = 2.2 \times 10^{-8}$$

Since K_{b2} is significantly larger than K_{a2} , the HCO_3^- ion is a stronger base than it is an acid.

Bicarbonate is an ampholyte

$$\text{pH} = \frac{pK_{a, \text{H}_2\text{CO}_3} + pK_{a, \text{HCO}_3^-}}{2} = (6.35 + 10.32) / 2 = 8.34$$

Answer provided by www.AssignmentExpert.com