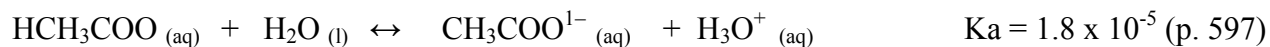


Unit 7, Lesson 06: Answers to Homework

Page 400, Q 19 - 24

19. Calculate the pH and percent dissociation of 0.83 M solution of vinegar:



	HCH ₃ COO _(aq)	⇌	CH ₃ COO ¹⁻ _(aq)	+	H ₃ O ⁺ _(aq)
I	0.83 M		0		0
C	- x		+ x		+ x
E	0.83 - x		x		x

$$K_a = \frac{[\text{CH}_3\text{COO}^{1-}][\text{H}_3\text{O}^+]}{[\text{HCH}_3\text{COO}]}$$

$$1.8 \times 10^{-5} = \frac{x^2}{0.83}$$

$$x = 3.865 \times 10^{-3} \text{ mol/L}$$

$$\text{so } [\text{H}_3\text{O}^+] \text{ at eq'm} = 3.865 \times 10^{-3} \text{ mol/L}$$

$$\text{pH} = -\log [3.865 \times 10^{-3}]$$

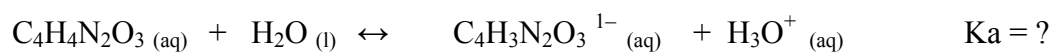
$$= 2.41 \text{ (2 decimals, because 2 sd)}$$

$$\begin{aligned} \text{\% dissociation} &= \frac{[\text{H}_3\text{O}^+]}{[\text{HCH}_3\text{COO}]} \times 100\% \\ &= \frac{3.9 \times 10^{-3} \text{ mol/L}}{0.83 \text{ mol/L}} \times 100\% \\ &= 0.47\% \text{ dissociated} \end{aligned}$$

Can we ignore the -x?

$$\frac{0.83}{1.8 \times 10^{-5}} \text{ is much greater than 500, so ignore -x}$$

20. Calculate the K_a of a 0.10 M solution of barbituric acid if it has a pH of 2.50.



	$\text{C}_4\text{H}_4\text{N}_2\text{O}_3 (\text{aq})$	\rightleftharpoons	$\text{C}_4\text{H}_3\text{N}_2\text{O}_3^{1-} (\text{aq})$	+	$\text{H}_3\text{O}^+ (\text{aq})$
I	0.10 M		0		0
C	- x		+ x		+ x
E	0.10 - x		x		x

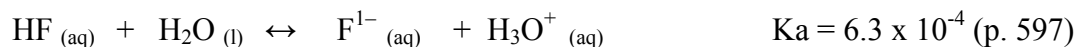
$$\begin{aligned} K_a &= \frac{[\text{C}_4\text{H}_3\text{N}_2\text{O}_3^{1-}] [\text{H}_3\text{O}^+]}{[\text{C}_4\text{H}_4\text{N}_2\text{O}_3]} \\ &= \frac{[0.003162]^2}{0.10} \\ &= 1.00 \times 10^{-4} \text{ mol/L} \end{aligned}$$

We know pH = 2.50, so we can find $[\text{H}_3\text{O}^+]$

$$\begin{aligned} [\text{H}_3\text{O}^+] &= 10^x [-2.50] \\ &= 0.003162 \text{ M} \\ &= x \end{aligned}$$

$$\begin{aligned} \% \text{ dissociation} &= \frac{[\text{H}_3\text{O}^+]}{[\text{C}_4\text{H}_4\text{N}_2\text{O}_3]} \times 100\% \\ &= \frac{0.003162 \text{ mol/L}}{0.10 \text{ mol/L}} \times 100\% \\ &= 3.2\% \text{ dissociated} \end{aligned}$$

21. Calculate the pH of 0.0100 M solution of HF:



	$\text{HF} (\text{aq})$	\rightleftharpoons	$\text{F}^{1-} (\text{aq})$	+	$\text{H}_3\text{O}^+ (\text{aq})$
I	0.0100 M		0		0
C	- x		+ x		+ x
E	0.0100 - x		x		x

$$\begin{aligned} K_a &= \frac{[\text{F}^{1-}] [\text{H}_3\text{O}^+]}{[\text{HF}]} \\ 6.3 \times 10^{-4} &= \frac{x^2}{(0.0100 - x)} \end{aligned}$$

Can we ignore the -x?

