What is the pH of 0.50 M HCl?

Name:________________________
Blk:____Date:__________________

$HCl + H_2O(l) \rightarrow H_3O^+ + Cl^-$

0.50   0.50   0.50

$pH = -\log(0.50)$

$= \boxed{0.30}$

$pH = -\log[H_3O^+]$

CHEMISTRY 12
ACID BASES UNIT
Lesson #12

Calculations Involving K_a and pH

IMPT: The pH scale is NOT LINEAR!!! When the pH is increased by 1, the [H_3O^+] is decreased by 10!!!

Recall that Strong Acids IONIZE 100 %, therefore the concentration of the strong acid will eqaul the concentration of H_3O^+ in solution!!

HOWEVER, Weak acids DO NOT IONIZE 100% in water, thereforewe must use an RICE table ! to determine the [H_3O^+] that is actually present in solution!

Generic Equation for a WEAK ACID in water:

$$HA + H_2O \longleftrightarrow A^- + H_3O^+$$

There are THREE types of problems that you can solve associated with a weak acid:

Type 1. Given the K_a and the concentration of the weak acid, solve for the pH

Example 1. What is the pH for a 0.500 M solution of CH_3COOH?

Step 1 . Write out the ionization equation with water

$CH_3COOH + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$

Step 2. Write out the K_a expression, identify the Ka value

$$K_a = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]} = 1.8 \cdot 10^{-5} \text{ (table)}$$

Step 3

R	CH_3COOH	+ ~~$H_2O(l)$~~	$\rightleftharpoons CH_3COO^-$	+ H_3O^+
I	0.500		Ø	Ø
C	$-x$		$+x$	$+x$
E	$0.500-x$		x	x

Let $0.500 - x \cong 0.500$

Step 4. Plug values into K_a and solve for x

$$0.500\,(1.8 \cdot 10^{-5}) = \left(\frac{x^2}{\cancel{0.500}}\right)\cancel{0.500} \rightarrow \sqrt{0.500\,(1.8 \cdot 10^{-5})} = \sqrt{x^2}$$

$$3.0 \cdot 10^{-3} = x = [H_3O^+]_E$$

Step 5. Use x to solve for pH

$$pH = -\log(3.0 \cdot 10^{-3}) = \boxed{2.52}$$

Type 2. Given the pH and the initial concentration of the weak acid, solve for the K_a value.

Example 2. If the pH of a 0.100 M solution of an unknown weak acid is 1.70, determine the K_a and identify the weak acid.

Step 1 . Write out the generic ionization with water

$$HA + H_2O(l) \rightleftharpoons A^- + H_3O^+$$

Step 2. Set up an ICE TABLE for the weak acid

R	HA	+ $\cancel{H_2O(l)}$	$\rightleftharpoons A^-$	+ H_3O^+
I	0.100		∅	∅
C	$-1.99\cdot10^{-2}$		$+1.99\cdot10^{-2}$	$+1.99\cdot10^{-2}$
E	$8.0\cdot10^{-2}$		$1.99\cdot10^{-2}$	$\boxed{1.99\cdot10^{-2}}$

Step 3. Use the pH to fill in the values of the ICE TABLE

$$[H_3O^+]_E = \text{antilog}(-pH)$$
$$= \text{antilog}(-1.70)$$
$$\sim 1.99\cdot10^{-2}$$

Step4. Use the equilibrium values to calculate the K_a

$$K_a = \frac{(1.99\cdot10^{-2})^2}{(0.100-1.99\cdot10^{-2})} \approx 5.0\cdot10^{-3}$$

$K_a <$ Citric acid, Iron III

Type 3. Given pH and K_a, determine the initial concentration of the weak acid.

Example 3. What concentration of H_2CO_3 is required to produce a pH of 4.18?

Step 1 . Write out the ionization with water

$$H_2CO_3 + H_2O(l) \rightleftharpoons HCO_3^- + H_3O^+$$

Step 2. Write out the K_a expression, look up K_a value

$$K_a = \frac{[HCO_3^-][H_3O^+]}{[H_2CO_3]} = 4.3\cdot10^{-7} \text{ (table)}$$

Step 3.

R	H_2CO_3	+ $\cancel{H_2O(l)}$	$\rightleftharpoons HCO_3^-$	+ H_3O^+
I	✗		∅	∅
C	-0.000066		$+6.6\cdot10^{-5}$	$+6.6\cdot10^{-5}$
E	let $x - 0.000066$ ≅ ✗		$6.6\cdot10^{-5}$	$\boxed{6.6\cdot10^{-5}}$

Step 4. Use pH to fill in the table

$$[H_3O^+] = \text{antilog}(-pH) \rightarrow \text{antilog}(-4.18) = 6.6\cdot10^{-5}$$

Step 5. Use K_a to solve for initial concentration

$$x(4.3\cdot10^{-7}) = \left(\frac{(6.6\cdot10^{-5})^2}{x}\right)x$$

$$\frac{x\cdot4.3\cdot10^{-7}}{4.3\cdot10^{-7}} = \frac{(6.6\cdot10^{-5})^2}{4.3\cdot10^{-7}}$$

$$\boxed{0.010\ M\ H_2CO_3}$$

0.000066

0.0066
0.000066

Seatwork/Homework: Exercises 74-78, 80, 82-83

PLO's: Part of M3 and M5 for K_a