

Chapter 8



when HBr is added the H^+ from the strong acid pushes the equilibrium towards the reactants (H_2O)



$$[H^+] = 1.0 \times 10^{-3} M \Rightarrow pH = \boxed{3}$$



$$[OH^-] = 1.0 \times 10^{-2}$$

$$pOH = 2 \Rightarrow pH = \boxed{12}$$

③ Dilute strong acid, use systematic method

mass balance: $5.0 \times 10^{-8} = [ClO_4^-]$

charge balance: $[H^+] = [OH^-] + [ClO_4^-]$

equilibrium: $[H^+][OH^-] = 10^{-14}$

unknowns: $[H^+], [OH^-], [ClO_4^-]$ equations: 3

Solve: $[OH^-] = \frac{10^{-14}}{[H^+]}$

$$[H^+] = \frac{10^{-14}}{[H^+]} + 5.0 \times 10^{-8}$$

$$[H^+]^2 - 5.0 \times 10^{-8} [H^+] - 10^{-14} = 0$$

$$[OH^-] = \frac{10^{-14}}{1.28 \times 10^{-7}} = 7.81 \times 10^{-8} M$$

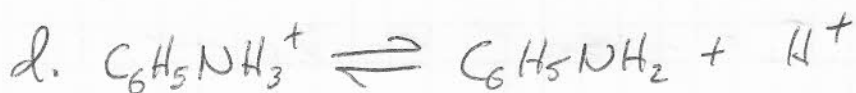
$$[H^+] = 1.28 \times 10^{-7} M \quad pH = \boxed{6.89}$$

$$\frac{7.81 \times 10^{-8}}{1.28 \times 10^{-7}} \times 100 = \boxed{61\%}$$

$H_2O \rightleftharpoons H^+ + OH^-$
amount of H^+ from H_2O



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$$\frac{[H^+][A^-]}{[HA]} = 1.00 \times 10^{-5}$$

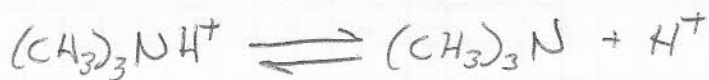
$$\frac{x^2}{0.100 - x} \overset{\text{ignore}}{=} 1.00 \times 10^{-5}$$

$$x = 1.0 \times 10^{-3} \quad \boxed{pH = 3}$$

	HA	H ⁺	A ⁻
I	0.100	0	0
C	-x	+x	+x
E	0.100-x	x	x

$$\alpha = \frac{1.0 \times 10^{-3}}{0.100} = \boxed{1.0 \times 10^{-2}} = 1\%$$

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$$\frac{[H^+][(CH_3)_3N]}{[(CH_3)_3NH^+]} = 1.59 \times 10^{-10}$$

$$\frac{x^2}{0.060 - x} \overset{\text{ignore}}{=} 1.59 \times 10^{-10}$$

$$x = 3.09 \times 10^{-6} M = [H^+]$$

	$(CH_3)_3NH^+$	H ⁺	$(CH_3)_3N$
I	0.060	0	0
C	-x	+x	+x
E	0.060-x	x	x

$pH = 5.51$

~~$[H^+] =$~~

$[(CH_3)_3N] = 3.09 \times 10^{-6} M$

$[(CH_3)_3NH^+] = 0.060 M$

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$$\frac{[H^+][A^-]}{[HA]} = 9.8 \times 10^{-5}$$

$$a. \frac{x^2}{10^{-2.00} - x} = 9.8 \times 10^{-5}$$

$$x^2 + 9.8 \times 10^{-5}x - 9.8 \times 10^{-7} = 0$$

$$x = 9.4 \times 10^{-4} = [H^+] \quad \boxed{pH = 3.03}$$

$$\alpha = \frac{9.4 \times 10^{-4}}{10^{-2.00}} \times 100 = \boxed{9.4\%}$$

b. To be the most precise use the systematic method - very difficult to solve in this case so assume $pH = \boxed{7.00} \Rightarrow [H^+] = 10^{-7}$

$$[HA] = \frac{[H^+][A^-]}{9.8 \times 10^{-5}} = \frac{10^{-7}[A^-]}{9.8 \times 10^{-5}} = 1.02 \times 10^{-3}[A^-]$$

$$\alpha = \frac{[A^-]}{[A^-] + [HA]} = \frac{[A^-]}{[A^-] + 1.02 \times 10^{-3}[A^-]} = \frac{1}{1 + 1.02 \times 10^{-3}} \times 100 = \boxed{99.9\%}$$



$$\frac{[OH^-][BH^+]}{[B]} = K_b = 1.00 \times 10^{-5}$$

$$\frac{x^2}{\cancel{1.0 \times 10^{-5}} - x} = 1.00 \times 10^{-5}$$

square

$$x = 1.0 \times 10^{-3} = [OH^-]$$

$$\alpha = \frac{1.0 \times 10^{-3}}{0.100} \times 100$$

$$= \boxed{1\%} \text{ or } 1.0 \times 10^{-2}$$

$$[H^+] = \frac{10^{-14}}{1.0 \times 10^{-3}} = 1.0 \times 10^{-11}$$

$$\boxed{pH = 11.0}$$

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$$\frac{[(CH_3)_3NH^+][OH^-]}{[(CH_3)_3N]} = \frac{10^{-14}}{1.59 \times 10^{-10}} = 6.3 \times 10^{-5}$$

$$\frac{x^2}{0.060 - x} = 6.3 \times 10^{-5}$$

$$x^2 + 6.3 \times 10^{-5}x - 3.78 \times 10^{-6} = 0$$

$$x = 1.91 \times 10^{-3} M = [OH^-] \quad pH = -\log \frac{10^{-14}}{1.91 \times 10^{-3}} = \boxed{11.3}$$

$$\begin{aligned} [(CH_3)_3NH^+] &= 1.91 \times 10^{-3} M \\ [(CH_3)_3N] &= 0.058 M \end{aligned}$$

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$$\frac{[CH_3CO_2H][OH^-]}{[CH_3CO_2^-]} = \frac{10^{-14}}{1.75 \times 10^{-5}} = 5.71 \times 10^{-10}$$

$$\frac{x^2}{0.100 - x} = 5.71 \times 10^{-10}$$

$$\alpha = \frac{7.56 \times 10^{-6}}{0.100} = \boxed{7.56 \times 10^{-5}} = \boxed{0.00756\%}$$

ignore

$$x = 7.56 \times 10^{-6}$$

$$\frac{x^2}{1 \times 10^{-2} - x} = 5.71 \times 10^{-10}$$

$$\alpha = \frac{2.39 \times 10^{-6}}{1 \times 10^{-2}} = \boxed{2.39 \times 10^{-4}} = \boxed{0.0239\%}$$

For $1.00 \times 10^{-2} M$ follow example of problem 13

$$[CH_3CO_2H] = \frac{5.71 \times 10^{-10} [CH_3CO_2^-]}{1.0 \times 10^{-7}} = 5.71 \times 10^{-3} [CH_3CO_2^-]$$

$$\alpha = \frac{5.71 \times 10^{-3} [CH_3CO_2^-]}{5.71 \times 10^{-3} [CH_3CO_2^-] + [CH_3CO_2^-]} = \boxed{5.68 \times 10^{-3}} = \boxed{0.57\%}$$

α increases with dilution