15.28 The factors affecting the ease with which the hydrogen will be donated (and therefore acidic) are the electronegativity of the element Y and the number of oxygen atoms attached to the element Y.

15.29 A Lewis acid is an electron pair acceptor. A Lewis base is an electron pair donor.

15.30 A Lewis acid has an empty orbital (or can rearrange electrons to create an empty orbital) that can accept an electron pair. A Lewis base has a lone pair of electrons it can donate to the Lewis acid.

15.31 The combustion of fossil fuels produces oxides of sulfur and nitrogen, which react with oxygen and water to form sulfuric and nitric acids. These acids then combine with rain to form acid rain. Acid rain is a significant problem in the northeastern United States.

15.32 Acid rain corrodes man-made structures and also damages aquatic environments and forests. Environmental legislation has helped stabilize the amount of acid rain being produced.

Problems by Topic

The Nature and Definitions of Acids and Bases

15.33 (a) acid $HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$

(b) acid $NH_4^+(aq) \rightarrow H^+(aq) + NH_3(aq)$

(c) base $KOH(aq) \rightarrow K^+(aq) + OH^-(aq)$

(d) acid $HC_2H_3O_2(aq) \rightarrow H^+(aq) + C_2H_3O_2^-(aq)$

15.34 (a) base $NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$

(b) acid $H_2SO_4(aq) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$

(c) acid $HBr(aq) \rightarrow H^+(aq) + Br^-(aq)$

(d) base $Sr(OH)_2(aq) \rightarrow Sr^{2+}(aq) + 2OH^-(aq)$

15.35 (a) Since H_2CO_3 donates a proton to H_2O, it is the acid. After H_2CO_3 donates the proton, it becomes HCO_3^-, the conjugate base. Since H_2O accepts a proton, it is the base. After H_2O accepts the proton, it becomes H_3O^+, the conjugate acid.

(b) Since H_2O donates a proton to NH_3, it is the acid. After H_2O donates the proton, it becomes OH^-, the conjugate base. Since NH_3 accepts a proton, it is the base. After NH_3 accepts the proton, it becomes NH_4^+, the conjugate acid.

(c) Since HNO_3 donates a proton to H_2O, it is the acid. After HNO_3 donates the proton, it becomes NO_3^-, the conjugate base. Since H_2O accepts a proton, it is the base. After H_2O accepts the proton, it becomes H_3O^+, the conjugate acid.

(d) Since H_2O donates a proton to C_5H_5N, it is the acid. After H_2O donates the proton, it becomes OH^-, the conjugate base. Since C_5H_5N accepts a proton, it is the base. After C_5H_5N accepts the proton, it becomes $C_5H_5NH^+$, the conjugate acid.

15.36 (a) Since HI donates a proton to H_2O, it is the acid. After HI donates the proton, it becomes I^-, the conjugate base. Since H_2O accepts a proton, it is the base. After H_2O accepts the proton, it becomes H_3O^+, the conjugate acid.

(b) Since H_2O donates a proton to CH_3NH_2, it is the acid. After H_2O donates the proton, it becomes OH^-, the conjugate base. Since CH_3NH_2 accepts a proton, it is the base. After CH_3NH_2 accepts the proton, it becomes $CH_3NH_3^+$, the conjugate acid.

(c) Since H_2O donates a proton to CO_3^{2-}, it is the acid. After H_2O donates the proton, it becomes OH^-, the conjugate base. Since CO_3^{2-} accepts a proton, it is the base. After CO_3^{2-} accepts the proton, it becomes HCO_3^-, the conjugate acid.

(d) Since HBr donates a proton to H_2O, it is the acid. After HBr donates the proton, it becomes Br^-, the conjugate base. Since H_2O accepts a proton, it is the base. After H_2O accepts the proton, it becomes H_3O^+, the conjugate acid.

