Ka & Kb

I. General :

a. Basic Ka

i. Ka – the acid ionization constant in the equilibrium constant for the ionization of an acid

1. General chemical reaction: $H_2O(e) + \ HA \ \rightleftharpoons A^- + H^+(or\ H_3O^+)$
2. $Ka = \frac{[H^+][A^-]}{[HA]}$
3. The larger the Ka the greater the ionization of the acid, the stronger it is, the lower the pH.

In the other words, pH inversely proportional to Ka

ii. pKa = -log Ka; Another indication of strength

1. because of –log, pH is proportional to pKa, so if pKa is low, pH is low, therefore acidic solution

b. basic Kb:

i. Kb – the base ionization constant is the equilibrium constant for ionization of a base in an aqueous solution.

1. Chemical reaction: $B + H_2O(e) \rightleftharpoons BH^+ + OH^+$
2. $Kb = \frac{[BH^+][OH^-]}{[B]}$

c. $Kw = Ka \times Kb$

i. True for any conjugate acid/base pair.

ii. 1×10^{-14}=Ka × Kw

1. From this and some algebra

pKa +pKa = 14.00

iii. These equations can be used to calculate

1. Ka given Kb or Kb given Ka
2. pKa given pKb or pKb given pKa

iv. a solution is basic if

1. Kb < Ka
2. $[OH]$ > $[H_3O^+]$
3. pKb > pka

- note: If acidic, it is vice versa

II. Example Calculations:

a. Basic

i. Given [ ]'s calculate Ka:

Q: given [HA]=.10M; [A]=1.2e-3 M, calculate Ka.

A: $HA \ \rightleftharpoons A + H^+$ *Note: because $[AC^-] = [H^+], [H^+] = \ 1.2 \times 10^{-3}M$

Ka = $\frac{[AC^-][H^+]}{[HAC]} = \frac{[1.2\times10^{-3}]^2}{.1} = 1.4 \times 10^{-5}$

ii. Given one calculate the other:

Q: what is the KB of acetic acid if the Ka is 1.76e-5?

A: $\mathrm{Ka} \times \mathrm{Kb} = \mathrm{Kw} \rightarrow 1.76\mathrm{e} - 5 \times \mathrm{Kb} = 1 \times 10^{-14}$ $\mathrm{Kb} = 5.7\mathrm{e} - 4$