16.53a) $HBrO_2 \rightleftharpoons BrO_2^- + H^+$ $K_a = \frac{[BrO_2^-][H^+]}{[HBrO]}$

$HBrO_2 + H_2O \rightleftharpoons BrO_2^- + H_3O^+$ $K_a = \frac{[BrO_2^-][H_3O^+]}{[HBrO]}$

b) $C_2H_5COOH \rightleftharpoons C_2H_5COO^- + H^+$ $K_a = \frac{[C_2H_5COO^-][H^+]}{[C_2H_5COOH]}$

$C_2H_5COOH + H_2O \rightleftharpoons C_2H_5COO^- + H_3O^+$ $K_a = \frac{[C_2H_5COO^-][H_3O^+]}{[C_2H_5COOH]}$

16.55 pH = -log[H+] 2.44 = -log[H^+] H^+ = 0.00363M

	$CH_3CH(OH)COOH$	$\rightleftharpoons$	$CH_3CH(OH)COO^-$	+	H^+
I	0.10M		0 M		0 M
C	-0.00363M		+0.00363M		+0.00363M
E	0.09637M		0.00363M		0.00363M

$$K_a = \frac{[CH_3CH(OH)COO^-][H^+]}{CH_3CH(OH)COOH} = \frac{(0.00363)^2}{0.09637} = 1.4 \times 10^{-4}$$

16.57 $\% \text{ Ion} = \frac{[H+]}{[HA_{initial}]}$ $0.11 = \frac{[H+]}{[0.100M]}$ [H+] = 0.0110*M*

	$ClCH_2COOH$	$\rightleftharpoons$	$ClCH_2COO^-$	+	H^+
I	0.10M		0 M		0 M
C	-0.0110M		+0.0110M		+0.0110M
E	0.0890M		0.0110M		0.0110M

$$K_a = \frac{[ClCH_2COO^-][H^+]}{ClCH_2COOH} = \frac{(0.0110)^2}{0.0890} = 1.4 \times 10^{-3}$$

16.59 pH = -log[H+] 2.90 = -log[H^+] H^+ = 0.00126M

	CH_3COOH	$\rightleftharpoons$	CH_3COO^-	+	H^+
I	x		0 M		0 M
C	-0.00126M		+0.00126M		+0.00126M
E	x - 0.00126M		0.00126		0.00126

$$K_a = \frac{[CH_3COO^-][H^+]}{CH_3COOH} \qquad 1.8 \times 10^{-5} = \frac{(0.00126)^2}{x - 0.00126}$$

$1.8 \times 10^{-5}x - 2.27 \times 10^{-8} = 1.58 \times 10^{-6}$ $x = 0.089M$

16.61

	C_6H_5COOH	$\rightleftharpoons$	$C_6H_5COO^-$	+	H^+
I	0.050M		0 M		0 M
C	-x		+x		+x
E	0.050 - x		x		x

$$K_a = \frac{[C_6H_5COO^-][H^+]}{[C_6H_5COOH]} \qquad 6.3 \times 10^{-5} = \frac{x^2}{0.050 - x}$$

← $0.050 - x \approx 0.050$ because K_a is small

$x = 0.0018$ $\therefore$ $[C_6H_5COO^-]$ & $[H^+]$ = 0.0018M and

$[C_6H_5COOH] = 0.050 - 0.0018 = 0.048M$

16.63a

	C_2H_5COOH	$\rightleftharpoons$	$C_2H_5COO^-$	+	H^+
I	0.095M		0 M		0 M
C	-x		+x		+x
E	0.095 - x		x		x

$$K_a = \frac{[C_2H_5COO^-][H^+]}{[C_2H_5COOH]} \qquad 1.3 \times 10^{-5} = \frac{x^2}{0.095 - x}$$

$0.095 - x \approx 0.095$ because K_a is small

$x = 0.0011$

$pH = -\log[H^+]$ $\quad pH = -\log[0.0011]$ $\quad pH = 2.96$

16.63c

	C_5H_5N	+	H_2O	$\rightleftharpoons$	$C_5H_5NH^+$	+	OH^-
I	0.120M		———		0 M		0 M
C	-x		———		+x		+x
E	0.120 - x		———		x		x

$$K_b = \frac{[C_5H_5NH^+][OH^-]}{C_5H_5N} \qquad 1.7 \times 10^{-9} = \frac{x^2}{0.120 - x}$$

$0.120 - x \approx 0.120$ because K_b is small

$x = 1.4 \times 10^{-5}$

$pOH = -\log[OH^-]$ $\quad pOH = -\log[1.4 \times 10^{-5}]$ $\quad pOH = 4.85$

$pH + pOH = 14$ $\quad pH = 14 - pOH$ $\quad pH = 14 - 4.85 = 9.15$

16.65

	$HNC_7H_4SO_3$	$\rightleftharpoons$	$NC_7H_4SO_3^-$	+	H^+
I	0.10		0 M		0 M
C	-x		+x		+x
E	0.10 - x		x		x

