Covid 11 Covid Chem

1) Cont K_A notes & problems

2) K_A lab Ms Purcell. = due Friday

3) Predicting the Direction of Acid-Base Rxns Using K_A Values

2) Ka?

[HF] = 0.267M   pH 5.62.   ~~$H_2S \leftrightharpoons H^+ + HS^-$~~

Hydrofluoric A

~~HCl~~

$$HF_{(aq)} \leftrightharpoons \overset{2.4\times10^{-6}M}{H^+_{(aq)}} + \overset{2.4\times10^{-6}M}{F^-_{(aq)}}$$

$$? K_A = \frac{[H^+][F^-]}{[HF]} = \frac{(2.4\times10^{-6})^2}{(0.267)} = 2.2\times10^{-11}$$

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

$5.62 = ?$ 2nd fxn shift $\log(-5.62)$

↑

2sf

$\therefore [H^+] = 2.4\times10^{-6}\ \frac{mol}{L}$

$$\%\text{ ion} = \frac{[H^+]}{[HF]} \cdot 100 = \frac{(2.4\times10^{-6}M)}{(0.267M)} \times 100$$

$$= 8.9\times10^{-4}\ \%$$

$\leq 5\%$

NO RICE TABLE

Normally we assume the acid is so weak that there is so little change to the concentration of the acid at equilibrium that we can safely use the initial concentration of the acid. as "constant"

formi methanoic acid

-COOH carboxylic acid group

(a) Calculate K_a for formic acid at this temperature.

HCOOH

Acid Ionization Equation: $HCOOH_{(aq)} \rightleftharpoons H^+_{(aq)} + HCOO^-_{(aq)}$

0.10 M 4.2×10^{-3} M 4.2×10^{-3} M

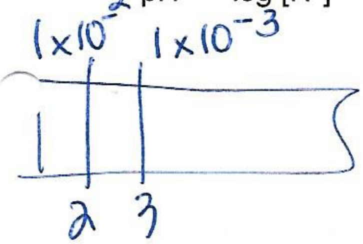
Ka Expression: $= \dfrac{[H+][HCOO^-]}{[HCOOH]} = \dfrac{(4.2 \times 10^{-3})(4.2 \times 10^{-3})}{(0.10)} = ?$

$1.8 \times 10^{-4} = K_A$

[HCOOH] ← 0.10 M

pH = - log [H⁺]   pH 2.38   2 sf

$pH = -\log[H+]$ $[H+] = 4.17 \times 10^{-3}$ mol/L (4.2)



$1 \times 10^{-2} > 4.17 \times 10^{-3} > 1 \times 10^{-3}$

pH 2 pH 3

(b) What percentage of the acid is ionized in a 0.10 *M* solution?

$\% \text{ ion} = \dfrac{[H+]}{[HCOOH]} \times 100 = \left(\dfrac{4.2 \times 10^{-3} M}{0.10 M}\right) \times 100$

$= 4.2\% < 5\%$

5 % Rule:

If you calculate the percent ionization to be greater than 5 % ($\geq$) you must then repeat the calculations using a **RICE** table! You must subtract the concentration of hydrogen ion from the initial concentration of the acid.

RICE

100 \$ → 97 \$ % 3

X 95 \$ 5% X

initially

Problem 2: Hypothetical acid HX with a concentration of 0.025 mol/L has a pH of 2.82.

at $\leftrightharpoons$!!

What is the Ka of HX?

$HX_{(aq)} \rightleftharpoons H^+_{(aq)} + X^-_{(aq)}$

$K_A = \frac{[H^+][X^-]}{[HX]}$ $K_a < 1$

Percent Ionization = $\frac{[H^+]}{[HX]} \times 100 = \frac{(1.5 \times 10^{-3} M)}{(0.025 M)} \times 100 = 6\%$

$pH = -\log[H^+]$ 2nd log (-2.82) $\therefore [H^+] = 1.5 \times 10^{-3} M$

R	HX $\rightleftharpoons$	H^+ +	X^-
I	0.025 M	$\emptyset$	$\emptyset$
BCE C MR	-1.5×10^{-3} M	$+1.5 \times 10^{-3}$ M	$+1.5 \times 10^{-3}$ M BCE
E	0.0235 M	1.5×10^{-3} M	1.5×10^{-3} M

Ka = $\frac{[H^+][X^-]}{[HX]} = \frac{(1.5 \times 10^{-3})^2}{(0.0235)} = 9.6 \times 10^{-5}$

Problem 3: HW

Consider acetic acid, $HC_2H_3O_2$, the substance responsible for the characteristic odor and acidity of vinegar. Let's calculate the pH of a 0.30 *M* solution of acetic acid at 25.0 °C after referring to a Ka table -- the Ka of acetic acid is 1.8 $\times$ 10^{-5}. < 1

5% v/v acetic acid vinegar dilute soln of a WA.

Base Dissociation Constant (Kb)

A similar equilibrium exists when a weak base is dissolved in water.

The base will accept a proton from water and form a conjugate acid, BH^+.

disc

$$B + H_2O(l) \leftrightarrow OH^-(aq) + BH^+(aq)$$

M-OH

This equilibrium has its own special constant, K_b, known as the **base dissociation constant**.

Like the acid dissociation constant, it is defined as the equilibrium constant multiplied by the concentration of water.

Sodium hydroxide

Arrh B

\+ −

dissociation

$NaOH(s) + H_2O(l)$	$\rightleftharpoons$	$Na^+(aq) + OH^-(aq)$
Virtually 0% at equilibrium		Virtually 100% at equilibrium

$NaOH(aq) \rightleftharpoons Na^+(aq) + OH^-(aq)$

Kb = $[Na^+][OH^-] = K_{sp}$

Ammonia

BL

B p^+ acc

A p^+ donor

ionization

$NH_3(g) + H_2O(l)$	$\rightleftharpoons$	$NH_4^+(aq) + OH^-(aq)$
> 99%		< 1%

Ø conc

Kb = $\dfrac{[NH_4^+][OH^-]}{[NH_3]}$

$\Delta [g]$

$\Delta [aq]$

The Ka of Boric Acid

Purpose:

- To determine the Ka of a weak acid, H_3BO_3

Materials:

- 2 Burets—READ TO 2 DECIMAL PLACES!!
- Erlenmeyer flask
- White paper
- BTB or ~~PHTH~~
- Funnel
- Beaker of NaOH
- Boric Acid
- Various pH papers

*a titration is not an $\neq$ situation

* titration the run to completion

Procedure:

- Determine the pH of the boric acid
- Titrate the boric acid with NaOH

0.030 M

Observations:

A + B = S + W

A or B

0.030 M NaOH

10.00 mL H_3BO_3

BTB yellow in H_3BO_3 → green "neutralized"

1.50 mL NaOH start

31.80 mL NaOH end

pH 3.6 of H_3BO_3

1 sf! don't use for sig figs