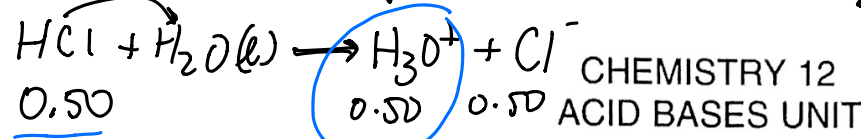


What is the pH of 0.50 M HCl? Name: _____
 Blk: _____ Date: _____



CHEMISTRY 12
 ACID BASES UNIT
 Lesson #12

$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log(0.50) = 0.30$$

Calculations Involving K_a and pH

IMPT: The pH scale is NOT LINEAR!!! When the pH is increased by 1, the $[\text{H}_3\text{O}^+]$ is decreased by 10!!!

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

HOWEVER, Weak acids DO NOT IONIZE 100% in water, therefore we must use an RICE table! to determine the $[\text{H}_3\text{O}^+]$ that is actually present in solution!

Generic Equation for a WEAK ACID in water:

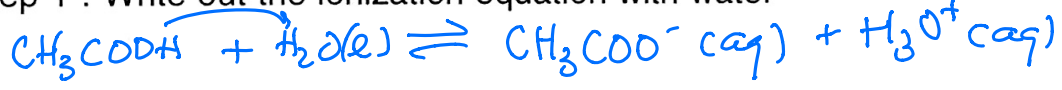


There are THREE types of problems that you can solve associated with a weak acid:

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



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

$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]} = 1.8 \cdot 10^{-5} \text{ (table)}$$

Step 3

R	CH_3COOH	$+$	$\text{H}_2\text{O}(\ell)$	\rightleftharpoons	CH_3COO^-	$+$	H_3O^+
I	0.500				0		0
C	$-x$				$+x$		$+x$
E	$0.500 - x$				x		x

Let $0.500 - x \approx 0.500$

Step 4. Plug values into K_a and solve for x

$$0.500 (1.8 \cdot 10^{-5}) = \frac{x^2}{0.500} \rightarrow \sqrt{0.500 (1.8 \cdot 10^{-5})} = \sqrt{x^2} = x = [\text{H}_3\text{O}^+]_E$$

$3.0 \cdot 10^{-3}$

Step 5. Use x to solve for pH

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

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



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

R	HA + H ₂ O(l)	⇌	A ⁻	+ H ₃ O ⁺
I	0.100		0	0
C	-1.99·10 ⁻²		+1.99·10 ⁻²	+1.99·10 ⁻²
E	8.0·10 ⁻²		1.99·10 ⁻²	1.99·10 ⁻²

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

$$[H_3O^+]_E = \text{antilog}(-pH)$$

$$= \text{antilog}(-1.70)$$

$$\approx 1.99 \cdot 10^{-2}$$

Step 4. Use the equilibrium values to calculate the K_a

$$K_a = \frac{(1.99 \cdot 10^{-2})^2}{(0.100 - 1.99 \cdot 10^{-2})} = 5.0 \cdot 10^{-3}$$

Citric acid
Iron III

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

Example 3. What concentration of H₂CO₃ is required to produce a pH of 4.18?

Step 1. Write out the ionization with water



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

$$K_a = \frac{[HCO_3^-][H_3O^+]}{[H_2CO_3]} = 4.3 \cdot 10^{-7} \text{ (table)}$$

R	H ₂ CO ₃ + H ₂ O(l)	⇌	HCO ₃ ⁻	+ H ₃ O ⁺
I	X		0	0
C	-0.000066		+6.6·10 ⁻⁵	+6.6·10 ⁻⁵
E	let X = 0.000066		6.6·10 ⁻⁵	6.6·10 ⁻⁵

Step 4. Use pH to fill in the table

$$[H_3O^+] = \text{antilog}(-pH) \rightarrow \text{antilog}(-4.18) = 6.6 \cdot 10^{-5}$$

Step 5. Use K_a to solve for initial concentration

$$4.3 \cdot 10^{-7} = \frac{(6.6 \cdot 10^{-5})^2}{X}$$

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

PLO's: Part of M3 and M5 for K_a

$$X \cdot 4.3 \cdot 10^{-7} = \frac{(6.6 \cdot 10^{-5})^2}{4.3 \cdot 10^{-7}}$$

0.000066
0.000066
0.000066
1.0·10⁻² M H₂CO₃