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

(b) HSO_3^- $H_2SO_3(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + HSO_3^-(aq)$

(c) CHO_2^- $HCHO_2(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + CHO_2^-(aq)$

(d) F^- $HF(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + F^-(aq)$

15.38 (a) NH_4^+ $NH_3(aq) + H_2O(l) \leftrightarrows NH_4^+(aq) + OH^-(aq)$

(b) $HClO_4$ $HClO_4(aq) + H_2O(l) \rightarrow H_3O^+(aq) + ClO_4^-(aq)$

(c) H_2SO_4 $H_2SO_4(aq) + H_2O(l) \rightarrow H_3O^+(aq) + HSO_4^-(aq)$

(d) HCO^- $HCO_3^-(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + CO_3^-(aq)$

15.39 $H_2PO_4^-(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + HPO_4^{2-}(aq)$

$H_2PO_4^-(aq) + H_2O(l) \leftrightarrows H_3PO_4(aq) + OH^-(aq)$

15.40 $HCO_3^-(aq) + HS^-(aq) \leftrightarrows H_2S(aq) + CO_3^{2-}(aq)$

$HCO_3^-(aq) + HS^-(aq) \leftrightarrows H_2CO_3(aq) + S^{2-}(aq)$

Acid Strength and K_a

15.41 (a) HNO_3 is a strong acid.

(b) HCl is a strong acid.

(c) HBr is a strong acid.

(d) H_2SO_3 is a weak acid. $H_2SO_3(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + HSO_3^-(aq)$

$$K_a = \frac{[H_3O^+][HSO_3^-]}{[H_2SO_3]}$$

15.42 (a) HF is a weak acid. $HF(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + F^-(aq)$

$$K_a = \frac{[H_3O^+][F^-]}{[HF]}$$

(b) $HCHO_2$ is a weak acid. $HCHO_2(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + CHO_2^-(aq)$

$$K_a = \frac{[H_3O^+][CHO_2^-]}{[HCHO_2]}$$

(c) H_2SO_4 is a strong acid.

(d) H_2CO_3 is a weak acid. $H_2CO_3(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + HCO_3^-(aq)$

$$K_{a_1} = \frac{[H_3O^+][HCO_3^-]}{[H_2CO_3]}$$

15.43 (a) contains no HA, 10 H^+, and 10 A^-

(b) contains 3 HA, 3 H^+, and 7 A^-

(c) contains 9 HA, 1 H^+, and 1 A^-

So, solution a > solution b > solution c.

15.44 HCl is a strong acid, $K_a(HF) = 3.5 \times 10^{-4}$, $K_a(HClO) = 2.9 \times 10^{-8}$, $K_a(HC_6H_5O) = 1.3 \times 10^{-10}$.

The larger the value of K_a the stronger the acid and the greater the $[H_3O^+]$.

The order of decreasing $[H_3O^+]$ is HCl > HF > HClO > HC_6H_5O.

15.45 (a) F^- is a stronger base than Cl^-.
F^- is the conjugate base of HF (a weak acid), Cl^- is the conjugate base of HCl (a strong acid), the weaker the acid, the stronger the conjugate base.

(b) NO_2^- is a stronger base than NO_3^-.
NO_2^- is the conjugate base of HNO_2 (a weak acid), NO_3^- is the conjugate base of HNO_3 (a strong acid), the weaker the acid, the stronger the conjugate base.

(c) ClO^- is a stronger base than F^-.
F^- is the conjugate base of HF ($K_a = 3.5 \times 10^{-4}$), ClO^- is the conjugate base of HClO ($K_a = 2.9 \times 10^{-8}$) HClO is the weaker acid, the weaker the acid, the stronger the conjugate base.

15.46 (a) ClO_2^- is a stronger base than ClO_4^-.
ClO_2^- is the conjugate base of $HClO_2$ (a weak acid), ClO_4^- is the conjugate base of $HClO_4$ (a strong acid), the weaker the acid, the stronger the conjugate base.

(b) H_2O is a stronger base than Cl^-.
H_2O is the conjugate base of H_3O^+, Cl^- is the conjugate base of HCl (a strong acid), the weaker the acid, the stronger the conjugate base.

(c) CN^- is stronger base than ClO^-.
CN^- is the conjugate base of HCN ($K_a = 4.9 \times 10^{-10}$), ClO^- is the conjugate base of HClO ($K_a = 2.9 \times 10^{-8}$), the weaker the acid, the stronger the conjugate base.

Autoionization of Water and pH

15.47 (a) **Given:** $K_w = 1.0 \times 10^{-14}$, $[H_3O^+] = 1.2 \times 10^{-8}$ M **Find:** $[OH^-]$
Conceptual Plan: $[H_3O^+] \rightarrow [OH^-]$
$K_w = 1.0 \times 10^{-14} = [H_3O^+][OH^-]$
Solution:
$K_w = 1.0 \times 10^{-14} = (1.2 \times 10^{-8})[OH^-]$
$[OH^-] = 8.3 \times 10^{-7}$M
$[OH^-] > [H_3O^+]$ so the solution is basic.

