17.2 Buffers

How do they Resist Drastic pH Change

Buffer: 1M CH_3CO_2H / 1M $CH_3CO_2^-$
Addition of Strong Acid – H_3O^+

CH_3CO_2H / $CH_3CO_2^-$
Weak Acid — Conjugate Base
'Buff Acid' — 'Buffer Base'

$H_3O^+ + CH_3CO_2^- = H_2O(l) + CH_3CO_2H(aq)$
'Buffer Base' — 'Buffer Acid'

Overall Changes:

$[CH_3CO_2^-]$: ↓ ... Reacts with the added H_3O^+

$[CH_3CO_2H]$: ↑ ... Product of the reaction that removes the H_3O^+

$[H_3O^+]$: ↑ ... Slight increase ... the result of the ['Buffer Acid'] ↑

pH: ↓ ... From the slight increase in $[H_3O^+]$

17.2 Buffers

How do they Resist Drastic pH Change

Buffer: 1M CH_3CO_2H / 1M $CH_3CO_2^-$
Addition of Strong Base – OH^-

CH_3CO_2H / $CH_3CO_2^-$
Weak Acid — Conjugate Base
'Buffer Acid' — 'Buffer Base'

$OH^- + CH_3CO_2H(aq) = H_2O(l) + CH_3CO_2^-$
'Buffer Acid' — 'Buffer Base'

OVERALL CHANGES

$[CH_3CO_2H]$: ↓ ... Reacts with the added OH^-.

$[CH_3CO_2^-]$: ↑ ... Product of the reaction that removed OH^-

$[OH^-]$: ↑ ... Slight increase ... a result of the [Buffer Base] ↑

pH : ↑ ... From the slight increase in $[OH^-]$

ers
do they Resist Drastic pH Change

A buffer solution made from **HF** and **KF** has a pH = 2.84.
Addition of OH^- will cause –

a) Increase or Decrease significantly
b) Increase slightly
c) Decrease slightly
d) Increase
e) Decrease

a) pH  b) ✓ ... adding a base

b) [HF]  e) ✓ ... used in the removal of OH^-

$OH^- + HF(aq) = H_2O(l) + F^-$

'Buffer acid' 'Buffer base'

17.2 Buffers

Identifying Buffers

Which of the following aqueous solutions are good buffer systems?

- [x] 0.34 M ammonium bromide (NH_4^+) + 0.36 M ammonia (NH_3)
- [ ] 0.22 M nitric acid (Strong acid) + 0.16 M potassium nitrate
- [x] 0.32 M nitrous acid (HNO_2) + 0.21 M potassium nitrite (NO_2^-)
- [ ] 0.18 M barium hydroxide (Strong Base) + 0.21 M barium bromide
- [x] 0.14 M hydrofluoric acid (HF) + 0.20 M sodium fluoride (F^-)

2 To Look Out For:

1.) Strong Acid + Weak Base
2.) Weak Acid + Strong Base

} When the concentration of the Strong is < concentration of the Weak.

$$0.07\,M\ NaOH(aq) + 0.14\,M\ HF(aq) = H_2O(l) + NaF(aq)$$

0.07 M HF (aq) — Weak Acid

0.07 M NaF(aq) → 0.07 M Na^+ + 0.07 M F^- (Conjugate Base of HF(aq))

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.
Ka HF = 7.2×10^{-4}

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

	$HF_{(aq)}$	+ $H_2O_{(l)}$ ⇔	H_3O^+	+ F^-
I	0.133		0	0.243
C	$-x$		x	x
E	$0.133 - x$		x	$0.243 + x$

$[HF]_i > 100(7.2 \times 10^{-4})$, $0.133 - x \approx 0.133$

$[F^-]_i > 100(7.2 \times 10^{-4})$, $0.243 + x \approx 0.243$

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

$$7.2 \times 10^{-4} = \frac{x(0.243)}{0.133}$$

$$0.243x = 0.133(7.2 \times 10^{-4})$$

$$x = \frac{0.133(7.2 \times 10^{-4})}{0.243} = 3.94 \times 10^{-4} = [H_3O^+]$$

$$pH = -\log_{10}(3.94 \times 10^{-4}) = 3.40$$

17.2 Buffers

Buffer pH – The Henderson–Hasselbalch Equation

$$HA_{(aq)} + H_2O_{(l)} \Leftrightarrow H_3O^+ + A^-$$

'Buffer Acid' — 'Buffer Base'

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

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

$$= K_a\left(\frac{[\text{Weak acid}]}{[\text{Conjugate base}]}\right)$$

$$= K_a\left(\frac{[\text{Buffer Acid}]}{[\text{Buffer Base}]}\right)$$

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

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

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

* **BEWARE:-** This formula works only if:
[Buffer Base] > 100 K_a and:
[Buffer Acid] > 100 K_a
Yet at no stage is this checked.

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