Announcements – Lecture XVII – Thursday, Mar 29th

1. iClicker:  Pick any letter a-e

2. Quiz 7: Place in basket on the front bench.

3. **Exam II: Moved to Saturday, April 7th.**

Identifying Buffers

Which of the following aqueous solutions are good buffer systems?

- [x] 0.34 M ammonium bromide + 0.36 M ammonia (NH_4^+ / NH_3)
- [ ] 0.22 M nitric acid + 0.16 M potassium nitrate (HNO_3) — HNO_3: Strong acid
- [x] 0.32 M nitrous acid + 0.21 M potassium nitrite (HNO_2 / NO_2^-)
- [ ] 0.18 M barium hydroxide + 0.21 M barium bromide ($Ba(OH)_2$) — $Ba(OH)_2$: Strong base
- [x] 0.14 M hydrofluoric acid + 0.20 M sodium fluoride (HF / F^-)

Looking for a weak acid-base conjugate pair.

17.2 Buffers

Buffer pH – ICE

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.
K_a HF = 7.2×10^{-4}

Shortest way to approach this is from the Buffer Acid equilibrium as this directly determines pH.

	HF	+	H_2O	$\Leftrightarrow$	H_3O^+	+	F^-
I	0.133				0		0.243
C	$-x$				x		x
E	$0.133-x$				x		$0.243+x$

$[HF]_i > 100K_a$, $0.133 - x \approx 0.133$
$[F^-]_i > 100K_a$, $0.243 + x \approx 0.243$

$$K_a = \frac{[H_3O^+][F^-]}{[HF]}$$

$$7.2\times10^{-4} = \frac{x(0.243)}{0.133}$$

$$0.234x = 0.133(7.2\times10^{-4}) = 9.576\times10^{-5}$$

$$x = \frac{9.576\times10^{-5}}{0.234} = 3.941\times10^{-4} = [H_3O^+]$$

$$pH = -\log_{10}(3.941\times10^{-4})$$

$$pH = 3.40$$

17.2 Buffers

Buffer pH – The Henderson–Hasselbalch Equation

$$HA(aq) + H_2O(l) \Leftrightarrow H_3O^+ + A^-$$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

$$[H_3O^+] = K_a\left(\frac{[HA]}{[A^-]}\right)$$

$$= K_a \frac{[\text{Weak acid}]}{[\text{Conjugate base}]}$$

$$= K_a \frac{[\text{Buffer acid}]}{[\text{Buffer base}]}$$

$$-\log_{10}[H_3O^+] = -\log_{10}K_a - \log_{10}\frac{[\text{Buffer acid}]}{[\text{Buffer base}]}$$

$$pH = pK_a + \log_{10}\frac{[\text{Buffer base}]}{[\text{Buffer acid}]}$$ *1

*1: BEWARE: This formula works only if:
[Buffer base] > 100 Ka
[Buffer acid] > 100 Ka
Yet at no stage is this checke!

The ICE method is much safer as it check the above to see if you can use the short cut or whether you need to solve the quadratic.