(b) **Given:** $K_w = 1.0 \times 10^{-14}$, $[H_3O^+] = 8.5 \times 10^{-5}$ M **Find:** $[OH^-]$
Conceptual Plan: $[H_3O^+] \rightarrow [OH^-]$
$K_w = 1.0 \times 10^{-14} = [H_3O^+][OH^-]$
Solution:
$K_w = 1.0 \times 10^{-14} = (8.5 \times 10^{-5})[OH^-]$
$[OH^-] = 1.2 \times 10^{-10}$M
$[H_3O^+] > [OH^-]$ so the solution is acidic.

(c) **Given:** $K_w = 1.0 \times 10^{-14}$, $[H_3O^+] = 3.5 \times 10^{-2}$ M **Find:** $[OH^-]$
Conceptual Plan: $[H_3O^+] \rightarrow [OH^-]$
$K_w = 1.0 \times 10^{-14} = [H_3O^+][OH^-]$
Solution:
$K_w = 1.0 \times 10^{-14} = (3.5 \times 10^{-2})[OH^-]$
$[OH^-] = 2.9 \times 10^{-13}$ M
$[H_3O^+] > [OH^-]$ so the solution is acidic.

15.48 (a) **Given:** $K_w = 1.0 \times 10^{-14}$, $[OH^-] = 1.1 \times 10^{-9}$ M **Find:** $[H_3O^+]$
Conceptual Plan: $[OH^-] \rightarrow [H_3O^+]$
$K_w = 1.0 \times 10^{-14} = [H_3O^+][OH^-]$

(c) **Given:** pH = 2.87 **Find:** $[H_3O^+]$, $[OH^-]$

Conceptual Plan: pH → $[H_3O^+]$ → $[OH^-]$

pH = $-\log[H_3O^+]$ $K_w = 1.0 \times 10^{-14} = [H_3O^+][OH^-]$

Solution: $pH = -\log[H_3O^+]$ $2.87 = -\log[H_3O^+]$

$-2.87 = \log[H_3O^+]$ $10^{-2.87} = 10^{\log[H_3O^+]}$

$10^{-2.87} = [H_3O^+]$ $[H_3O^+] = 1.3 \times 10^{-3}$

$K_w = 1 \times 10^{-14} = (1.3 \times 10^{-3})[OH^-]$

$[OH^-] = 7.4 \times 10^{-12}$ M

15.51 pH = $-\log[H_3O^+]$ $K_w = 1.0 \times 10^{-14} = [H_3O^+][OH^-]$

$[H_3O^+]$	$[OH^-]$	pH	Acidic or basic
7.1×10^{-4}	1.4×10^{-11}	**3.15**	acidic
3.7×10^{-9}	2.7×10^{-6}	8.43	basic
8×10^{-12}	1×10^{-3}	**11.1**	basic
6.2×10^{-4}	**1.6×10^{-11}**	3.20	acidic

$[H_3O^+] = 10^{-3.15} = 7.1 \times 10^{-4}$ $[OH^-] = \dfrac{1.0 \times 10^{-14}}{7.1 \times 10^{-4}} = 1.4 \times 10^{-11}$

$[OH^-] = \dfrac{1.0 \times 10^{-14}}{3.7 \times 10^{-9}} = 2.7 \times 10^{-6}$ $pH = -\log(3.7 \times 10^{-9}) = 8.43$

$[H_3O^+] = 10^{-11.1} = 8 \times 10^{-12}$ $[OH^-] = \dfrac{1.0 \times 10^{-14}}{8 \times 10^{-12}} = 1 \times 10^{-3}$

$[H_3O^+] = \dfrac{1.0 \times 10^{-14}}{1.6 \times 10^{-11}} = 6.2 \times 10^{-4}$ $pH = -\log(6.2 \times 10^{-4}) = 3.20$

15.52 pH = $-\log[H_3O^+]$ $K_w = 1.0 \times 10^{-14} = [H_3O^+][OH^-]$

$[H_3O^+]$	$[OH^-]$	pH	Acidic or basic
3.5×10^{-3}	2.9×10^{-12}	2.46	acidic
2.6×10^{-8}	**3.8×10^{-7}**	7.58	basic
1.8×10^{-9}	5.6×10^{-6}	8.74	basic
7.1×10^{-8}	1.4×10^{-7}	**7.15**	basic

