

3) Since these are all buffer solutions, you can use the Henderson-Hasselbalch equation

$$a) \text{ pH} = \text{pKa} + \log \frac{[\text{Base}]}{[\text{Acid}]}$$

$$\text{pH} = (-\log 1.8 \times 10^{-5}) + \log \frac{0.25}{0.10}$$

$$\text{pH} = 4.74 + 0.40 = 5.14$$

$$b) \text{ pH} = \text{pKa} + \log \frac{[\text{Base}]}{[\text{Acid}]}$$

$$\text{pH} = 4.74 + \log \frac{0.10}{0.25}$$

$$\text{pH} = 4.34$$

$$c) \text{ pH} = \text{pKa} + \log \frac{[\text{Base}]}{[\text{Acid}]}$$

$$\text{pH} = 4.74 + \log \frac{0.20}{0.080}$$

$$\text{pH} = 5.14$$

$$d) \text{ pH} = \text{pKa} + \log \frac{[\text{Base}]}{[\text{Acid}]}$$

$$\text{pH} = 4.74 + \log \frac{0.080}{0.20}$$

$$\text{pH} = 4.34$$

$$33) [HC_7H_5O_2] = \frac{21.5g HC_7H_5O_2 \frac{1mol}{122.12g}}{0.2000L} = 0.880M$$

$$[C_7H_5O_2^-] = \frac{327g NaC_7H_5O_2 \frac{1mol NaC_7H_5O_2}{144.10g NaC_7H_5O_2} | 1mol C_7H_5O_2^-}{0.2000L} = 1.31M$$

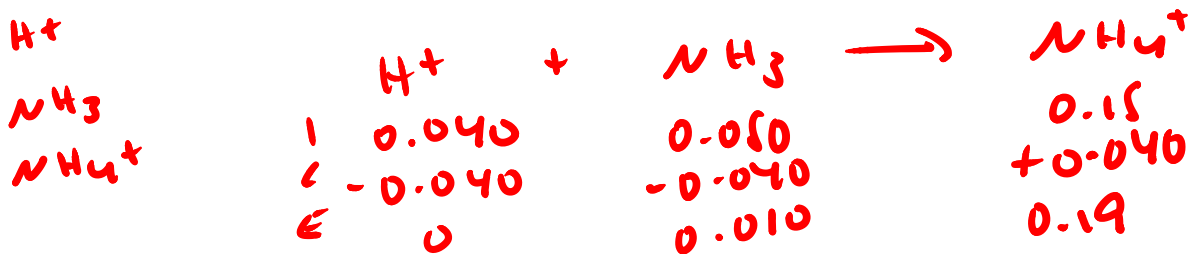
* since it's a buffer, use Henderson-Hasselbalch

$$pH = pK_a + \log \frac{[C_7H_5O_2^-]}{[HC_7H_5O_2]}$$

$$pH = -\log(6.4 \times 10^{-5}) + \log \frac{1.31}{0.880}$$

$$pH = 4.19 + 0.173 = 4.36$$

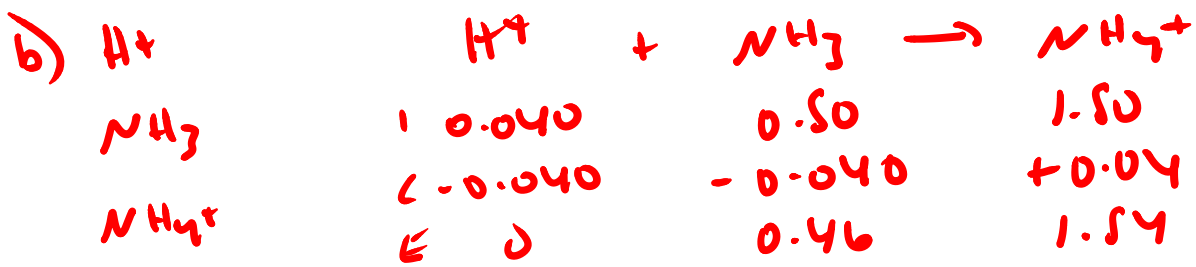
$$35 a) [H^+] = \frac{0.010mol}{0.2500L} = 0.040M$$



$$pH = pK_a + \log \frac{[Base]}{[Acid]}$$

$$= -\log(5.6 \times 10^{-10}) + \log \frac{0.010}{0.19}$$

$$= 9.25 - 1.29 = 7.97$$



$$pH = pK_a + \log \frac{[Base]}{[Acid]}$$

$$= 9.25 + \log \frac{0.46}{1.54} = 8.73$$

both buffers have the same initial pH however their buffer capacities are different. (B) has a larger buffer capacity since it has larger buffer components. Buffers with larger buffer capacities can absorb more added H⁺ or OH⁻

3) $pH = pK_a + \log \frac{[Base]}{[Acid]}$

* Since buffer components are in same volume,

$$\frac{[Base]}{[Acid]} = \frac{\text{mol Base}}{\text{mol Acid}}$$

$$5.00 = -\log(1.8 \times 10^{-5}) + \log \frac{\text{mol Base}}{\text{mol Acid}}$$

mole H₂C₂O₄ = (0.5000 L)(0.200 M) = 0.100 mo

$$0.26 = \log \frac{x}{0.100} \quad 10^{-0.26} = \frac{x}{0.100}$$

$$x = 0.18 \text{ mole NaC}_2\text{H}_3\text{O}_2 \quad \frac{82.03 \text{ g}}{1 \text{ mol}} = 15 \text{ g NaC}_2\text{H}_3\text{O}_2$$



$$K_a = \frac{1.0 \times 10^{-4}}{1.7 \times 10^{-9}} = 5.9 \times 10^{-6}$$

$$pK_a = -\log(5.9 \times 10^{-6}) = 5.23$$

$$pH = pK_a + \log \frac{[\text{C}_5\text{H}_5\text{N}]}{[\text{C}_5\text{H}_5\text{NH}^+]}$$

$$a) 4.8 = 5.23 + \log \frac{[\text{Base}]}{[\text{Acid}]}$$

$$10^{-0.43} = \frac{[\text{Base}]}{[\text{Acid}]} = 0.37$$

$$b) 5.00 = 5.23 + \log \frac{[\text{Base}]}{[\text{Acid}]}$$

$$10^{-0.23} = \frac{[\text{Base}]}{[\text{Acid}]} = 0.59$$

$$c) 5.23 = 5.23 + \log \frac{[\text{Base}]}{[\text{Acid}]}$$

$$10^{-0.00} = \frac{[\text{Base}]}{[\text{Acid}]} = 1.0$$

$$d) 5.50 = 5.23 + \log \frac{[\text{Base}]}{[\text{Acid}]}$$

$$10^{0.27} = \frac{[\text{Base}]}{[\text{Acid}]} = 1.9$$

41. A best buffer has large and equal quantities of weak acid and conjugate base. Since $[\text{Acid}] = [\text{Base}]$ for the best buffer...

$$\text{pH} = \text{pK}_a + \log \frac{[\text{Base}]}{[\text{Acid}]} = \text{pK}_a + 0$$

*** for the best buffer solution, pH is equal to pKa (or close to it)

If we want a buffer with a pH = 7.00, we want a pKa as close to 7 as possible (Ka around 1.0×10^{-7})

HOCl has a $K_a = 3.5 \times 10^{-8}$ and $\text{pK}_a = 7.46$

$$7.00 = 7.46 + \log \frac{[\text{Base}]}{[\text{Acid}]} = 10^{-0.46} = 0.35$$

So any $[\text{OCl}^-]/[\text{HOCl}]$ value with a ratio of 0.35 to 1 will work