

pH Review! 1. Find the pH of each solution. (You will need a Ka/Kb chart for some of these)

a. 0.30 M HNO₃ strong acid! [H⁺] = 0.30 M. pH = -log(.30) = **0.52**

b. 0.30 M NaNO₃ neutral! pH = **7.00**

c. 0.30 M LiOH strong base! [OH⁻] = 0.30 M pOH = -log(.30) = 0.52
pH = 14 - pOH = **13.48**

d. 0.30 M K₂O strong base, di basic. [OH⁻] = 2(.30) = 0.60 M pOH = 0.22
O²⁻ + H₂O → 2OH⁻ pH = **13.78**

e. 0.30 M Ca(OH)₂ strong base [OH⁻] = 2(.30) = 0.60 M, pH = **13.78**, as above

f. 0.30 M NaH strong base! H⁻ + H₂O → H_{2(g)} + OH⁻ [OH⁻] = 0.30 M pOH = 0.52
pH = **13.48**

g. 0.30 M HC₄H₇O₂ (butanoic acid) weak acid! must use a Ka. look up: Ka = 1.5 × 10⁻⁵
HC₄H₇O₂ ⇌ H⁺ + C₄H₇O₂⁻ Ka = $\frac{x^2}{.30-x} = 1.5 \times 10^{-5}$

assume x << .30, get x = 0.0021(2) M = [H⁺]
pH = -log(.0021(2)) = **2.67**

h. A solution containing butanoic acid dissolved at 0.30 M, and sodium butanoate dissolved at 0.20 M

HC₄H₇O₂ ⇌ H⁺ + C₄H₇O₂⁻ So [Na⁺] = .20 M, [C₄H₇O₂⁻] = 0.20 M
.30 - x x .20 + x

Ka = $\frac{x(.20+x)}{.30-x} = 1.5 \times 10^{-5}$ $\frac{x(.20)}{.30} = 1.5 \times 10^{-5}$ x = [H⁺] = 2.2(5) × 10⁻⁵ M
assume x << .30, .20 pH = **4.65**

i. 0.30 M sodium butanoate, NaC₄H₇O₂

Since butanoic acid is a weak acid, its conjugate base, butanoate, will be significantly basic.

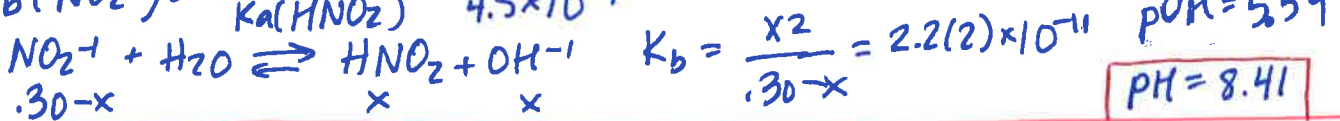
C₄H₇O₂⁻ + H₂O ⇌ HC₄H₇O₂ + OH⁻ Kb = $\frac{x^2}{.30-x} = \frac{K_w}{K_a(\text{HC}_4\text{H}_7\text{O}_2)}$
.30 - x x x = 6.67 × 10⁻¹⁰

j. 0.30 M HNO₂ weak acid Ka = 4.5 × 10⁻⁴
HNO₂ ⇌ H⁺ + NO₂⁻ assume x << .30, $\frac{x^2}{.30} = 6.67 \times 10^{-10} = 6.67 \times 10^{-10}$
.30 - x x x x = [OH⁻] = 0.00014(14) M. pOH = 4.85
pH = **9.15**
 $\frac{x^2}{.30-x} = 4.5 \times 10^{-4}$ $\frac{x^2}{.30} = 4.5 \times 10^{-4}$ $\frac{x^2}{.30-.01162} = 4.5 \times 10^{-4}$ [H⁺] = .01140
assume x << .30 x = .01162... x = .01139, .01140, pH = **1.94**

k. 0.30 Molar LiNO₂

Since HNO₂ is a weak acid, NO₂⁻ ions basic.

$$K_b(\text{NO}_2^-) = \frac{K_w}{K_a(\text{HNO}_2)} = \frac{1.00 \times 10^{-14}}{4.5 \times 10^{-4}} = 2.2(2) \times 10^{-11}$$

