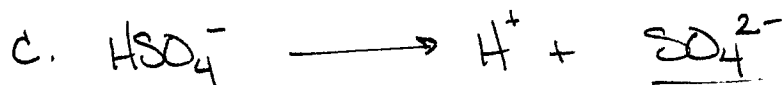
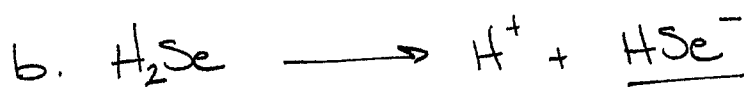
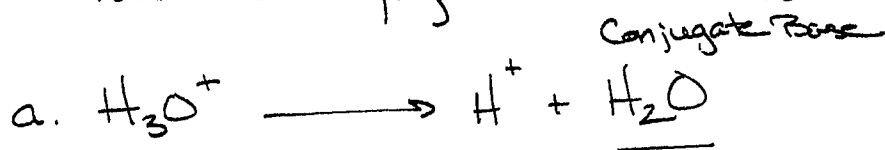
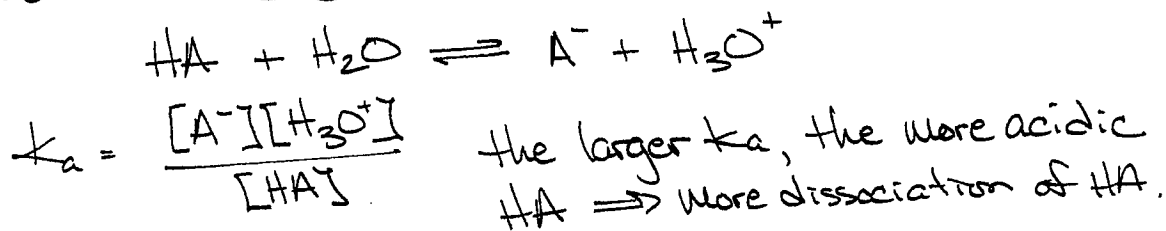
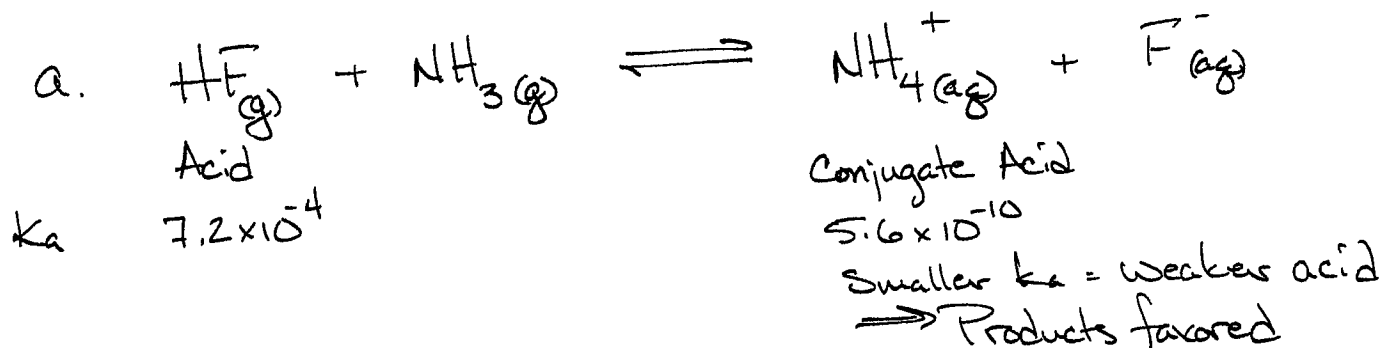


