17.2 Buffers

Buffer pH – The Henderson–Hasselbalch Equation

Calculate the pH of a buffer solution made from 1.00 L of a 0.133 M hydrofluoric acid and 0.243 mol of sodium fluoride.
Ka HF = 7.2×10^{-4}

$$pH = pK_a + \log_{10}\frac{[\text{Buffer Base}]}{[\text{Buffer Acid}]}$$

$[\text{Buffer Base}] = [F^-] = 0.243$; $[\text{Buffer Acid}] = [HF] = 0.133$

$$pH = -\log_{10}(7.2\times10^{-4}) + \log_{10}\left(\frac{0.243}{0.133}\right)$$

$$= 3.143 + \log_{10}(1.827)$$

$$= 3.143 + 0.262$$

$$= 3.40$$

17.2 Buffers

Making Buffer Solutions

$$pH = pK_a + \log_{10}\frac{[\text{Buffer Base}]}{[\text{Buffer Acid}]}$$

but when [Buffer Base] = [Buffer Acid]

$$pH = pK_a + \log_{10}(1)$$

$$pH = pK_a$$

Preparing Buffer Solutions

$[HCO_2H]$	$[NaHCO_2]$
0.10 M =	0.10 M

- (selected) $HCO_2H/NaHCO_2$
- $H_2CO_3/NaHCO_3$
- $HOCl/NaOCl$
- H_3BO_3/NaH_2BO_3
- NH_4Cl/NH_3
- $NaHCO_3/Na_2CO_3$

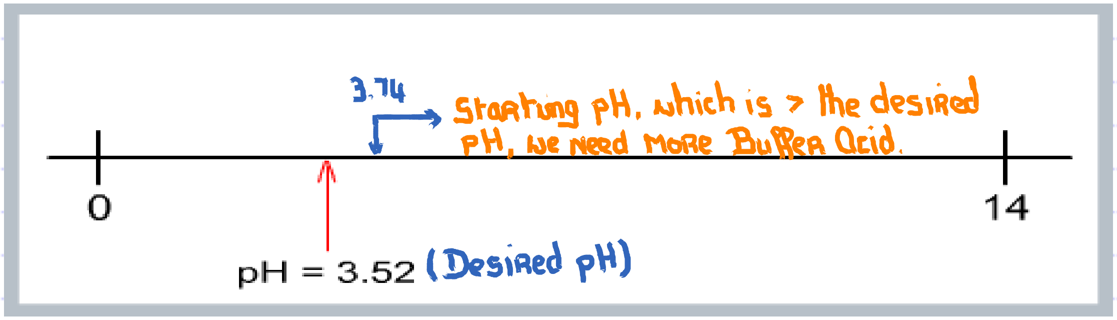
New Target

See Class Web Site.

$K_a = 1.8 \times 10^{-4}$	$pK_a = 3.74$
$K_a = 4.2 \times 10^{-7}$	$pK_a = 6.38$
$K_a = 3.5 \times 10^{-8}$	$pK_a = 7.46$
$K_a = 7.3 \times 10^{-10}$	$pK_a = 9.14$
$K_a = 5.6 \times 10^{-10}$	$pK_a = 9.25$
$K_a = 4.8 \times 10^{-11}$	$pK_a = 10.32$

1) Chose an acid/base combination whose acid pKa is closest to the desired pH.

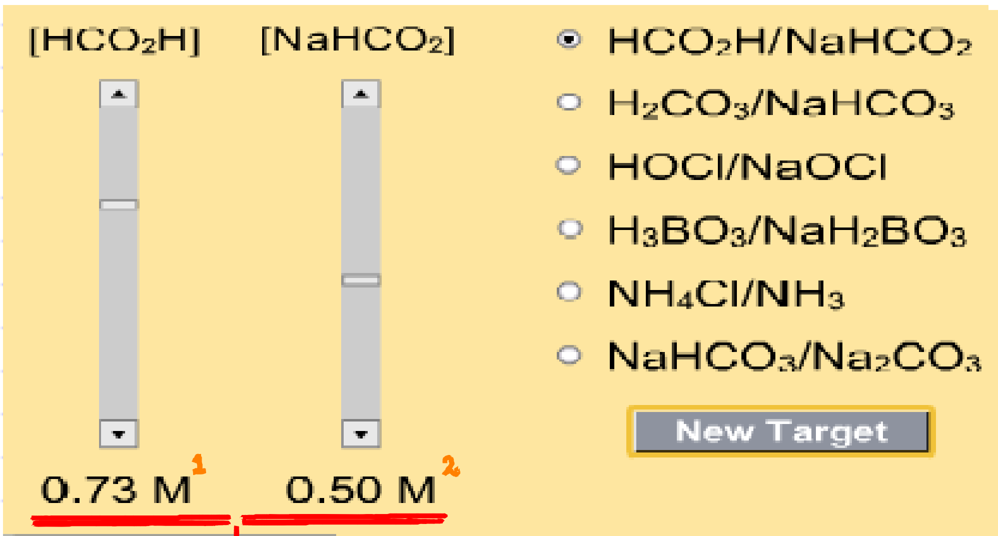
2) Adjust the acid or base concentration to get the desired pH.



17.2 Buffers

Buffer Capacity

Preparing Buffer Solutions



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Buffer Capacity

1. Maximum amount of OH^- that can be removed = [HCO_2H] = 0.73 M
2. Maximum amount of H_3O^+ that can be removed = [HCO_2^-] = 0.50 M