10. pH and pOH Calculations with Weak Acids and Bases

a) Weak Acids

i) Example: What is the pH of 0.400 M H_2S?

① Not 100% dissociated so $[H_2S] \neq [H_3O^+]$; so can't just pH = -log[0.400]

② The weak acid will form an equilibrium:

	H_2S + $H_2O_{(l)}$ ⇌	HS^- +	H_3O^+
I	0.400	0	0
C	-X	+X	+X
E	0.400 – X	X	X

③ $Ka = 9.1 \times 10^{-8} = \dfrac{[HS^-][H_3O^+]}{[H_2S]} = \dfrac{X^2}{0.400 - X}$

④ Assume 0.400 –X = 0.400, because it is a weak acid and % dissociation < 5%

⑤ $9.1 \times 10^{-8} = \dfrac{X^2}{0.400}$ $\quad X = 1.91 \times 10^{-4} = [H_3O^+]$

⑥ pH = -log[1.91×10^{-4}] = **3.719**

ii) Example: What is pH of 0.200 M CH_3COOH?

① Not 100% dissociated so $[CH_3COOH] \neq [H_3O^+]$; so can't just pH = -log[0.200]

② The weak acid will form an equilibrium:

	CH_3COOH + $H_2O_{(l)}$ ⇌	CH_3COO^- +	H_3O^+
I	0.200	0	0
C	-X	+X	+X
E	0.200 – X	X	X

③ $Ka = 1.8 \times 10^{-5} = \dfrac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]} = \dfrac{X^2}{0.200 - X}$

④ Assume 0.200 –X = 0.200, because it is a weak acid and % dissociation < 5%

⑤ $1.8 \times 10^{-5} = \dfrac{X^2}{0.200}$ $\quad X = 1.90 \times 10^{-3} = [H_3O^+]$

⑥ pH = -log[1.90×10^{-3}] = **2.721**

iii) Example: What is Ka for an unknown acid which gives a pH of 2.50 for a 0.300M solution?

① Equilibrium $[H_3O^+]$ = antilog(-2.50) = 3.2×10^{-3} M

② HA + $H_2O_{(l)}$ ⇌	A^- +	H_3O^+
I 0.300	0	0
C -3.2×10^{-3}	$+3.2 \times 10^{-3}$	$+3.2 \times 10^{-3}$
E 0.297	3.2×10^{-3}	3.2×10^{-3}

③ $Ka = \dfrac{[H_3O^+][A^-]}{[HA]} = \dfrac{(3.2 \times 10^{-3})^2}{0.297} = 3.4 \times 10^{-5}$

iv) Example: What is pH for a solution containing 0.300 M NH_4NO_3?

① $NH_4NO_3 \rightarrow NH_4^+ + NO_3^-$; NO_3^- is a spectator and NH_4^+ acts as a weak acid

② $NH_4^+ + H_2O \rightleftharpoons H_3O^+ + NH3$

	NH_4^+	H_2O	H_3O^+	NH3
I	0.300		0	0
C	-X		+X	+X
E	0.300 - X		X	X

③ $Ka = 5.6 \times 10^{-10} = \frac{[NH_3][H_3O^+]}{[NH_4^+]} = \frac{X^2}{0.300 - X}$

④ Assume 0.300 –X = 0.300, because it is a weak acid and % dissociation < 5%

⑤ $5.6 \times 10^{-5} = \frac{X^2}{0.300}$ $X = 1.30 \times 10^{-5} = [H_3O^+]$

⑥ pH = -log[1.30×10^{-5}] = **4.886**

b) Weak Bases

i) Example: What is the pOH and pH of 0.150 M NH_3?

① Not 100% dissociated so $[NH_3] \neq [OH^-]$

② $NH_3 + H_2O_{(l)} \rightleftharpoons OH^- + NH_4^+$

	NH_3	$H_2O_{(l)}$	OH^-	NH_4^+
I	0.150		0	0
C	-X		+X	+X
E	0.150 – X		X	X

③ Must 1st find the Kb. $Kb(NH_3) = \frac{Kw}{Ka(NH_4^+)} = \frac{1.00 \times 10^{-14}}{5.6 \times 10^{-10}} = 1.78 \times 10^{-5}$

④ $Kb = \frac{[OH^-][NH_4^+]}{[NH_3]} = 1.78 \times 10^{-5} = \frac{(X)^2}{0.150 - X}$

⑤ Assume 0.150 – X = 0.150

⑥ $1.78 \times 10^{-5} = \frac{X^2}{0.150}$ $X = 1.63 \times 10^{-3}$ M = $[OH^-]$

⑦ pOH = -log(1.63×10^{-3}) = **2.788**; pH = 14 – 2.788 = **11.212**

ii) Example: What is the pH of a solution containing 0.100 M $NaNO_2$?

① $NaNO_2 \rightarrow Na^+ + NO_2^-$; Na^+ is a spectator and NO_2^- acts as a weak base

② $NO_2^- + H_2O \rightleftharpoons OH^- + HNO_2$

	NO_2^-		OH^-	HNO_2
I	0.100		0	0
C	-X		+X	+X
E	0.100 - X		X	X

③ $Kb(NO_2^-) = \frac{Kw}{Ka(HNO_2)} = \frac{1.00 \times 10^{-14}}{4.6 \times 10^{-4}} = 2.17 \times 10^{-11}$

④ $Kb = \frac{[OH^-][HNO_2]}{[NO_2^-]} = 2.17 \times 10^{-11} = \frac{(X)^2}{0.100 - X}$.

⑤ Assume $0.100 - X = 0.100$

⑥ $2.17 \times 10^{-11} = \frac{X^2}{0.100}$ $X = 1.47 \times 10^{-6}$ M $= [OH^-]$

⑦ $pOH = -\log(1.47 \times 10^{-6}) =$ **5.833**; $pH = 14 - 5.833 =$ **8.167**

iii) Example: A 0.400 M solution of an unknown base, A^-, has a pH of 11.50. What is the K_b for A^-?

① $pOH = 14 = 11.50 = 2.50$

② Equilibrium $[OH^-] = \text{antilog}(-2.50) = 3.2 \times 10^{-3}$ M

③ $A^- + H_2O \rightleftharpoons OH^- + HA$

	A^-		OH^-	HA
I	0.400		0	0
C	-3.2×10^{-3}		$+3.2 \times 10^{-3}$	$+3.2 \times 10^{-3}$
E	0.397		3.2×10^{-3}	3.2×10^{-3}

④ $Kb = \frac{[OH^-][HA]}{[A^-]} = \frac{(3.2 \times 10^{-3})^2}{0.397} =$ **2.6×10^{-5}**

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