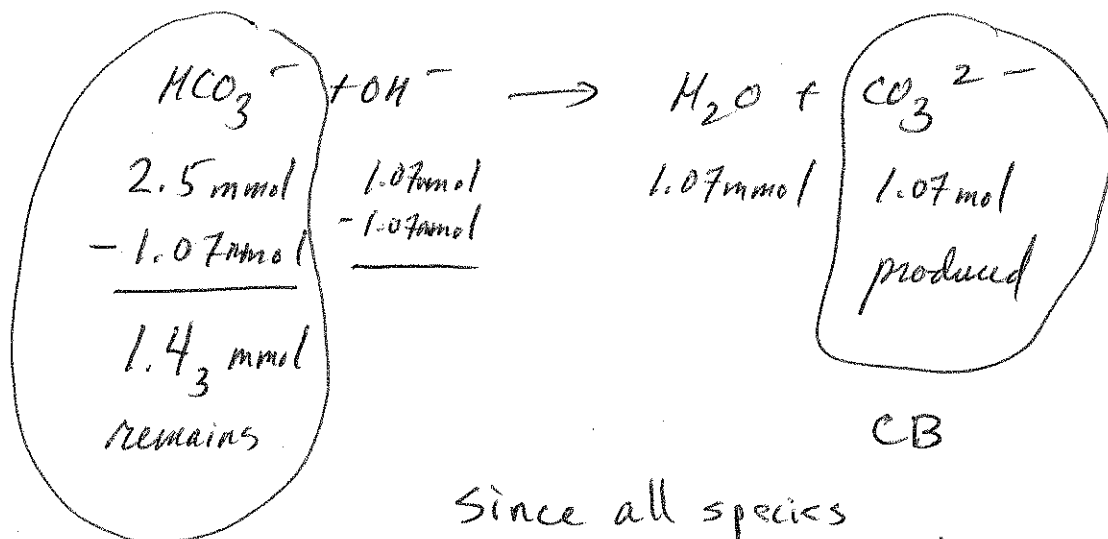


$$\text{NaHCO}_3 \quad 50.0 \text{ mL}, 0.050 \text{ M} = 2.5 \text{ mmol HCO}_3^-$$

$$\text{NaOH} \quad 10.7 \text{ mL}, 0.10 \text{ M} = 1.07 \text{ mmol OH}^-$$



WA

CB

Since all species are in the same volume, one can use mmols rather than molarity below.

$$a) \quad K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} \Rightarrow [\text{H}_3\text{O}^+] = \frac{K_a [\text{HA}]}{[\text{A}^-]}$$

$$[\text{H}_3\text{O}^+] = (4.69 \times 10^{-11}) \left( \frac{1.4_3}{1.07} \right) = 6.2_7 \times 10^{-11}$$

$$\text{pH} = -\log(6.2_7 \times 10^{-11}) = 10.20 \quad \checkmark$$

Using HH

$$\begin{aligned} \text{pH} &= \text{p}K_a + \log \frac{[\text{A}^-]}{[\text{HA}]} = 10.33 + \log \left( \frac{1.07}{1.4_3} \right) \\ &= 10.33 + \log(0.74_8) \\ &= 10.33 + -0.12_6 \\ &= 10.22 \quad \checkmark \end{aligned}$$

b) pH of buffer = 10.22

new pH = 10.22 - 0.07 units = 10.15

↑  
why  
the negative  
sign?

why not  
positive?

