1. Buffer capacity is $\approx \pm$ 1 pH unit away from pKa. of the conj. acid

CH_3COOH $K_a = 1.8 \times 10^{-5}$ $pK_a = 4.74$

H-Cl is a strong acid, so we would not use it as a buffer.

NH_4^+ $K_a = 5.6 \times 10^{-10}$ $pK_a = 9.25$ ✓ would work to buffer around pH 9

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2. Na_2HPO_4 / NaH_2PO_4 equimolar concentrations.

[conj. base] = [conj. acid]

$[HPO_4^{2-}] = [H_2PO_4^-]$

$$pH = pK_a + \log\left[\frac{[\text{conj base}]}{[\text{conj acid}]}\right] = pK_a + \cancel{\log(1)}^{0} = pK_a$$

$$pH = pK_a = -\log(pK_a) = -\log(6.2 \times 10^{-8})$$

looked up Ka of H_2PO_4 in table

$pH = 7.21$

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3. $pH = -\log[H^+]$

a) $pH = -\log(4.5 \times 10^{-3}) = 2.3$

pKa b/w 1.3 + 3.3

Ka b/w $10^{-1.3}$ + $10^{-3.3}$

0.05 + 5×10^{-4}

$H_3PO_4 / H_2PO_4^-$ looks good

$pK_a = 2.12$

b.) $pH = -\log(9.7 \times 10^{-14}) = 10.0$

HCO_3^- / CO_3^{2-} $pK_a = 10.32$

4. Acid-base pair is NH_4^+ / NH_3 $pKa = 9.25$

$$pH = pka + \log \left[\frac{NH_3}{NH_4^+}\right]$$ final pH is 9.00

$$PH - pka = \log \left(\frac{NH_3}{NH_4^+}\right)$$

$$9.00 - 9.25 = \log \frac{[NH_3]}{[NH_4^+]} \rightarrow 10^{-0.25} = \frac{[NH_3]}{[NH_4^+]}$$

since the volume is the same; we can see that the molar ratio of NH_3 to NH_4 is $10^{-0.25}$.

moles NH_3: $\frac{0.10 \text{ moles } NH_3}{1L} \times 0.5L = 0.05 \text{ mol } NH_3$

$$10^{-.25} = \frac{0.05 \text{ mol } NH_3}{x \text{ mol } NH_4^+}$$ $x = 0.0889 \text{ mol } NH_3^+$

$$\frac{0.0889 \text{mol } NH_3Cl}{1} \times \frac{53.49 g}{mol} \approx \boxed{4.8g\ NH_4^+Cl \text{ needed}}$$

5. Initial condition

$$pH = pka + \log \frac{[HPO_4^{2-}]}{[H_2PO_4^{1-}]}$$

because $[H_2PO_4^{-1}] = [HPO_4^{2-}]$

$pH = pka = 7.21$ (looked up pka from table)

if $pH = 7.21$, what is initial $[H_3O^+]$?

$$10^{-7.21} = [H_3O^+] = 6.2 \times 10^{-8} M$$

~~Set up ICE table~~

$[H_3O^+] = 6.2 \times 10^{-8} \frac{\text{moles}}{L}$ → since this is a 2.00L solution we have $\boxed{1.24 \times 10^{-7} \text{ moles } H_3O^+}$

-add $\frac{0.425 g \text{ NaOH}}{1} \times \frac{\text{mole}}{40 g} = 0.0106$ mol NaOH → $\boxed{0.0106 \text{ mol } OH}$

$NaOH + H_3O^+ \rightarrow Na^+ + 2 H_2O$ $[H_3O^+]$ is obviously the limiting reagent.

thus remaining amount of OH = initial amount − consumed amount

$0.0106 \text{ mol OH} - 1.2 \times 10^{-7} \text{ mol OH} \approx 0.0106 \text{ mol OH}$

thus $pOH = -\log\left(\frac{0.0106 \text{ mol OH}}{2L}\right) = 2.27$

$\therefore pH = 11.73$.

