Sample Exercise: Buffers[10]

Determine the $[H^+]$ and the pH of a buffer containing 1.0 M HF ($K_a = 7.2 \times 10^{-4}$) and 1.0 M NaF. Calculate the percent dissociation of HF in this solution.

The major species initially in solution are: HF, F^-, Na^+ and H_2O.

Na^+ is neither acidic nor basic; H_2O is a *very weak* acid.

The equilibrium expression is : $Ka = \dfrac{[F^-][H^+]}{[HF]} = 7.2 \times 10^{-4}$

Initial concentration (mol/L)		Equilibrium concentration
$[HF]_0 = 1.0$	x mol/L HF dissociates	$[HF] = 1.0 - x$
$[F^-]_0 = 1.0$	→	$[F^-] = 1.0 + x$
$[H^+]_0 = 0$		$[H^+] = x$

Thus at equilibrium:

$Ka = 7.2 \times 10^{-4} = \dfrac{[F^-][H^+]}{[HF]} = \dfrac{(1+x)(x)}{[(1-x)} \approx \dfrac{(1.0)(x)}{1.0}$ Since x is expected to be << 1.0 M

$x = 7.2 \times 10^{-4} \dfrac{1.0}{1.0} = 7.2 \times 10^{-4} = [H^+]$; pH = -log $[H^+]$ = 3.14

The percent dissociation of HF is $\dfrac{[H^+]}{[HF]} \times 100 = \dfrac{7.2 \times 10 - 4}{1.0} \times 100 = 0.072\%$

[10] From "Chemistry" 6th edition by Steven S. Zumdahl and Susan A. Zumdahl