$$10.15 = 10.33 + \log \left( \frac{\text{mmol } A^- - \text{mmol } H^+}{\text{mmol } HA + \text{mmol } H^+} \right)$$

$$10.15 = 10.33 + \log \left( \frac{0.441 - z}{0.589 + z} \right)$$

25 mL sample of buffer

$$A^- : \frac{1.07 \text{ mmol}}{60.7 \text{ mL}} = \frac{x}{25.0 \text{ mL}}$$

$$x = 0.441 \text{ mmol } A^-$$

$$HA : \frac{1.43 \text{ mmol}}{60.7} = \frac{y}{25.0 \text{ mL}}$$

$$y = 0.589 \text{ mmol } HA$$

$$10^{-0.18} = \frac{0.441 - z}{0.589 + z} \Rightarrow 0.661 = \frac{0.441 - z}{0.589 + z}$$

$$0.389 + 0.661z = 0.441 - z$$

$$1.661z = 0.052$$

$$z = 0.031 \text{ mmol HCl needed}$$

1.14 g HCl

$$\times 36.46 \frac{\text{g}}{\text{mol}}$$

(1)

a) What is the pH of an aqueous solution that is 0.551 M phenol and 0.377 sodium phenolate?

b) Explain your answer in part a.

c) What will be the pH of the solution in part a after 0.100 mol of HCl gas is dissolved in 1.00 L of the solution? Assume the volume of the solution does not change.

d) Explain your answer in part c.

e) What will be the pH of the solution in part a after 0.125 mol of solid KOH is dissolved in 1.00 L of the solution? Assume the volume of the solution does not change.

f) Explain your answer in part e.

g) How many grams of HCl can be dissolved in the original solution (part a) while still being considered a buffer? This is called buffer capacity.

h) The assumption is made in parts c and e that the volume of the resulting solution does not change. While this is a reasonable assumption, what does happen to the volume when either HCl or NaOH is added? Explain.

a)  $K_a \text{ phenol} = 1.6 \times 10^{-10}$

$$K_b \text{ phenolate} = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{1.6 \times 10^{-10}} = 6.25 \times 10^{-5}$$

Since  $K_b > K_a$  the pH of the buffer will be  $pK_b \pm 1 = 9.80 \pm 1$

or  $pOH = pK_b \pm 1$

$$\begin{aligned} pH &= pK_a + \log \frac{[A^-]}{[HA]} \\ &= 9.80 + \log \frac{(0.377)}{(0.551)} \\ &= 9.80 + -0.165 \\ &= 9.64 \end{aligned}$$

$$\begin{aligned} pOH &= pK_b + \log \frac{[HA]}{[A^-]} \\ pOH &= 4.20 + \log \frac{(0.551)}{(0.377)} \\ &= 4.20 + 1.65 \times 10^{-1} \\ &= 4.36 \\ pH &= 14.00 - 4.36 = 9.64 \end{aligned}$$

b)

c) Since there is one Liter of sol'n  $\rho$  of a monoprotic acid  
moles of HA = 0.551  
moles of A<sup>-</sup> = 0.377  
added moles of H<sub>3</sub>O<sup>+</sup> (from HCl reacting w/ H<sub>2</sub>O) = 0.100 mol

HA ↑ 0.651 mol

A<sup>-</sup> ↓ 0.277 mol

$$pH = 9.80 + \log \frac{(0.277)}{(0.651)} = 9.43$$

d) Adding strong acid (a small amount) increases HA at the expense of A<sup>-</sup> but the ratio A<sup>-</sup>/HA is still within the same order of magnitude

p) Similar to c except that  $\text{pH} > 9.80$

$$\text{pH} = 9.80 + \log \frac{5.02 \times 10^{-1}}{4.26 \times 10^{-1}} = 9.87$$

f) similar to d but now  $\text{A}^- \uparrow$  and  $\text{HA} \downarrow$   
due to the addition of strong base

g) Since HCl is added the limit will be  $\text{pH}$   
which is 8.80.

$$8.80 = 9.80 + \log \frac{\overset{\text{moles}}{\text{A}^- - x}}{\underset{\text{moles}}{\text{HA} + x}}$$

$$-1.00 = \log \left( \frac{\text{A}^- - x}{\text{HA} + x} \right)$$

$$0.1 = \frac{\text{A}^- - x}{\text{HA} + x}$$

$$0.1 \text{ HA} + 0.1x = \text{A}^- - x$$

$$1.1x = \text{A}^- - 0.1 \text{ HA}$$

h) I'll let you think  
about this. What does it mean  
for a substance to dissolve?

