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ICE ICE, BABY, ICE ICE

$$pH = pK_a + \log\left(\frac{[base]}{[acid]}\right)$$
$$pK_a = -\log(1.8 \times 10^{-5}) = 4.74$$
$$pH = 4.74 + \log\left(\frac{[base]}{[acid]}\right)$$

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ICE ICE, BABY, ICE ICE

HOAc + H₂O ↔ OAc⁻ + H₃O⁺

I	0.100 M	-	0.100 M	0
C	-x	-	+x	+x
E	0.100-x	-	0.100+x	x

$$pH = 4.74 + \log\left(\frac{[base]}{[acid]}\right)$$
$$pH = -\log(x) = 4.74 + \log\left(\frac{0.100 + x}{0.100 - x}\right)$$

Henderson-Hasselbach doesn't help me much!!!

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The Equilibrium Constant Expression

Except, what if we make our assumption?

$x \ll 0.100$

$$pH = 4.74 + \log\left(\frac{0.100 + x}{0.100 - x}\right)$$
$$pH = 4.74 + \log\left(\frac{0.100}{0.100}\right) = 4.74$$

Henderson-Hasselbach is just the solution to your ICE chart if YOU MAKE THE ASSUMPTION!

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I do each of them,
one at a time, from
biggest K to smallest
K!

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My TWO (or 4) reactions!

$NH_3(aq) + HCl(aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$

$NH_3(aq) + H_2O(l) \leftrightarrow NH_4^+(aq) + OH^-(aq)$

$HCl(aq) + H_2O(l) \rightarrow Cl^-(aq) + H_3O^+(aq)$

$H_2O(l) + H_2O(l) \rightarrow H_3O^+(aq) + OH^-(aq)$

Do I know the K's?

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My TWO (or 4) reactions!

$NH_3(aq) + HCl(aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$

$NH_3(aq) + H_2O(l) \leftrightarrow NH_4^+(aq) + OH^-(aq)$
 $K_a = 1.8 \times 10^{-5}$

$HCl(aq) + H_2O(l) \rightarrow Cl^-(aq) + H_3O^+(aq)$
 $K_a = \text{HUGE, nearly infinite}$

$H_2O(l) + H_2O(l) \rightarrow H_3O^+(aq) + OH^-(aq)$
 $K_w = 1.0 \times 10^{-14}$

What about the first reaction?

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What about the first reaction?

$$NH_3(aq) + HCl(aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$$

It is NOT NOT NOT a K_a !
 How do I know?
 No water, no K_a !
 It's actually a combo of the other 3 reactions!

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What about the first reaction?

$$NH_3(aq) + HCl(aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$$

If I add the first 2 and subtract the 3rd one, I get...

$$NH_3(aq) + H_2O(l) \leftrightarrow NH_4^+(aq) + OH^-(aq)$$

$$HCl(aq) + H_2O(l) \rightarrow Cl^-(aq) + H_3O^+(aq)$$

$$NH_3(aq) + H_2O(l) + HCl(aq) + H_2O(l) \rightarrow NH_4^+(aq) + OH^-(aq) + Cl^-(aq) + H_3O^+(aq)$$

$$H_3O^+(aq) + OH^-(aq) \rightarrow H_2O(l) + H_2O(l)$$

$$NH_3(aq) + H_2O(l) + HCl(aq) + H_2O(l) + H_2O(l) + OH^-(aq) \rightarrow NH_4^+(aq) + OH^-(aq) + Cl^-(aq) + H_3O^+(aq) + H_2O(l) + H_2O(l)$$

$$\rightarrow NH_4^+(aq) + Cl^-(aq) + H_2O(l)$$

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What about the first reaction?

$$NH_3(aq) + HCl(aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$$

So, I could get the K by combining the K's from the other reactions!

$$K = \frac{K_a(HCl) \times K_b(NH_3)}{K_w} = \frac{FB \times 1.8 \times 10^{-5}}{1.0 \times 10^{-14}}$$

$$= FB \times 1.8 \times 10^9 = FB \times FB = SFB$$

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Calculating buffer capacity

I mix 250 mL of 0.5 M HOAc and 500 mL of 0.5 M NaOAc. What is the pH of the buffer?
What is the buffer capacity?
($K_a = 1.8 \times 10^{-5}$)

$$pH = pK_a + \log\left(\frac{[base]}{[acid]}\right)$$

0.25 L * 0.5 M = 0.125 moles HOAc
0.500 L * 0.5 M = 0.25 moles OAc⁻

$$pH = 4.74 + \log\left(\frac{[0.25\text{moles}]}{[0.125\text{moles}]}\right)$$

$$pH = 5.04$$

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HOAc + H₂O = OAc⁻ + H₃O⁺

$$pH = pK_a + \log\left(\frac{[OAc^-]}{[HOAc]}\right)$$

If the pH=5.04, it is "balanced" until pH = 4.04 or pH = 6.04

Let's do pH=4.04 1st.

Add HCl (acid lowers pH)
HCl + OAc⁻ = HOAc + Cl⁻

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HOAc + H₂O ↔ OAc⁻ + H₃O⁺

I	0.125 mol	-	0.250 mol	0
C	+ acid		-acid	
I	0.125 + acid		0.250 - acid	

$$pH = pK_a + \log\left(\frac{[base]}{[acid]}\right) = 4.74 + \log\left(\frac{0.250 - x}{0.125 + x}\right)$$

$$4.04 = 4.74 + \log\left(\frac{0.250 - x}{0.125 + x}\right)$$

$$-0.70 = \log\left(\frac{0.250 - x}{0.125 + x}\right)$$

$$10^{-0.70} = 0.1995 = \frac{0.250 - x}{0.125 + x}$$

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$$0.1995 = \frac{0.250 - x}{0.125 + x}$$
$$0.025 + 0.1995 = 0.250 - x$$
$$1.1995x = 0.2475$$
$$X = 0.206 \text{ moles Acid}$$

So, if I add 0.206 moles of acid to my buffer, it is "unbuffered".

Buffer capacity = moles acid/L solution
Buffer capacity = 0.206 moles/0.75 L = 0.275M acid

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How much base, I can add?

$$6.04 = 4.74 + \log\left(\frac{0.250 + x}{0.125 - x}\right)$$
$$1.30 = \log\left(\frac{0.250 + x}{0.125 - x}\right)$$
$$10^{1.30} = \frac{0.250 + x}{0.125 - x}$$
$$19.95 = \frac{0.250 + x}{0.125 - x}$$
$$2.49 - 19.95x = 0.25 + x$$
$$2.24 = 20.95x$$
$$X = 0.1 \text{ moles}$$

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Buffer capacity

$$\text{Buffer capacity} = \frac{0.1 \text{ mol base absorbed}}{0.75 \text{ L}} = 0.133 \text{ M}$$

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