$\frac{0.0100}{6.3 \times 10^{-4}}$ is less than than 500, so must use the quadratic formula

$$-x^2 - 6.3 \times 10^{-4}x + 6.3 \times 10^{-6} = 0 \quad \text{or} \quad x^2 + 6.3 \times 10^{-4}x - 6.3 \times 10^{-6} = 0$$

$$x = 0.002215 \text{ or } x = -0.002845 \text{ (inadmissible)} \quad \text{so, } [\text{H}_3\text{O}^+] = 0.002215 \text{ mol/L}$$

$$\text{pH} = -\log [0.002215]$$

$$= 2.65 \text{ (2 decimals, because 2 sd)}$$

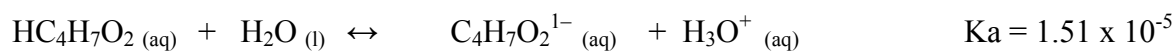
22. What is K_a for HClO if a 0.40 M solution is 0.027% dissociated?

if HClO is 0.027% dissociated, then $[\text{H}_3\text{O}^+] = 0.00027 \times 0.40 \text{ M}$

$$= 0.000108 \text{ mol/L} = [\text{ClO}^-]$$

$$\begin{aligned} K_a &= \frac{[\text{ClO}^{1-}] [\text{H}_3\text{O}^+]}{[\text{HClO}]} \\ &= \frac{[0.000108] [0.000108]}{[0.40]} \\ &= 2.9 \times 10^{-8} \end{aligned}$$

23. Calculate the pH of $1.0 \times 10^{-2} \text{ M}$ solution of butanoic acid:



	$\text{HCH}_3\text{COO} \text{ (aq)}$	\rightleftharpoons	$\text{CH}_3\text{COO}^{1-} \text{ (aq)}$	+	$\text{H}_3\text{O}^+ \text{ (aq)}$
I	0.010 M		0		0
C	- x		+ x		+ x
E	0.010 - x		x		x

$$K_a = \frac{[\text{C}_4\text{H}_7\text{O}_2^{1-}] [\text{H}_3\text{O}^+]}{[\text{HC}_4\text{H}_7\text{O}_2]}$$

$$1.51 \times 10^{-5} = \frac{x^2}{0.010}$$

$$x = 3.89 \times 10^{-4} \text{ mol/L}$$

so $[\text{H}_3\text{O}^+] \text{ at eq'm} = 3.89 \times 10^{-4} \text{ mol/L}$

$$\text{pH} = -\log [3.89 \times 10^{-4}]$$

$$= 3.41 \text{ (2 decimals, because 2 sd)}$$

Can we ignore the -x?

$$\frac{0.010}{1.51 \times 10^{-5}} \text{ is much greater than 500, so ignore -x}$$

24. Calculate the concentration of a solution of caproic acid if it has a pH of 2.94.



We know $\text{pH} = 2.94$, so we can find $[\text{H}_3\text{O}^+]$

$$[\text{H}_3\text{O}^+] = 10^x [-2.94]$$

$$= 0.001148 \text{ M} = [\text{C}_5\text{H}_{11}\text{COO}^{1-}]$$

If the original concentration of $\text{C}_5\text{H}_{11}\text{COOH}$ before dissociation was “x”, then after dissociation, the $[\text{C}_5\text{H}_{11}\text{COOH}] = x - 0.001148 \text{ mol/L}$