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6. $pH = pKa + \log\frac{[\text{propanoate}]}{[\text{propanoic acid}]}$ $pKa = 4.85$

$pH = 4.85 + \log\left(\frac{0.3}{0.2}\right) = 5.02$

-Add HCl. what is the limiting reagent?   or $HCl + \text{propanoate} \rightleftharpoons \text{propanoic acid} + Cl^-$

moles HCl: $\frac{0.10 \text{ moles HCl}}{1 \text{ Liter}} \times \frac{1 \times 10^{-3} L}{1} \rightarrow 1 \times 10^{-4} \text{ moles } [H^+]$

moles propanoate: $\frac{0.30 \text{ moles}}{\text{Liter}} \times \frac{0.01 L}{1} \rightarrow 0.003 \text{ moles } [H^+]$

$\therefore [H^+]$ is the lim. reagent.

time to set up ICE

$[H^+]_0 = \frac{1\times10^{-4} \text{ moles } H^+}{0.011 L} = 0.0091 M\ H^+$ $[HA]_0 = 0.182 M$ $[A^-]_0 = 0.273 M$

	A^-	$+ H_3O^+ \rightleftharpoons$	$HA + H_2O$
I	0.273	0.0091	0.182
C	-0.0091	-0.0091	+0.0091
E	0.2639	0	0.1911

$K_a = 1.4\times10^{-5}$

$$pH = pKa + \log\left(\frac{0.2639}{0.1911}\right)$$

$$pH = 4.85 + 0.14 = 4.99$$

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7. moles HCl : $\frac{0.150M}{1} \times 0.050L = 0.0075 \text{ mol HCl}$

moles NaOH : (a) $\frac{0.150M}{1} \times 0.025L = 0.00375 \text{ mol NaOH}$

*NaOH is the limiting reagent → 0.00375 moles HCl left

$[H^+] = \frac{0.00375 \text{ moles}}{(0.075 L)} = 0.05M\ H_3O^+$

pH = 1.3

(b) $0.150M \times 0.0501L = 0.00752 \text{ mol NaOH}$

*HCl is Lim. Reagent! 2×10^{-5} moles OH^- left

$[OH^+] = \frac{2\times10^{-5} \text{ moles}}{0.1001} = 1.998\times10^{-4} \longrightarrow$ pOH = 3.7

pH = 10.3

8. 30.00-mL 0.100 M HA add 0.1 M NaOH (10, 30, 40)

$$\frac{0.1 \text{ moles HA}}{L} \times \frac{0.030 L}{} = 0.003 \text{ moles HA}$$

$$\frac{0.1 \text{ moles NaOH}}{L} \times \frac{0.010 L}{1} = 0.001 \text{ moles NaOH}$$ * lim reagent

$[OH]_0$ $$\frac{0.001 \text{ moles NaOH}}{0.04} = 0.025 M \text{ NaOH} \rightarrow 0.025 M \ OH^-$$

$[HA]_0$ $$\frac{0.003 \text{ moles HA}}{0.04} = 0.075 M \text{ HA}$$

	HA	NaOH	→ A^-	H_2O
I	0.075 M	+ 0.025 M		
C	-0.025	-0.025	+0.025	
E	0.05	0	0.025	

$\boxed{\text{Have } HA \approx A^- \therefore \text{buffer}}$

$$pH = pKa + \log \frac{A^-}{HA} = 4.2 - 0.30 = 3.9$$

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30 mL

$$\frac{0.1 \text{ moles NaOH}}{L} \times 0.03 L = 0.003 \text{ moles OH}$$

$[OH]_0$ $$\frac{0.003 \text{ moles OH}}{0.06 L} = 0.05 M \text{ OH}$$

$[HA]_0 =$ $$\frac{0.003 \text{ moles HA}}{0.06 L} = 0.05 M \text{ HA}$$

	HA	OH	⇄ A^-	H_2O
I	0.05 M	0.05 M	0	
C	-0.05	-0.05	+0.05	
E	0	0	0.05 M	

Salt calculation!

0.05 M A^-, what is pH?

$K_a = 6.3 \times 10^{-5}$

$K_b = \frac{K_w}{K_a} = \frac{1 \times 10^{-14}}{6.3 \times 10^{-5}} = 1.5 \times 10^{-10}$

	A^-	$+ H_2O$	$\rightleftharpoons$	HA	$+ OH^-$
I	0.05				
C	$-x$			$+x$	$+x$
E	$0.05-x$			x	x

$$K_b = \frac{[HA][OH]}{[A^-]} = \frac{x^2}{0.05-x} = 1.5 \times 10^{-10}$$

$$\xrightarrow[\text{cut}]{\text{short}} (0.05)(1.5 \times 10^{-10}) = x^2 \qquad x = 2.82 \times 10^{-6}$$

$$pOH = -\log(2.82 \times 10^{-6}) = 5.5$$

$$pH = 8.45$$