AP CHEMISTRY

TOPIC 7: ACIDS & BASES, PART B — EXAMPLES, PART II — Day 76:

- pH and pOH scale
- pH of strong acids
- pH of weak acids

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1. a) Calculate the [H^{+1}] concentration for a solution with a pH of 4.22

$$pH = 4.22\ ,\ \left[H^{+1}\right] = \text{antilog}\left(-4.22\right) = 6.03\times10^{-5}$$

b) Calculate the [OH^{-1}] concentration for a solution with a pH of 3.75

$$pH = 3.75\ ,\ \left[H^{+1}\right] = \text{antilog}\left(-3.75\right) = 1.78\times10^{-4}\ ,\ \left[OH^{-1}\right] = \frac{1.00\times10^{-14}}{\left[H^{+1}\right]} = \frac{1.00\times10^{-14}}{1.78\times10^{-4}} = 5.62\times10^{-11}$$

c) Calculate the pOH for a solution with a [H^{+1}] concentration of 7.83 x 10^{-12}.

$$\left[OH^{-1}\right] = \frac{1.00\times10^{-14}}{\left[H^{+1}\right]} = \frac{1.00\times10^{-14}}{7.83\times10^{-12}} = 1.28\times10^{-3}\ ,\ pOH = -\log\left(1.28\times10^{-3}\right) = 2.89$$

OR $$pOH = 14-\left(-\log\left(7.83\times10^{-12}\right)\right) = 2.89$$

d) Calculate the pH for a strong acid with an initial concentration of 5.55 x 10^{-3} *M*.

$$HA \rightarrow H^{+1} + A^{-1}$$

$$\left[H^{+1}\right] = 5.55\times10^{-3}\ M\ ,\ pH = -\log\left(5.55\times10^{-3}\right) = 2.26$$

2. A solution is prepared by adding 3.00 L of 0.253 *M* $HClO_4$ to 1.35 L of 1.13 *M* HBr. Calculate the concentrations of all the species in this solution. (hint: do a dissociation for each strong acid, AND recall what the solvent for these acids, H_2O.)

Answers: ***Strong Acids = 100 % dissociation, Total Volume = 4.35 L***

$HClO_4$: $$\frac{3.00\ L}{} \times \frac{0.253\ mol}{L} = 0.759\ mol$$

HBr : $$\frac{1.35\ L}{} \times \frac{1.13\ mol}{L} = 1.5255\ mol$$

$HClO_3 \rightarrow H^{+1} + ClO_3^{-1}$
$HBr \rightarrow H^{+1} + Br^{-1}$

[$HClO_4$] = [ClO_4^{-1}] = $$\frac{0.750\ mol}{4.35\ L} = 0.172\ M$$ **= [H^{+1}]**

[HBr] = [Br^{-1}] = $$\frac{1.5255\ mol}{4.35\ L} = 0.351\ M$$ **= [H^{+1}]**

[OH^{-1}] = $$\frac{1.00\times10^{-14}}{\left[H^{+}\right]} = 1.91\times10^{-14}\ M$$

[H^{+1}] 0.172 *M* + 0.351 *M*
[H^{+1}] = 0.523 *M*

2. Calculate the pH of a 5.40 *M* benzoic acid, $HC_7H_5O_2$, solution - a weak acid with a $K_a = 6.52 \times 10^{-5}$ at 25 ^{0}C.

Step 1: write the dissociation reaction for $HC_7H_5O_2$:

$$HC_7H_5O_{2\,(aq)} \rightleftharpoons H^{+1}{}_{(aq)} + C_7H_5O_2^{-1}{}_{(aq)}$$

Step 2: write the equilibrium expression:

$$K_a = 6.52 \times 10^{-5} = \frac{[\ H^+\ ][\ C_7H_5O_2^-\ ]}{[\ HC_7H_5O_2\ ]}$$

Step 3: make an ICE chart:

	[$HC_7H_5O_2$]	↔	[H^{+1}]	+	[$C_7H_5O_2^{-1}$]
I	5.40 *M*		0		0
C	$-x$		$+x$		$+x$
E	5.40 *M* - x		x		x

Step 4: set up the algebra to solve for "x"

$$K_a = 6.52 \times 10^{-5} = \frac{(\ x\ )(\ x\ )}{(\ 5.40\ -\ x\ )}$$

again, since x will be small, the equilibrium concentration for HF will go from (1.0 – x) to (1.0)

$$K_a = 6.52 \times 10^{-5} = \frac{(\ x\ )(\ x\ )}{(\ 5.40\ )}\ ;\ 3.52 \times 10^{-4} = x^2\ ;\ \sqrt{3.52 \times 10^{-4}} = x = 0.0188$$

Step 5: write the equilibrium concentrations for each:

[$HC_7H_5O_2$] = 5.40 *M*, [H^{+1}] = 0.0188 *M*, [$C_7H_5O_2^{-1}$] = 0.0188 *M*

Step 6: calculate the pH from [H^+]

pH = - log (0.0188) = 1.73

4. Calculate the pOH of a 0.50 *M* hypochlorous acid, HClO, solution - a weak acid with a $K_a = 3.00 \times 10^{-8}$ at 25 ^{0}C.

Step 1: write the dissociation reaction for HClO:

$$HClO_{(aq)} \rightleftharpoons H^{+1}{}_{(aq)} + ClO^{-1}{}_{(aq)}$$

Step 2: write the equilibrium expression:

$$K_a = 3.00 \times 10^{-8} = \frac{[\ H^+\ ][\ ClO^-\ ]}{[\ HClO\ ]}$$

Step 3: make an ICE chart:

	[HClO]	↔	[H^{+1}]	+	[ClO^{-1}]
I	0.50 *M*		0		0
C	$-x$		$+x$		$+x$
E	0.50 *M* - x		x		x

Step 4: set up the algebra to solve for "x"

$$K_a = 3.00\times10^{-8} = \frac{(\ x\ )\ (\ x\ )}{(\ 0.50\ -\ x\ )}$$

again, since x will be small, the equilibrium concentration for HF will go from (1.0 – x) to (1.0)

$$K_a = 3.00\times10^{-8} = \frac{(\ x\ )\ (\ x\ )}{(\ 0.50\ )} \quad ; \quad 1.50\times10^{-8} = x^2 \quad ; \quad \sqrt{1.50\times10^{-8}} = x = 1.22\times10^{-4}$$

Step 5: write the equilibrium concentrations for each:

$$[\ HClO^{-1}\ ] = 0.50\ M,\ [\ H^+\ ] = 1.22\times10^{-4}\ M,\ [\ ClO^{-1}\ ] = 0.0188\ M$$

Step 6: calculate the pOH from [H^+]

$$pH = -\log(\ 1.22\times10^{-4}\ ) = 3.91$$

$$pOH = 14 - 3.91 = 10.09$$

OR

$$pOH = 14 - (-\log(\ 1.22\times10^{-4}\ )) = 10.09$$