Chapter 9 Homework Key

9.50 Draw the conjugate base of each acid.

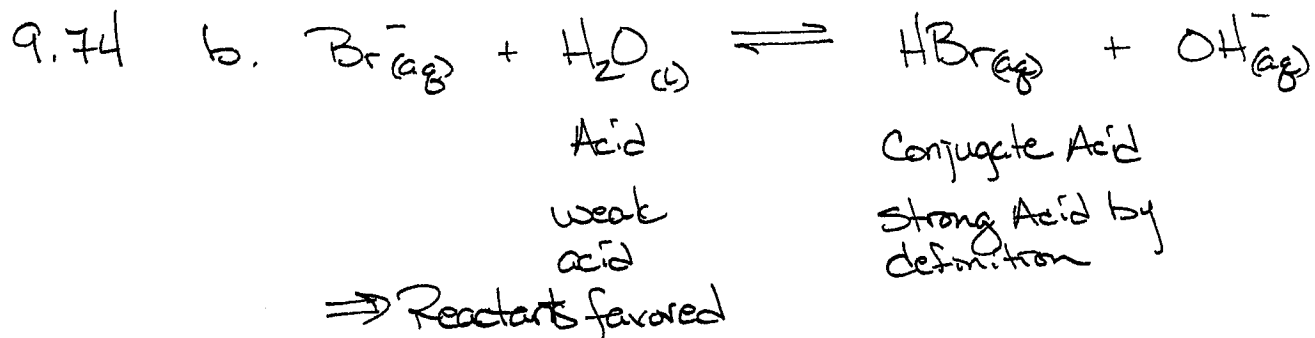


9.74 Label the acid in the reactants and the conjugate acid in the products in each reaction. Use the data in Tables 9.1 and 9.2 to determine whether the reactants or products are favored at equilibrium. Explain your reasoning.

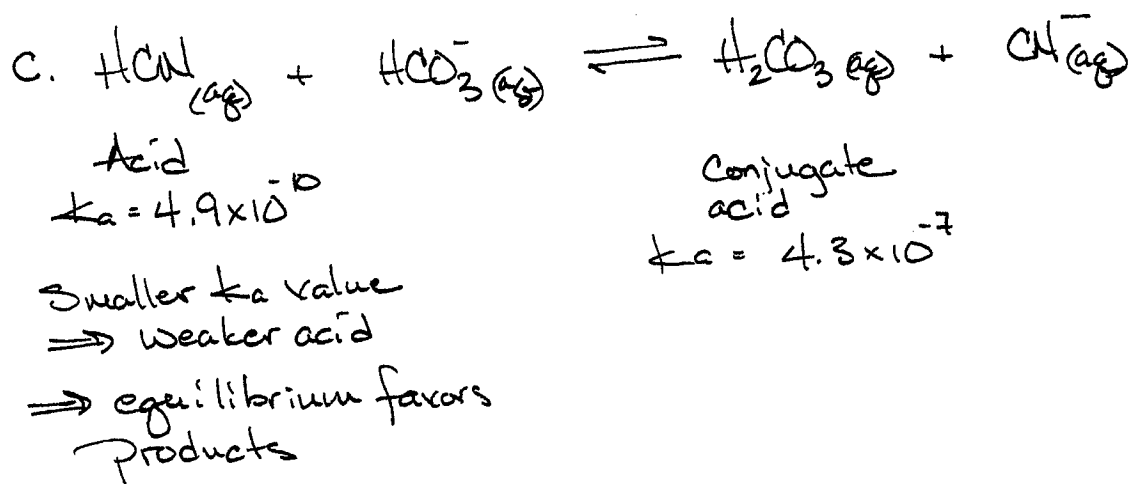


The weaker of the two acids will be the dominate side \Rightarrow smaller K_a value

$\left[\frac{\text{Large } K_a}{\Delta} \right] \quad \left[\text{small } K_a \right]$



Here we don't need the K_a values to make a conclusion. HBr is a strong acid like HCl. H_2O is a weak acid. the equilibrium will favor the side with water.



9.100 What is the molarity of an H_2SO_4 solution if 18.5 mL of 0.18 M NaOH are needed to neutralize 25.0 mL of the sample?



Volume \Rightarrow 25.0 mL 18.5 mL } moles
 ? M 0.18 M

$$\frac{\text{moles}}{\text{Volume}} = \frac{18.5 \text{ mL NaOH}}{25.0 \text{ mL H}_2\text{SO}_4} \times \frac{1 \text{ L NaOH}}{1000 \text{ mL NaOH}} \times \frac{0.18 \text{ moles NaOH}}{1 \text{ L NaOH}} \times \frac{1 \text{ mole H}_2\text{SO}_4}{2 \text{ moles NaOH}} \times \frac{1000 \text{ mL}}{1 \text{ L}}$$