$$K_a = \frac{[\text{C}_5\text{H}_{11}\text{COO}^{1-}][\text{H}_3\text{O}^+]}{[\text{C}_5\text{H}_{11}\text{COOH}]}$$

$$1.3 \times 10^{-5} = \frac{[0.001148]^2}{[x - 0.001148]}$$

$$x = 0.102 \text{ mol/L concentration of solution}$$

There were 100 mL of solution, so calculate moles:

$$n = C \times V$$

$$= 0.102 \text{ mol/L} \times 100 \text{ mL}$$

$$= 0.0102 \text{ mol of caproic acid}$$

Now find the mass of caproic acid needed to make this solution:

$$m = n \times \text{MM}$$

$$= 0.0102 \text{ mol} \times 116.18 \text{ g/mol}$$

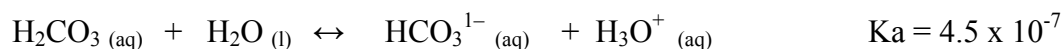
$$= 1.185 \text{ g of caproic acid in 100 mL of solution}$$

$$= 1.2 \text{ g of caproic acid (2 sig digs, from } K_a)$$

(I think the text book messed up, it should have been a volume of 100.0 mL or they considered 100 mL to have 3 sd).

Questions 25 – 28 on pages 402 – 403

25. Calculate the pH of 0.00050 mol/L solution of carbonic acid:



	$\text{H}_2\text{CO}_3(\text{aq})$	\rightleftharpoons	$\text{HCO}_3^{1-}(\text{aq})$	+	$\text{H}_3\text{O}^+(\text{aq})$
I	0.00050 M		0		0
C	- x		+ x		+ x
E	0.00050 - x		x		x

$$K_a = \frac{[\text{HCO}_3^{1-}][\text{H}_3\text{O}^+]}{[\text{H}_2\text{CO}_3]}$$

$$4.5 \times 10^{-7} = \frac{x^2}{0.00050}$$

$$x = 1.5 \times 10^{-5} \text{ mol/L}$$

so $[\text{H}_3\text{O}^+]$ at eq'm = 1.5×10^{-5} mol/L

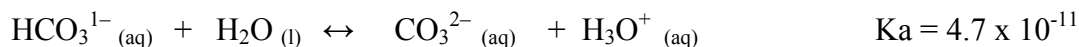
$$\text{pH} = -\log [1.5 \times 10^{-5}]$$

$$= 4.82 \text{ (2 decimals, because 2 sd)}$$

Can we ignore the -x?

$$\frac{0.00050}{4.5 \times 10^{-7}} \text{ is greater than 500, so ignore -x}$$

To calculate the concentration of CO_3^{2-} in solution, we need to use the second dissociation reaction, and the concentration of H_3O^+ and HCO_3^{1-} determined in the first calculation:



	$\text{HCO}_3^{1-}(\text{aq})$	\rightleftharpoons	$\text{CO}_3^{2-}(\text{aq})$	+	$\text{H}_3\text{O}^+(\text{aq})$
I	1.5×10^{-5} mol/L		0		1.5×10^{-5} mol/L
C	- x		+ x		+ x
E	$1.5 \times 10^{-5} - x$		x		$1.5 \times 10^{-5} + x$

$$K_a = \frac{[\text{CO}_3^{2-}][\text{H}_3\text{O}^+]}{[\text{HCO}_3^{1-}]}$$

$$4.7 \times 10^{-11} = \frac{[x][1.5 \times 10^{-5}]}{1.5 \times 10^{-5}}$$

$$\text{and } x = 4.7 \times 10^{-11}$$

So, the concentration of CO_3^{2-} in solution is 4.7×10^{-11} mol/L

The value of x for the second dissociation is going to be so small, we can ignore it for both of these terms.

The remaining problems are similar to those shown above. The number of ions contributed by the second dissociation is negligible, so calculate pH based on the first dissociation only.