$[OH^-] = \dfrac{1.0 \times 10^{-14}}{3.5 \times 10^{-3}} = 2.9 \times 10^{-12}$ $pH = -\log(3.5 \times 10^{-3}) = 2.46$

$[H_3O^+] = \dfrac{1.0 \times 10^{-14}}{3.8 \times 10^{-7}} = 2.6 \times 10^{-8}$ $pH = -\log(2.6 \times 10^{-8}) = 7.58$

$[OH^-] = \dfrac{1.0 \times 10^{-14}}{1.8 \times 10^{-9}} = 5.6 \times 10^{-6}$ $pH = -\log(1.8 \times 10^{-9}) = 8.74$

$[H_3O^+] = 10^{-7.15} = 7.1 \times 10^{-8}$ $[OH^-] = \dfrac{1.0 \times 10^{-14}}{7.1 \times 10^{-8}} = 1.4 \times 10^{-7}$

15.53 **Given:** $K_w = 2.4 \times 10^{-14}$ at 37°C **Find:** $[H_3O^+]$, pH

Conceptual Plan: K_w → $[H_3O^+]$ → pH

$K_w = [H_3O^+][OH^-]$ pH = $-\log[H_3O^+]$

Solution: $H_2O(l) + H_2O(l) \leftrightarrows H_3O^+(aq) + OH^-(aq)$

$K_w = [H_3O^+][OH^-]$

$[H_3O^+] = [OH^-] = \sqrt{K_w} = \sqrt{2.4 \times 10^{-14}} = 1.5 \times 10^{-7}$
$pH = -\log[H_3O^+] = -\log(1.5 \times 10^{-7}) = 6.81$

Check: The value of K_w increased indicating more products formed, so the $[H_3O^+]$ increases and the pH decreases from the values at 25°C.

15.54 The increasing value of K_w indicates more products are formed as the temperature increases. According to Le Châtelier, this means the heat is a reactant. Therefore, the autoionization of water is endothermic.

Acid Solutions

15.55 (a) **Given:** 0.25 M HCl (strong acid) **Find:** $[H_3O^+]$,$[OH^-]$, pH

Conceptual Plan: $[HCl] \rightarrow [H_3O^+] \rightarrow$ pH and then $[H_3O^+] \rightarrow [OH^-]$

$[HCl] \rightarrow [H_3O^+]$ $pH = -\log[H_3O^+]$ $[H_3O^+][OH^-] = 1.0 \times 10^{-14}$

Solution: $0.25\ M\ HCl = 0.25\ M\ H_3O^+$ $pH = -\log(0.25) = 0.60$
$[OH^-] = 1.0 \times 10^{-14}/0.25\ M = 4.0 \times 10^{-14}$

Check: HCl is a strong acid with a relatively high concentration, so we expect the pH to be low and the $[OH^-]$ to be small.

(b) **Given:** 0.015 M HNO_3 (strong acid) **Find:** $[H_3O^+]$,$[OH^-]$, pH

Conceptual Plan: $[HNO_3] \rightarrow [H_3O^+] \rightarrow$ pH and then $[H_3O^+] \rightarrow [OH^-]$

$[HNO_3] \rightarrow [H_3O^+]$ $pH = -\log[H_3O^+]$ $[H_3O^+][OH^-] = 1.0 \times 10^{-14}$

Solution: $0.015\ M\ HNO_3 = 0.015\ M\ H_3O^+$ $pH = -\log(0.015) = 1.82$
$[OH^-] = 1.0 \times 10^{-14}/0.015\ M = 6.7 \times 10^{-13}$

Check: HNO_3 is a strong acid, so we expect the pH to be low and the $[OH^-]$ to be small.