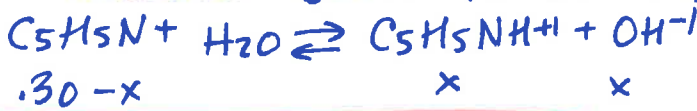


assume $x \ll .30$,
get $x = 2.5(8) \times 10^{-6} \text{ M}$
 $x = [\text{OH}^-]$
pOH = 5.59

PH = 8.41

l. 0.30 Molar pyridine, C₅H₅N

this is a base! $K_b = 1.7 \times 10^{-9}$ (looked up)



$$K_b = \frac{x^2}{.30-x} = 1.7 \times 10^{-9}$$

ass. $x \ll .30$
 $x = .000022(6) \text{ M} = [\text{OH}^-]$

pOH = 4.65, **PH = 9.35**

m. 0.30 M Pyridinium bromide, C₅H₅NHBr

this is the conjugate acid of pyridine; it is acidic.

$[\text{C}_5\text{H}_5\text{NH}^+] = .30 \text{ M}$, and $[\text{Br}^-] = .30 \text{ M}$

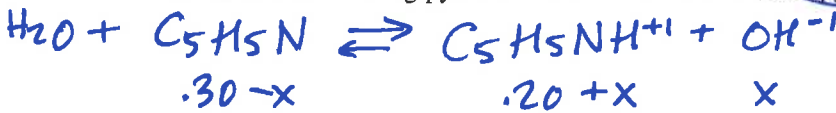


$$K_a(\text{C}_5\text{H}_5\text{NH}^+) = \frac{K_w}{K_b(\text{C}_5\text{H}_5\text{N})} = 5.8(82) \times 10^{-6}$$

$$\frac{x^2}{.30-x} = 5.882 \times 10^{-6} \quad x = [\text{H}^+] = .0013(3) \text{ M}$$

yoda yoda → **PH = 2.88**

n. A solution containing pyridine dissolved at 0.30 M, and pyridium bromide dissolved at 0.20 M.



pyridinium.
 $[\text{C}_5\text{H}_5\text{NH}^+] = .20 \text{ M}$, $[\text{Br}^-] = .20 \text{ M}$

$$K_b = \frac{(.20+x)(x)}{.30-x} = 1.7 \times 10^{-9}$$

assume $x \ll .3, .2$, $x = 2.5(5) \times 10^{-9} \text{ M} = [\text{OH}^-]$
pOH = 8.59, **PH = 5.41**

(or you could do this whole problem using K_a , w/ $[\text{C}_5\text{H}_5\text{NH}^+] \rightleftharpoons \text{H}^+ + (\text{C}_5\text{H}_5\text{N})!$)

o. A solution of 0.30 M KHCO₃,

$$K_a \text{ of } \text{HCO}_3^- = K_{a2} \text{ of } \text{H}_2\text{CO}_3 = 5.6 \times 10^{-11}$$

$$K_b \text{ of } \text{HCO}_3^- = \frac{K_w}{K_{a1} \text{ of } \text{H}_2\text{CO}_3} = \frac{1.00 \times 10^{-14}}{4.3 \times 10^{-7}} = 2.3(26) \times 10^{-8}$$

$K_b > K_a$ so HCO_3^- is a base.



$$K_b = 2.3(26) \times 10^{-8} = \frac{x^2}{.30-x}$$

$x = 8.3(5) \times 10^{-5} \text{ M} = [\text{OH}^-]$

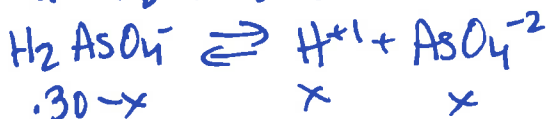
pOH = 4.08, **PH = 9.92**

p. A solution of 0.30 M NaH₂AsO₄.

$$K_a \text{ of } \text{H}_2\text{AsO}_4^- = K_{a2} \text{ of } \text{H}_3\text{AsO}_4 = 1.0 \times 10^{-7}$$

$$K_b \text{ of } \text{H}_2\text{AsO}_4^- = \frac{K_w}{K_{a1} \text{ of } \text{H}_3\text{AsO}_4} = \frac{1.00 \times 10^{-14}}{5.6 \times 10^{-3}} = 1.79 \times 10^{-12}$$

$K_a > K_b$ so it's acidic



$$\frac{x^2}{.30-x} = 1.0 \times 10^{-7}$$

$x \ll .30 \dots$

$x = .00017(32) \text{ M} = [\text{H}^+]$
PH = 3.76 !!!!!