

$$K_{eq} = \frac{[HCO_3^-][PO_4^{2-}]}{[CO_3^{2-}][H_2PO_4^-]} \Rightarrow K_{eq} = \frac{K_a(\text{acid on reactant})}{K_a(\text{acid on product})}$$

The Keq expression can be RE-WRITTEN AS:

$$K_{eq} = \frac{[H_3O^+][HPO_4^{2-}]}{[H_2PO_4^-]} \times \frac{[HCO_3^-]}{[H_3O^+][CO_3^{2-}]} \Rightarrow K_a(H_2PO_4^-) \times \frac{1}{K_a(HCO_3^-)}$$

OR SIMPLY AS:

$$K_{eq} = \frac{K_a(H_2PO_4^-)}{K_a(HCO_3^-)} = \frac{6.2 \times 10^{-8}}{5.6 \times 10^{-11}} = 1.1 \times 10^{+3} \quad (K_{eq} > 1)$$

PRODUCTS
are favoured

RECALL:

If the Keq value = 1 BOTH SIDES EQUALLY FAVOURED

If the Keq value > 1 PRODUCTS ARE FAVOURED

If the Keq value < 1 REACTANTS ARE FAVOURED

THE GENERIC Keq expression for acid-base equilibria is:

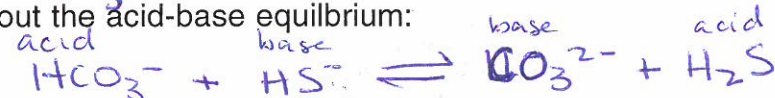
$$K_{eq} = \frac{K_a(\text{REACTANT ACID})}{K_a(\text{PRODUCT ACID})}$$

Example 3. When HS⁻ and HCO₃⁻ are mixed, does the resulting equilibrium favour the reactant or the products?

1. Choose which of these two reactants is going to act as the acid. (higher on table)

HCO₃⁻ will act as the acid

2. Write out the acid-base equilibrium:



3. Identify the TWO ACIDS involved in the equilibrium:

HCO₃⁻ and H₂S

a. Solve the problem using the WEAKER ACID rule

Because HCO₃⁻ is weaker the REACTANTS are favoured.

b. Solve the problem using the Keq equation

$$K_{eq} = \frac{K_a(\text{reactant})}{K_a(\text{product})} = \frac{K_a(HCO_3^-)}{K_a(H_2S)} = \frac{5.6 \times 10^{-11}}{9.1 \times 10^{-8}} = \boxed{6.2 \times 10^{-4}}$$

B/c Keq < 1 the Reactants are favoured!!!

SEATWORK/HOMEWORK: Exercises 38-46 in Hebden pg 131

PLO's: K8+K9