Chem 1062

Note Title 7/24/2008

* Last Day to Withdraw?
* D2L Quizzes

Find the pH of 0.10M NH_4NO_3.

acidic salt
- produce H_3O^+
- proton donor

$$NH_4NO_3(aq) \longrightarrow \boxed{NH_4^+(aq)} + NO_3^-(aq)$$

$$NH_4^+(aq) + H_2O(l) \rightarrow H_3O^+(aq) + NH_3(aq)$$

(equil M) $0.10 - x$ $\quad x \quad x$

$$\frac{K_w}{K_b} = K_a = \frac{[H_3O^+][NH_3]}{[NH_4^+]}$$

Assume $x << 0.10$ — valid assumption

$$\frac{1.0\times10^{-14}}{1.8\times10^{-5}} = 5.56\times10^{-10} = \frac{x^2}{0.10 \cancel{-x}}$$

$$\sqrt{5.56\times10^{-11}} = \sqrt{x^2} = x = 7.45\times10^{-6}M = [H_3O^+]$$

$$pH = -\log[H_3O^+] = -\log(7.45\times10^{-6}) = \boxed{5.13}$$

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(1) $\cancel{NH_4^+}(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + \cancel{NH_3}(aq)$ $\quad K_a$

(2) $\cancel{NH_3}(aq) + H_2O(l) \rightleftharpoons \cancel{NH_4^+}(aq) + OH^-(aq)$ $\quad K_b$

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(3) $H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$ $\quad K_w$

$K_aK_b = K_w$ — if conjugate pairs

$$K_a = \frac{K_w}{K_b} \qquad K_b = \frac{K_w}{K_a}$$

Find the pH of 0.10 M KCN.

basic salt
produces OH^-
proton acceptor

$$KCN_{(aq)} \rightarrow K^+_{(aq)} + \boxed{CN^-_{(aq)}}$$

$$CN^-_{(aq)} + H_2O_{(l)} \rightleftharpoons HCN_{(aq)} + OH^-_{(aq)}$$

(equil M) $0.10-x$ $\quad x \quad x$

$$\frac{K_w}{K_a} = K_b = \frac{[HCN][OH^-]}{[CN^-]}$$

Assume $x \ll 0.10$

Assumption valid

$$\frac{1.0\times10^{-14}}{4.9\times10^{-10}} = 2.04\times10^{-5} = \frac{x^2}{0.10\cancel{-x}}$$

$$\sqrt{2.04\times10^{-6}} = \sqrt{x^2} = x = 0.00143\ M = [OH^-]$$

$$pOH = -\log[OH^-] = -\log(0.00143) = 2.85$$

$$pH = 14.00 - pOH = 14.00 - 2.85 = \boxed{11.15}$$

Find the pH of 0.10 M NaCl.

$$\boxed{pH = 7.00}$$

Find the pH of 0.10 M HNO_3.

strong acid

$[H_3O^+] = 0.10M \quad -\log(0.10) = \boxed{1.00}$

Common Ion Effect

Find the pH of a solution that is both 0.20 M $HC_2H_3O_2$ & 0.10 M HCl.

end M

$$HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

H_3O^+: 0.10 M

$$HC_2H_3O_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + C_2H_3O_2^-(aq)$$

(equil M) $0.20 - x$ $\qquad$ $0.10 + x$ $\qquad$ x

weak acid & $\boxed{\text{strong acid}}$

pH due entirely to strong acid

$$K_a = \frac{[H_3O^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

$$1.7\times10^{-5} = \frac{(0.10\cancel{+x})(x)}{(0.20\cancel{-x})}$$

Assume $x \ll 0.10$

Assumption valid

$$\frac{1.7\times10^{-5}}{0.50} = \frac{\cancel{0.50}\,x}{\cancel{0.50}} = 3.4\times10^{-5} = x$$

$$0.10 + 3.4\times10^{-5} \cong 0.10\,M = [H_3O^+]$$

$$pH = -\log[H_3O^+] = -\log(0.10) = \boxed{1.00}$$

weak base & $\boxed{\text{strong base}}$

pH attributed entirely to the strong base

Find the pH of a solution that is both $0.10\,M\ NaC_2H_3O_2$ & $0.20\,M\ HC_2H_3O_2$.

basic salt

weak acid

conjugate pair

Buffer solution - contains weak acid & its conjugate base

" " base & " " acid

$$HC_2H_3O_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + C_2H_3O_2^-(aq)$$

(equil M) $0.20 - x$ $\qquad$ x $\qquad$ $0.10 + x$

$$K_a = \frac{[H_3O^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

$$1.7\times10^{-5} = \frac{(x)(0.10\cancel{+x})}{0.20\cancel{-x}}$$

Assume $x \ll 0.10$

$$\frac{1.7\times10^{-5}}{0.50} = \frac{\cancel{0.50}\,x}{\cancel{0.50}} = x = [H_3O^+] = 3.4\times10^{-5}\,M$$

$$pH = -\log[H_3O^+] = -\log(3.4\times10^{-5}) = \boxed{4.47}$$

Find the pH of a solution that is both $0.10\,M\ Na\boxed{C_2H_3O_2}$ & $0.20\,M\ \boxed{HC_2H_3O_2}$.

