AP CHEMISTRY

TOPIC 7: ACIDS & BASES, PART B

Day 76:

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

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1. Calculate the pH of each of the following strong acids in water.

   a) 0.10 *M* HCl **: [H^+] = 0.10 *M*, *pH = - log (0.10) = 1.00***

   b) 5.0×10^{-4} *M* HCl **: [H^+] = 5.0×10^{-4} *M*, *pH = - log (5.0×10^{-4}) = 3.30***

   c) 1.0×10^{-11} *M* HCl **: [H^+] = 1.0×10^{-11} *M*, *pH = - log (1.0×10^{-11}) = 11.00***

2. A solution is prepared by adding 50.0 mL of 0.0500 *M* HCl to 150. mL of 0.10 *M* HNO_3. 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 = 200 mL = 0.2 L***

HCl : $$\frac{50\ mL}{} \times \frac{1\ L}{1000\ mL} \times \frac{0.05\ mol}{L} = 0.0025\ mol$$

HNO_3 : $$\frac{150\ mL}{} \times \frac{1\ L}{1000\ mL} \times \frac{0.10\ mol}{L} = 0.015\ mol$$

$$HCl \rightarrow H^{+1} + Cl^{-1}$$

$$HNO_3 \rightarrow H^{+1} + NO_3^{-1}$$

$$[\ Cl^{-1}\ ] = \frac{0.0025\ mol}{0.2\ L} = 0.0125\ M$$

$$[\ NO_3^{-1}\ ] = \frac{0.015\ mol}{0.2\ L} = 0.075\ M$$

$$[\ H^{+1}\ ] = \frac{(0.015\ mol\ +\ 0.0025\ mol)}{0.2\ L} = 0.0875\ M$$

3. Calculate the concentration of an aqueous HCl solution that has a pH = 2.50.

$$[\ H^{+1}\ ] = \text{Antilog}(-2.50) = 3.16 \times 10^{-3}\ M$$

HCl is a strong acid, 100 % dissociation

$$HCl \rightarrow H^{+1} + Cl^{-1}$$

Therefore, **[HCl] = [Cl^{-1}] = [H^{+1}]**

4. A 0.20 M solution of acetic acid has a $K_a = 1.8 \times 10^{-5}$. Calculate the concentration of all the species present in the solution. Also, calculate the pH.

Answers: ***WEAK ACID …***

	$[HC_2H_3O_2]$	$\leftrightarrow$	$[H^{+1}]$	+	$[C_2H_3O_2^{-1}]$
I	0.20 M		0		0
C	$- x$		$+ x$		$+ x$
E	0.20 $M - x$		x		x

$$K_a = 1.8 \times 10^{-5} = \frac{\left[H^{+1} \right]\left[C_2H_3O_2^{-1} \right]}{\left[HC_2H_3O_2 \right]^2} = \frac{x^2}{0.20 - x} = \frac{x^2}{0.20}$$

$$1.8 \times 10^{-5} (0.2) = x^2 \; ; \; x = \sqrt{3.6\times10^{-6}} = 1.9\times10^{-3} \; M$$

$$\boldsymbol{pH = - \log (1.9 \times 10^{-3}) = 2.72}$$

5. Calculate the $[H^+]$ and $[OH^-]$ for each solution.

a) pH = 7.41 (normal pH for blood),

$[H^{+1}]$ = *Antilog (- 7.41)* = 3.89 x 10^{-8} *M* , $\left[OH^{-1} \right] = \dfrac{1.00\times10^{-14}}{3.89\times10^{-8}} = 2.57\times10^{-7} \; M$

b) pH = 13.00

$[H^{+1}]$ = *Antilog (- 13.4)* = 3.98 x 10^{-14} *M* , $\left[OH^{-1} \right] = \dfrac{1.00\times10^{-14}}{3.98\times10^{-14}} = 0.251 \; M$

c) pH = 3.20

$[H^{+1}]$ = *Antilog (- 3.2)* = 6.31 x 10^{-4} *M* , $\left[OH^{-1} \right] = \dfrac{1.00\times10^{-14}}{6.31\times10^{-4}} = 1.58\times10^{-11} \; M$

d) pOH = 3.20

$[OH^{-1}]$ = *Antilog (- 3.2)* = 6.31 x 10^{-4} *M* , $\left[H^{+1} \right] = \dfrac{1.00\times10^{-14}}{6.31\times10^{-4}} = 1.58\times10^{-11} \; M$

e) pOH = 9.30

$[OH^{-1}]$ = *Antilog (- 9.3)* = 5.01 x 10^{-10} *M* , $\left[H^{+1} \right] = \dfrac{1.00\times10^{-14}}{5.01\times10^{-10}} = 2.00\times10^{-5} \; M$

6. Boric acid, H_3BO_3 is commonly used in eyewash solutions in chemistry laboratories to neutralize bases splashed in the eye. It acts as a monoprotic acid, but the dissociation reaction is slightly different from that of other acids.

$$B(OH)_{3\,(aq)} + H_2O_{(l)} \leftrightarrow B(OH)_4{}^{-1}{}_{(aq)} + H^+{}_{(aq)} \qquad K_a = 5.8 \times 10^{-10}$$

Calculate the pH of a 0.50 *M* solution of boric acid

Answers: ***WEAK ACID …***

	$[\,B(OH)_3\,]$		$[\,H_2O\,]$	$\leftrightarrow$	$[\,B(OH)_4{}^{-1}\,]$	+	$[\,H^{+1}\,]$
I	0.5 *M*		-		0		0
C	- x		-		+ x		+ x
E	0.5 *M* - x		-		x		x

$$K_a = 5.8 \times 10^{-10} = \frac{\left[\,B(OH)_4^{-1}\,\right]\left[\,H^{+1}\,\right]}{\left[\,B(OH)_3\,\right]^2} = \frac{x^2}{0.50 - x} = \frac{x^2}{0.50}$$

$$5.8 \times 10^{-10}\,(\,0.50\,) = x^2 \;;\; x = \sqrt{2.90 \times 10^{-10}} = 1.70 \times 10^{-5}\ M$$

$$[\,H^{+1}\,] = 1.70 \times 10^{-5}\ M, \qquad \mathbf{pH} = -\log(\,1.70 \times 10^{-5}\,) = \mathbf{4.77}$$

7. Hypobromous acid, HOBr , is a weak acid that dissociates in water, as represented in the equation below:

$$HOBr_{(aq)} \leftrightarrow H^+{}_{(aq)} + OBr^-{}_{(aq)} \qquad K_a = 2.3 \times 10^{-9}$$

Calculate the pH of a 2.50 *M* HOBr solution

Answers: ***WEAK ACID …***

	$[\,HOBr\,]$	$\leftrightarrow$	$[\,H^{+1}\,]$	+	$[\,OBr^{-1}\,]$
I	2.50 *M*		0		0
C	- x		+ x		+ x
E	2.50 *M* - x		x		x

$$K_a = 2.3 \times 10^{-9} = \frac{\left[\,OBr^{-1}\,\right]\left[\,H^{+1}\,\right]}{\left[\,HOBr\,\right]^2} = \frac{x^2}{2.50 - x} = \frac{x^2}{2.50}$$

$$2.3 \times 10^{-9}\,(\,2.50\,) = x^2 \;;\; x = \sqrt{5.75 \times 10^{-9}} = 7.58 \times 10^{-5}\ M$$

$$[\,H^{+1}\,] = 7.58 \times 10^{-5}\ M, \qquad \mathbf{pH} = -\log(\,7.58 \times 10^{-5}\,) = \mathbf{4.12}$$