8.4 - Calculations involving Acidic Solutions.

p.513, #1+2.

#1). $[HCl] = \cancel{0.0}\ 0.700\ mol/L = [H^+] = 0.700\ mol/L$
Required = $[OH^-]$

Formula Analysis:

$$[H^+][OH^-] = 1.0 \times 10^{-14}.$$

$$[OH^-] = \frac{1.0 \times 10^{-14}}{0.700}$$

$$[OH^-] = 1.43 \times 10^{-13}\ mol/L.$$

Statement:
The $[OH^-]$ is $1.43 \times 10^{-13}\ \frac{mol}{L}$.

#2). Given: $[HBr] = 0.03\overset{5}{\cancel{8}}0\ mol/L$; $pH = 2.78$.

$[H^+] = 0.035 mol/L.$ because $HBr = \frac{0.070\ mol}{2\ L} = 0.035\ mol/L.$

$pH = -\log[H^+]$.
$= -\log(3.5 \times 10^{-2}) = 1.46$.

$pOH = 14.00 - 1.46$
$= 12.54$

The pH is (1.46) and pOH is 12.54.
for the HBr solution.

8.4.

p.516, #1 + 2.

#1. Given: $[HC_3H_5O_{2(aq)}] = 0.050$ mol.
$pH = 2.78$
Required: percent ionization
Analysis: $[H^+_{(aq)}] = 10^{-2.78}$
$= 0.00166$

$\frac{\text{percent ionization}}{100} = \frac{0.00166}{0.050} \times 100\%$

percent ionization $= 3.3\%$.

#2. Given: $[HF] = 0.100$ mol/L
% ionization $= 7.8\%$.
Analysis:
HF concentration (7.8%)(0.100 mol/L)
$= 0.078 \times 0.100$ mol/L
$= 0.0078$ mol/L.

	$HF_{(aq)} \rightleftharpoons$	$H^+_{(aq)}$ +	$F^-_{(aq)}$
Initial	0.100	0	0
Change	−0.0078	+0.0078	+0.0078
Equilibrium	0.0922	+0.0078	+0.0078

$Ka = \frac{(0.0078)(0.0078)}{(0.0922)}$

$Ka = 6.6 \times 10^{-4}$.

Statement:
The Ka for HF is 6.6×10^{-4}

p.S20, #1+2

1. Given. $[HCN] = 6.18 \times 10^{-3}$.

$K_A = 6.2 \times 10^{-10}$.

pH = ?

$$HCN_{(aq)} \rightleftharpoons H^+_{(aq)} + CN^-_{(aq)}$$

I.	6.18×10^{-3}	0	0
C.	$-x$	x	x
E.	$6.18 \times 10^{-\cancel{13}3} - x$	x	x

$$6.2 \times 10^{-10} = \frac{(x)(x)}{(6.18 \times 10^{\cancel{13}} - x)}$$

$$\frac{6.18 \times 10^{-\cancel{13}}}{6.2 \times 10^{-10}} < 100$$

$9,967741 < 100$

$\therefore 6.18 \times 10^{-3}$

$$6.2 \times 10^{-10} = \frac{x^2}{\cancel{6.2 \times 10}\ 6.18 \times 10^{3}}$$

$x^2 = 3.83 \times 10^{-12}$

$x = [H^+(aq)]$

$[H^+] = 1.96 \times 10^{-6}$ mol/L.

$\boxed{pH = +5.71}$

100th Rule.

$\frac{\text{concentration}}{K} > 100$

Can eliminate $6.18 \times 10^{\cancel{13}} x$ to just 6.18×10^{-13}

This simplifies the calculation.

#2 Given. $[HNO_2] = 0.10$ mol/L ; $K_A = 4.6\times10^{-4}$.
Required pH =

$$HNO_2 \rightleftharpoons H^+ + NO_2^-(aq).$$

I	0.10	0	0
C	$-x$	x	x
E	$0.10-x$	x	x

$$Ka = \frac{[H^+][NO_2^-]}{[HNO_2]}$$

$$4.6\times10^{-4} = \frac{(x)(x)}{(0.10-x)}$$

$$4.6\times10^{-4} = \frac{x^2}{0.10}$$

eliminate using 100-rule.

$\frac{C}{K} > 100.$

$\frac{0.10}{4.6\times10^{-4}} > 100.$

$217.39 > 100$

$\therefore$ 0.01 − x is 0.01.

$x^2 = 4.6\times10^{-5}$

$x = [H^+]$

$x = 6.78\times10^{-3}$ mol/L.

$pH = -\log(6.78\times10^{-3})$.

$pH = 2.17$.

8.5 p. 527, #1 + 2.

Given:

#1. $[KOH] = 0.00100$ mol/L

Required: $[H^+]$, $[OH^-(aq)]$.

$[H^+][OH^-] = 1.0\times10^{-14}$.

$[OH^-] = 0.00100$ mol/L.

$$[H^+] = \frac{1.0\times10^{-14}}{0.00100}$$

$[H^+] = 1.00\times10^{-11}$ mol/L ; $[OH^-] = 0.00100$ mol/L.

(8.5 calculation)

#2 $[Sr(OH)_{2(aq)}] = 0.042\ mol / 2.00\ L = 0.021\ mol/L$

$2[OH^-]$

$2(0.021) = [OH^-] = 0.042\ mol/L$.

$pOH = -\log[OH^-]$

$= \cancel{1.688} - \log(0.042)$.

$pOH = 1.37$.

$\boxed{\frac{C}{K} > 100}$.

$pH = 14.00 - 1.37$
$= 12.62$.