This assumes 1 L of soln  
w/ the original concentrations

$$1.1x = 0.377 - (0.1)(0.551)$$

$$x = \frac{0.222}{1.1} \text{ moles HCl}$$

$$0.202 \text{ moles HCl} \times \frac{36.46 \text{ g}}{\text{mol}} = 7.34 \text{ g}$$

$$x = 0.293 \text{ moles HCl}$$

$$0.293 \text{ moles} \times \frac{36.46 \text{ g}}{\text{mol}} = 10.7 \text{ g HCl}$$

(2) Some liquid household bleaches are about 5% by weight aqueous solutions of sodium hypochlorite, NaOCl, which completely dissociates to its ions in the solution. Hypochlorous acid, HOCl<sub>(aq)</sub>, has a pK<sub>a</sub> of 7.5.

(a) What is the molarity of sodium hypochlorite in bleach? Show how you get your answer.

(b) Write the reaction that hypochlorite anion, ClO<sup>-</sup><sub>(aq)</sub>, undergoes with water. Write the equilibrium constant expression for this reaction. What is the usual symbol we use to designate the equilibrium constant for a reaction like this? What is the numeric value of this equilibrium constant? Explain your reasoning.

(c) What is the pH of a liquid bleach solution? Explain.

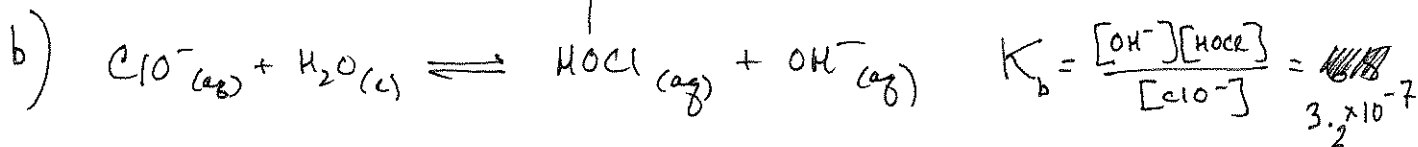
(d) If you wish to change the pH of bleach solution to 6.5, should you add sodium hydroxide or hydrochloric acid? Explain your reasoning.

(e) What is the ratio of the conjugate base to the conjugate acid in a bleach solution whose pH has been adjusted to 6.5? Explain.

$$pK_b = -pK_a + 14.0 = -7.5 + 14.0 = 6.5$$

a) 
$$\frac{5 \text{ g NaOCl}}{100 \text{ g solution}} \times \frac{1 \text{ g soln}}{1 \text{ mL soln}} \times \frac{1000 \text{ mL}}{1 \text{ L}} \times \frac{1 \text{ mol NaOCl}}{74.44 \text{ g NaOCl}} = 0.67 \text{ M NaOCl}$$

↑  
assume density ≈ 1 g/mL



c) 
$$\frac{[\text{OH}^-][\text{HOCl}]}{[\text{ClO}^-]} = K_b \quad \frac{x^2}{0.67 - x} = 3.2 \times 10^{-7} \quad x = [\text{OH}^-] = 4.6 \times 10^{-4} \text{ M}$$

pOH = 3.3  
pH = 14.0 - 3.3 = 10.7

d) pH decreases, so add HCl.  
HCl will protonate OCl<sup>-</sup> to form HOCl, resulting in a buffer.

e) 
$$6.5 = 7.5 + \log\left(\frac{[\text{OCl}^-]}{[\text{HOCl}]}\right) \quad \text{pH} \rightarrow \text{ratio of } \frac{\text{conj base}}{\text{acid}} = \frac{0.1}{1} = \frac{[\text{OCl}^-]}{[\text{HOCl}]}$$

$$7.5 = 6.5 + \log\left(\frac{[\text{HOCl}]}{[\text{OCl}^-]}\right) \quad \text{pOH}$$