Chapter 9 Homework Answer Key

Chapter 9 - 39, 41, 43, 45, 46, 51, 53, 57, 63, 65, 69, 73, 74, 75, 77, 81, 83, 85, 89, 91, 93, 95, 99, 101, 102, 109, 113, 114, 119, 124

46) Draw the conjugate base of each acid.

Conjugates differ by 1 proton (hydrogen ion). The conjugate base will have 1 less H^+ changing the hydrogen count and the charge.

a) $\underbrace{H_3O^+}_{\text{total charge } +1} \longrightarrow \underbrace{H_2O + H^+}_{\text{total charge } +1}$ ← Just to track

b) $\underbrace{H_2Se}_{\text{total charge } \varnothing} \longrightarrow \underbrace{HSe^- + H^+}_{\text{net charge } \varnothing}$

c) $\underbrace{HSO_4^-}_{\text{Total charge } -1} \longrightarrow \underbrace{SO_4^{2-} + H^+}_{\text{net charge } -1}$

74) Label the acid in the reactants and the Conjugate acid in the products in each reaction. Use data in Tables 9.2 and 9.3 to determine whether the reactants or products are favored at equilibrium. Explain your reasoning.

a) $HF_{(g)} + NH_{3(g)} \rightleftharpoons NH^{+}_{4(aq)} + F^{-}_{(aq)}$

Table 9.2	Weak Acid	stronger base	weak acid	Weak base
Table 9.3 K_a	7.2×10^{-4}		5.6×10^{-10}	

HF
Acid

NH_4^+
Conjugate acid

Equilibrium will favor the weaker side of the equation, the side with the weaker conjugates. You can also compare the K_a values for the two acids. The smaller K_a value will be favored $\Rightarrow$ Smaller K_a = Weaker acid.
It's a contest to get rid of the H^+. Stronger acid gets rid of the H^+, weaker acid is stuck with it.

$HF + NH_3 \rightleftharpoons NH_4^+ + F^-$

Products favored

74)

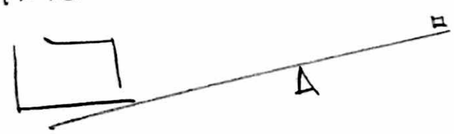
b) $Br^-_{(aq)} + H_2O_{(l)} \rightleftharpoons HBr_{(aq)} + OH^-_{(aq)}$

Table 9.2	Weakest acid	Almost Strongest acid
Ka	2.0×10^{-16}	$\sim 1.0 \times 10^{4}$

Equilibrium lies to the reactant side, the weaker acid side.

c) $HCN_{(aq)} + HCO_3^-{}_{(aq)} \rightleftharpoons H_2CO_3{}_{(aq)} + CN^-_{(aq)}$

Table 9.2	Weakest Acid	Stronger Acid
Ka	4.9×10^{-10}	4.3×10^{-7}



Equilibrium lies to the reactant side, the weaker acid side.

102) 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?

$$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$$

	H_2SO_4	$NaOH$
Vol	25.0 mL	18.5 mL
Conc.	?	0.18 M

2 steps → ① find moles H_2SO_4 from NaOH
② Divide moles H_2SO_4 by volume H_2SO_4 in L

① mL NaOH → L NaOH → moles NaOH → moles H_2SO_4

$$18.5\text{ mL NaOH} \times \frac{1\text{ L NaOH}}{1000\text{ mL NaOH}} \times \frac{\overset{2SF}{0.18}\text{ moles NaOH}}{1\text{ L NaOH}} \times \frac{1\text{ mole }H_2SO_4}{2\text{ moles NaOH}}$$

$= 0.001665$ moles H_2SO_4 2SF

② $$\frac{0.001665\text{ moles }H_2SO_4}{25.0\text{ mL }H_2SO_4} \times \frac{1000\text{ mL }H_2SO_4}{1\text{ L }H_2SO_4} = 0.0666\text{ mol/L }H_2SO_4$$

$$= \boxed{0.067\text{ mol/L }H_2SO_4}$$

114) 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) 0.20 M b) 0.40 M c) 0.10 M

General Acid/Base

$$HA + H_2O \rightleftharpoons A^- + H_3O^+ \quad \Big| \quad K_a = \frac{[A^-][H_3O^+]}{[HA]}$$

For Carbonate Buffer

$$HCO_3^- + H_2O \rightleftharpoons CO_3^{2-} + H_3O^+ \quad \Big| \quad K_a = \frac{[CO_3^{2-}][H_3O^+]}{[HCO_3^-]}$$

Now using the K_a expression:

$$K_a = \frac{[CO_3^{2-}][H_3O^+]}{[HCO_3^-]}$$ solve in terms of $[H_3O^+]$

$$[H_3O^+] = K_a \frac{[HCO_3^-]}{[CO_3^{2-}]}$$ Now take $-\log$ both Sides

$$-\log[H_3O^+] = -\log\left(K_a \frac{[HCO_3^-]}{[CO_3^{2-}]}\right)$$ Distribute $-\log$ and clean up

$$\underbrace{-\log[H_3O^+]}_{pH} = \underbrace{-\log K_a}_{pK_a} + -\log\frac{[HCO_3^-]}{[CO_3^{2-}]}$$

$$\Longrightarrow pH = pK_a - \log\frac{[HCO_3^-]}{[CO_3^{2-}]}$$

114 cont.) $pH = pK_a - \log \frac{[HCO_3^-]}{[CO_3^{2-}]}$

pK_a for Carbonate = 10.25

Now given $[HCO_3^-] = 0.20\,\underline{M}$ always our equation becomes:

$$pH = 10.25 - \log\left(\frac{0.20\,M}{[CO_3^{2-}]}\right)$$

a) $CO_3^{2-} = 0.20\,\underline{M}$

$$pH = 10.25 - \log\left(\frac{0.20\,\underline{M}}{0.20\,\underline{M}}\right) = \boxed{10.25}$$

b) $CO_3^{2-} = 0.40\,\underline{M}$

$$pH = 10.25 - \log\left(\frac{0.20\,\underline{M}}{0.40\,M}\right) = \boxed{10.55}$$

c) $CO_3^{2-} = 0.10\,\underline{M}$

$$pH = 10.25 - \log\left(\frac{0.20\,\underline{M}}{0.10\,\underline{M}}\right) = \boxed{9.95}$$

124) Most buffer solutions 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 M H_3PO_4 and 0.10 M NaOH.

This is a harder conceptual problem

$$H_3PO_4 + NaOH \rightarrow Na_3PO_4 + 3H_2O$$

for neutralization. However there isn't enough NaOH to deprotonate (neutralize) all of the H_3PO_4 $\Rightarrow$ Reaction is incomplete

	H_3PO_4 +	NaOH	$\rightarrow$ H_2O +	H_3PO_4 +	NaH_2PO_4
Assume 1 L Sol	0.20 mols	0.10 mols *Limiting Reagent	0.10 mols	0.10 moles	0.10 moles

Because NaOH is Limiting it forms an equivalent of NaH_2PO_4 $\Rightarrow$ The conjugate base of H_3PO_4.

Thus by adding some NaOH we make a buffer by making a solution of H_3PO_4 (a weak acid) and NaH_2PO_4 (the conjugate base).