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}

$$pH = pK_a + \log_{10} \frac{[\text{Buffer base}]}{[\text{Buffer acid}]}$$

[Buffer base] = $[F^-]$ = 0.243 [Buffer acid] = [HF] = 0.133

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

$$pH = 3.143 + \log_{10} 1.827$$

$$= 3.143 + 0.262$$

$$pH = 3.40$$

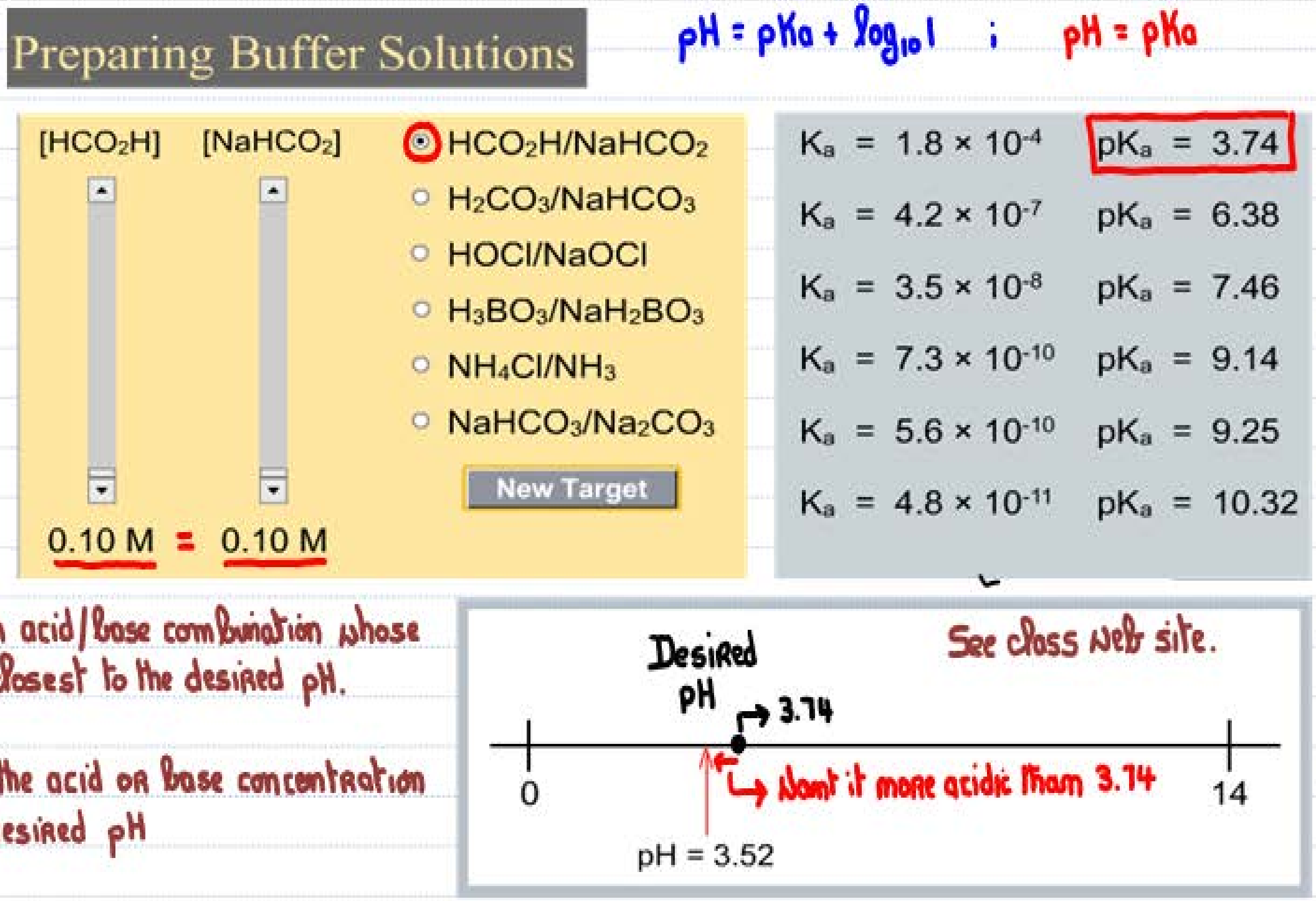
17.2 Buffers

Making Buffer Solutions

$$pH = pKa + \log_{10}\frac{[\text{Buffer base}]}{[\text{Buffer acid}]}$$

[Buffer base] = [Buffer acid]

$pH = pKa + \log_{10} 1$; $pH = pKa$



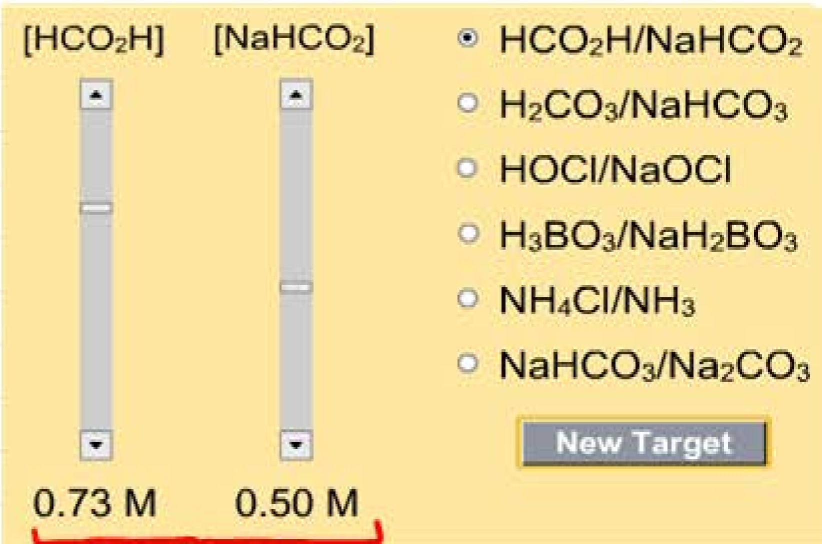
I) Chose an acid/base combination whose pKa is closest to the desired pH.

II) Adjust the acid or base concentration to get desired pH

17.2 Buffers

Buffer Capacity

Preparing Buffer Solutions



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

Buffer Capacity.

i) Maximum amount of H_3O^+ that can be removed = $[HCO_2^-]$ = 0.50 M.

ii) Maximum amount of OH^- that can be removed = $[HCO_2H]$ = 0.73 M