

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
-

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

$$pH = 4.22, [H^+] = \text{antilog}(-4.22) = 6.03 \times 10^{-5}$$

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

$$pH = 3.75, [H^+] = \text{antilog}(-3.75) = 1.78 \times 10^{-4}, [OH^-] = \frac{1.00 \times 10^{-14}}{[H^+]} = \frac{1.00 \times 10^{-14}}{1.78 \times 10^{-4}} = 5.62 \times 10^{-11}$$

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

$$[OH^-] = \frac{1.00 \times 10^{-14}}{[H^+]} = \frac{1.00 \times 10^{-14}}{7.83 \times 10^{-12}} = 1.28 \times 10^{-3}, pOH = -\log(1.28 \times 10^{-3}) = 2.89$$

$$\text{OR } pOH = 14 - (-\log(7.83 \times 10^{-12})) = 2.89$$

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



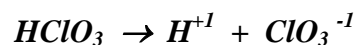
$$[H^+] = 5.55 \times 10^{-3} M, pH = -\log(5.55 \times 10^{-3}) = 2.26$$

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

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

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



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

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

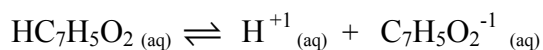
$$[OH^-] = \frac{1.00 \times 10^{-14}}{[H^+]} = 1.91 \times 10^{-14} M$$

$$[H^+] = 0.172 M + 0.351 M$$

$$[H^+] = 0.523 M$$

2. Calculate the pH of a 5.40 M benzoic acid, $\text{HC}_7\text{H}_5\text{O}_2$, solution - a weak acid with a $K_a = 6.52 \times 10^{-5}$ at 25 °C.

Step 1: write the dissociation reaction for $\text{HC}_7\text{H}_5\text{O}_2$:



Step 2: write the equilibrium expression:

$$K_a = 6.52 \times 10^{-5} = \frac{[\text{H}^+][\text{C}_7\text{H}_5\text{O}_2^-]}{[\text{HC}_7\text{H}_5\text{O}_2]}$$

Step 3: make an ICE chart:

	$[\text{HC}_7\text{H}_5\text{O}_2]$	\leftrightarrow	$[\text{H}^+]$	+	$[\text{C}_7\text{H}_5\text{O}_2^-]$
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:

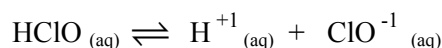
$$[\text{HC}_7\text{H}_5\text{O}_2] = 5.40 \text{ M}, [\text{H}^+] = 0.0188 \text{ M}, [\text{C}_7\text{H}_5\text{O}_2^-] = 0.0188 \text{ M}$$

Step 6: calculate the pH from $[\text{H}^+]$

$$\text{pH} = -\log(0.0188) = 1.73$$

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 °C.

Step 1: write the dissociation reaction for HClO :



Step 2: write the equilibrium expression:

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

Step 3: make an ICE chart:

	$[\text{HClO}]$	\leftrightarrow	$[\text{H}^+]$	+	$[\text{ClO}^-]$
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"

$$K_a = 3.00 \times 10^{-8} = \frac{(x)(x)}{(0.50 - x)}$$

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

$$K_a = 3.00 \times 10^{-8} = \frac{(x)(x)}{(0.50)} ; 1.50 \times 10^{-8} = x^2 ; \sqrt{1.50 \times 10^{-8}} = x = 1.22 \times 10^{-4}$$

Step 5: write the equilibrium concentrations for each:

$$[\text{HClO}^{-1}] = 0.50 \text{ M}, [\text{H}^{+}] = 1.22 \times 10^{-4} \text{ M}, [\text{ClO}^{-1}] = 0.0188 \text{ M}$$

Step 6: calculate the pOH from $[\text{H}^{+}]$

$$\text{pH} = -\log(1.22 \times 10^{-4}) = 3.91$$

$$\text{pOH} = 14 - 3.91 = 10.09$$

OR

$$\text{pOH} = 14 - (-\log(1.22 \times 10^{-4})) = 10.09$$