$pK_a = -\log[K_a]$ $2.32 = -\log[K_a]$ $K_a = 0.00479$

$K_a = \frac{[NC_7H_4SO_3^-][H^+]}{HNC_7H_4SO_3}$ $4.79\text{x}10^{-3} = \frac{x^2}{0.10 - x}$

$x^2 + 0.00479x - 0.000479 = 0$ $x = 0.0196$

$pH = -\log[H^+]$ $pH = -\log[0.0196]$ $pH = 1.71$

16.67a

	HN_3	$\rightleftharpoons$	N_3^-	+	H^+
I	0.400		0 M		0 M
C	-x		+x		+x
E	0.400 - x		x		x

$K_a = \frac{[N_3^-][H^+]}{HN_3}$ $1.9 \text{ x } 10^{-5} = \frac{x^2}{0.400 - x}$ ← $0.400 - x \approx 0.400$ because K_a is small

$x = 2.8 \text{ x } 10^{-3}$ $\% \text{ Ion} = \frac{[2.8 \text{ x } 10^{-3}]}{[0.400]} = 0.70\%$

16.67b

$K_a = \frac{[N_3^-][H^+]}{HN_3}$ $1.9 \text{ x } 10^{-5} = \frac{x^2}{0.100 - x}$ ← $0.100 - x \approx 0.100$ because K_a is small

$x = 1.4 \text{ x } 10^{-3}$ $\% \text{ Ion} = \frac{[1.4 \text{ x } 10^{-3}]}{[0.100]} = 1.4\%$

16.63 a&b reflect the idea that the higher the initial concentration, the smaller the % ionization, as seen in 16.4

16.71 a) NH_3OH^+

b) When hydroxylamine acts as a base, the nitrogen accepts the protons. Nitrogen is an effective proton acceptor because of its unbonded electrons.

c) In hydroxylamine, both the oxygen and the nitrogen have unbonded electrons which could be shared with hydrogen. This said, nitrogen will more likely share electrons because its electronegativity is lower.

16.73a)

$(CH_3)_2NH + H_2O \rightleftharpoons (CH_3)_2NH_2^+ + OH^-$ $\quad K_b = \dfrac{[(CH_3)_2NH_2^+]\,[OH^-]}{[(CH_3)_2NH]}$

b)

$CO_3^{2-} + H_2O \rightleftharpoons HCO_3^- + OH^-$ $\quad K_b = \dfrac{[HCO_3^-]\,[OH^-]}{[CO_3^{2-}]}$

c)

$CHOO^- + H_2O \rightleftharpoons CHOOH + OH^-$ $\quad K_b = \dfrac{[CHOOH]\,[OH^-]}{[CHOO^-]}$

16.75

	$C_2H_5NH_2$	+	H_2O	$\rightleftharpoons$	$C_2H_5NH_3^+$	+	OH^-
I	0.075M		———		0 M		0 M
C	-x		———		+x		+x
E	0.075 - x		———		x		x

$K_b = \dfrac{[C_2H_5NH_3^+][OH^-]}{[C_2H_5NH_2]}$ $\quad 6.4 \times 10^{-4} = \dfrac{x^2}{0.075 - x}$ ← $0.075 - x \approx 0.120$ because K_a is small

$x = 0.0069$

pOH = -log[OH^-] $\quad$ pOH = -log[0.0069] $\quad$ pOH = 2.16

pH + pOH = 14 $\quad$ pH = 14 - pOH $\quad$ pH = 14 - 2.16 = 11.84

16.77 pH + pOH = 14 pOH = 14 - pH pOH = 14 - 11.33 = 2.67

pOH = -log[OH-] 2.67 = -log[OH-] OH^- = 0.0021M

	$C_{10}H_{15}ON$ +	H_2O	⇌ $C_{10}H_{15}ONH^+$ +	OH^-
I	0.035	———	0 M	0 M
C	-0.0021	———	+0.0021	+0.0021
E	0.0329	———	0.0021	0.0021

$$K_b = \frac{[C_{10}H_{15}ONH^+][OH^-]}{[C_{10}H_{15}ON]} = \frac{(0.0021)^2}{(0.0329)} = 0.00013$$

16.79 a) $C_6H_5OH \leftrightharpoons C_6H_5O^- + H^+$

b)

$K_a \times K_b = K_w$

$1.3\times10^{-10} \times K_b = 1.0\times10^{-14}$

$K_b = 7.7\times10^{-5}$

c) Phenol is a weak acid than water, as indicated by it's Ka value being less that 1×10^{-7}.

16.81a) Acetic acid (K_a=1.8×10^{-5}) is a stronger weak acid than hypochlorous acid (K_a=3.0×10^{-8}).

b) Weaker weak acids have stronger conjugate bases. As such, the conjugate base of the weaker hypochlorous acid (hypochlorite ion) is the stronger base.

c)

$K_a \times K_b = K_w$

$1.8\times10^{-5} \times K_b = 1.0\times10^{-14}$

$K_b = 5.6\times10^{-10}$

$K_a \times K_b = K_w$

$3.0\times10^{-8} \times K_b = 1.0\times10^{-14}$

$K_b = 3.3\times10^{-7}$