basic salt

weak acid

conjugate

$$C_2H_3O_2^-{}_{(aq)} + H_2O(l) \rightleftharpoons HC_2H_3O_{2\,(aq)} + OH^-{}_{(aq)}$$

(equil M) $\quad 0.10-x \qquad 0.20+x \qquad x$

$$\frac{K_w}{K_a} = K_b = \frac{[HC_2H_3O_2][OH^-]}{[C_2H_3O_2^-]}$$

Assume $x \ll 0.10$

$$\frac{1.0\times10^{-14}}{1.7\times10^{-5}} = 5.88\times10^{-10} = \frac{(0.20\cancel{+x})(x)}{(0.10\cancel{-x})}$$

$$\frac{5.88\times10^{-10}}{2.0} = \frac{\cancel{2.0}\,x}{\cancel{2.0}} = 2.94\times10^{-10} = [OH^-]$$

$$pOH = -\log[OH^-] = -\log(2.94\times10^{-10}) = 9.53$$

$$pH = 14.00 - pOH = 14.00 - 9.53 = \boxed{4.47}$$

Find the pH of a solution that is both $0.10M\ Na\boxed{C_2H_3O_2}$ & $0.20M\ \boxed{HC_2H_3O_2}$.

basic salt

conjugate

weak acid

$$\boxed{pH = pK_a + \log\frac{[base]}{[acid]}}$$

$$pH = -\log(1.7\times10^{-5}) + \log\left(\frac{0.10}{0.20}\right) = \boxed{4.47}$$

Find the ratio of $\overset{B}{[NH_3]}$ to $\overset{A}{[NH_4^+]}$ in a buffer solution that has a $pH = 10.00$.

$\hookrightarrow pOH = 4.00$

$$pOH = pK_b + \log\frac{[A]}{[B]}$$

$$4.00 = -\log(1.8\times10^{-5}) + \log\frac{[A]}{[B]}$$

$$4.00 = 4.745 + \log\frac{[A]}{[B]}$$

$$-0.745 = \log\frac{[A]}{[B]}$$

$$10 \qquad 10$$

$$0.18 = \frac{[A]}{[B]} \Rightarrow \frac{[B]}{[A]} = \boxed{5.6 \text{ to } 1 = \frac{[NH_3]}{[NH_4^+]}}$$

Na_2HPO_4 & NaH_2PO_4

HPO_4^{2-} & $H_2PO_4^-$

conjugate pairs

What is the pH of a solution prepared by mixing 65.0 mL of 0.10 M NH_3 with 25.0 mL of 0.20 M NH_4Cl?

NH_3

$$\frac{C_1V_1}{V_2} = \frac{C_2\cancel{V_2}}{\cancel{V_2}} = \frac{(0.10\,M)(65.0\,\cancel{mL})}{(90.0\,\cancel{mL})} = 0.0722\,M\,NH_3$$

NH_4^+

$$\frac{C_1V_1}{V_2} = C_2 = \frac{(0.20\,M)(25.0\,\cancel{mL})}{(90.0\,\cancel{mL})} = 0.0556\,M\,NH_4^+$$

$$pOH = pK_b + \log\frac{[A]}{[B]}$$

$$= -\log(1.8\times10^{-5}) + \log\left(\frac{0.0556\,M}{0.0722\,M}\right) = 4.63$$

$$pH = 14.00 - pOH = 14.00 - 4.63 = \boxed{9.37}$$

What is the pH of this buffer solution after the addition of 5.0 mL of 0.15 M HCl?

0.15 M H_3O^+

strong acid, produce H_3O^+

react with each other

90.0 mL of 0.0722 M NH_3 & 0.0556 M NH_4^+

$$H_3O^+(aq) + NH_3(aq) \rightarrow NH_4^+(aq) + H_2O(l)$$

(mol)	H_3O^+	NH_3	NH_4^+
Init	0.00075	0.00650	0.00500
Change	−0.00075	−0.00075	+0.00075
End	0	0.00575	0.00575

H_3O^+ $\quad 5.0\,\cancel{mL} \times \frac{0.15\text{ mol } H_3O^+}{1000\,\cancel{mL}} = 0.00075\text{ mol } H_3O^+$ initially

NH_3 $\quad 90.0\,\cancel{mL} \times \frac{0.0722\text{ mol } NH_3}{1000\,\cancel{mL}} = 0.00650\text{ mol } NH_3$ initially

NH_4^+ $\quad 90.0\,\cancel{mL} \times \frac{0.0556\text{ mol } NH_4^+}{1000\,\cancel{mL}} = 0.00500\text{ mol } NH_4^+$ initially

$$pOH = pK_b + \log\frac{[A]}{[B]}$$

mole ratio = molarity ratio

$$= -\log(1.8\times10^{-5}) + \log\left(\cancel{\frac{0.00575}{0.00575}}\right)$$

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$$pOH = 4.74 \quad + 0$$

$$pH = 14.00 - pOH = 14.00 - 4.74 = \boxed{9.26}$$