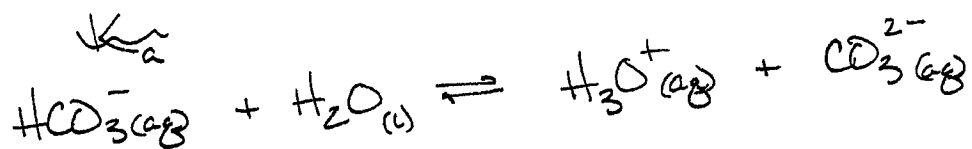
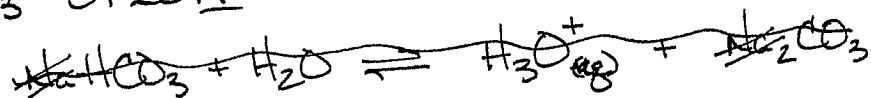
$$= \frac{0.0666 \text{ moles H}_2\text{SO}_4}{1 \text{ L H}_2\text{SO}_4} = \frac{0.067 \text{ moles H}_2\text{SO}_4}{1 \text{ L H}_2\text{SO}_4}$$

$$= \boxed{0.067 \text{ M H}_2\text{SO}_4}$$

9.112 Calculate the pH of a bicarbonate/carbonate buffer in which the concentration of sodium bicarbonate (NaHCO_3) is always 0.20 M , but the concentration of sodium carbonate (Na_2CO_3) corresponds to each of the following values.

a. NaHCO_3 0.20 M

Na_2CO_3 0.20 M



Na^+ just a spectator

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CO}_3^{2-}]}{[\text{HCO}_3^-]} = 5.6 \times 10^{-11} \quad \left| \text{solve for } [\text{H}_3\text{O}^+] \right.$$

$$[\text{H}_3\text{O}^+] = \frac{K_a [\text{HCO}_3^-]}{[\text{CO}_3^{2-}]} \quad \left| \text{take } -\log \text{ both sides} \right.$$

$$-\log [\text{H}_3\text{O}^+] = -\log K_a - \log \left(\frac{[\text{HCO}_3^-]}{[\text{HCO}_3^-]} \right)$$

$$\text{pH} = -\log K_a - \log \left(\frac{[\text{HCO}_3^-]}{[\text{CO}_3^{2-}]} \right)$$

$$\text{pH} = -\log(5.6 \times 10^{-11}) - \log \left(\frac{0.20 \text{ M}}{0.20 \text{ M}} \right)$$

$$\text{pH} = 10.25 \quad \left| \begin{array}{c} \uparrow \\ 1812 \end{array} \right.$$

$$\boxed{\text{pH} = 10.25}$$

9.112 b. NaHCO_3 0.20M

Na_2CO_3 0.40M

Now we just use the equation from before with the new values.

$$\text{pH} = -\log k_a - \log \left(\frac{[\text{HCO}_3^-]}{[\text{CO}_3^{2-}]} \right)$$

$$\text{pH} = -\log(5.6 \times 10^{-11}) - \log \left(\frac{0.20 \text{ M}}{0.40 \text{ M}} \right)$$

$$\text{pH} = 10.55 \overline{2842}$$

$$\boxed{\text{pH} = 10.55}$$

c. NaHCO_3 0.20M

Na_2CO_3 0.10M

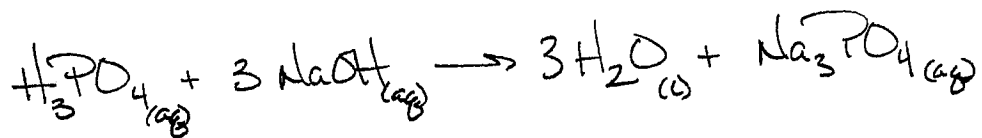
$$\text{pH} = -\log k_a - \log \left(\frac{[\text{HCO}_3^-]}{[\text{CO}_3^{2-}]} \right)$$

$$\text{pH} = -\log(5.6 \times 10^{-11}) - \log \left(\frac{0.20 \text{ M}}{0.10 \text{ M}} \right)$$

$$\text{pH} = 9.95 \overline{782}$$

$$\boxed{\text{pH} = 9.95}$$

9.124 Most buffers are prepared using a weak acid and a salt of its conjugate base. Explain how the following combination can also form a buffer solution: $0.20\text{ M H}_3\text{PO}_4$ and 0.10 M NaOH .



Lets assume 1L solution and look at what happens.

$$1\text{ L solution} \times \frac{0.10\text{ mols NaOH}}{1\text{ L solution}} \times \frac{1\text{ mole H}_3\text{PO}_4\text{ neutralized}}{3\text{ moles NaOH}} = \frac{0.03}{\cancel{0.3}}\text{ moles H}_3\text{PO}_4\text{ neutralized}$$

with only ~~0.3~~ 0.03 moles H_3PO_4 neutralized this means that there is 0.03 moles PO_4^{3-} and still 0.17 moles H_3PO_4 . Thus by adding a small amount of NaOH we caused a small amount of the phosphate salt to be made. This is the same as adding $0.17\text{ M H}_3\text{PO}_4$ and $0.03\text{ M Na}_3\text{PO}_4$, there by making a buffer solution.