(c) **Given:** 0.052 M HBr and 0.020M HNO_3 (strong acids) **Find:** $[H_3O^+]$,$[OH^-]$, pH

Conceptual Plan: $[HBr] + [HNO_3] \rightarrow [H_3O^+] \rightarrow$ pH and then $[H_3O^+] \rightarrow [OH^-]$

$[HBr] + [HNO_3] \rightarrow [H_3O^+]$ $pH = -\log[H_3O^+]$ $[H_3O^+][OH^-] = 1.0 \times 10^{-14}$

Solution: $0.052\ M\ HBr = 0.052\ M\ H_3O^+$ and $0.020\ M\ HNO_3 = 0.020\ M\ H_3O^+$
Total $H_3O^+ = 0.052\ M + 0.020\ M = 0.072\ M$ $pH = -\log(0.072) = 1.14$
$[OH^-] = 1.0 \times 10^{-14}/0.072\ M = 1.4 \times 10^{-13}$

Check: HBr and HNO_3 are both strong acids and completely dissociate. This gives a relatively high concentration, so we expect the pH to be low and the $[OH^-]$ to be small.

(d) **Given:** $HNO_3 = 0.655\%$ by mass, $d_{solution} = 1.01\ g/mL$ **Find:** $[H_3O^+]$,$[OH^-]$, pH

Conceptual Plan:

% mass $HNO_3 \rightarrow$ g $HNO_3 \rightarrow$ mol HNO_3 and then g soln $\rightarrow$ mL soln $\rightarrow$ L soln $\rightarrow$ M HNO_3

$\frac{\%}{100}$ $\frac{mol\ HNO_3}{63.018\ g\ HNO_3}$ $\frac{1.01\ g\ soln}{mL\ soln}$ $\frac{1000\ mL\ soln}{L\ soln}$ $\frac{mol\ HNO_3}{L\ soln}$

$\rightarrow$ M $H_3O^+ \rightarrow$ pH and then $[H_3O^+] \rightarrow [OH^-]$

$[HNO_3] \rightarrow [H_3O^+]$ $pH = -\log[H_3O^+]$ $[H_3O^+][OH^-] = 1.0 \times 10^{-14}$

Solution: $\frac{0.655\ \cancel{g\ HNO_3}}{100\ \cancel{g\ soln}} \times \frac{1\ mol\ HNO_3}{63.018\ \cancel{g\ HNO_3}} \times \frac{1.01\ \cancel{g\ soln}}{\cancel{mL\ soln}} \times \frac{1000\ \cancel{mL\ soln}}{L\ soln} = 0.105\ M\ HNO_3$

$0.105\ M\ HNO_3 = 0.105\ M\ H_3O^+$ $pH = -\log(0.105) = 0.979$
$[OH^-] = 1.00 \times 10^{-14}/0.105\ M = 9.52 \times 10^{-14}$

Check: HNO_3 is a strong acid and completely dissociates. This gives a relatively high concentration, so we expect the pH to be low and the $[OH^-]$ to be small.

15.56 (a) **Given:** 0.048 M HI (strong acid) **Find:** $[H_3O^+]$,$[OH^-]$, pH

Conceptual Plan: $[HI] \rightarrow [H_3O^+] \rightarrow$ pH and then $[H_3O^+] \rightarrow [OH^-]$

$[HI] \rightarrow [H_3O^+]$ $pH = -\log[H_3O^+]$ $[H_3O^+][OH^-] = 1.0 \times 10^{-14}$

Solution: $0.048\ M\ HI = 0.048\ M\ H_3O^+$ $pH = -\log(0.048) = 1.32$
$[OH^-] = 1.0 \times 10^{-14}/0.048\ M = 2.1 \times 10^{-13}$

Check: HI is a strong acid with a relatively high concentration, so we expect the pH to be low and the $[OH^-]$ to be small.

15.62 **Given:** 0.200 M formic acid. $K_a = 1.8 \times 10^{-4}$ **Find:** $[H_3O^+]$, pH

Conceptual Plan: Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x. Determine $[H_3O^+]$ and pH.

Solution: $HCH_2O(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + CH_2O^-(aq)$

I 0.200 M 0.0 0.0

C $-x$ x x

E $0.200 - x$ x x

$$K_a = \frac{[H_3O^+][CH_2O^-]}{[HCH_2O]} = \frac{(x)(x)}{(0.200 - x)} = 1.8 \times 10^{-4}$$

Assume x is small compared to 0.200.

$x^2 = (1.8 \times 10^{-4})(0.200)$ $x = 6.0 \times 10^{-3}\text{ M} = [H_3O^+]$

Check assumption: $\frac{6.0 \times 10^{-3}}{0.200} \times 100\% = 3.0\%$, assumption valid.

$\text{pH} = -\log(6.0 \times 10^{-3}) = 2.22$

15.63 (a) **Given:** 0.500 M HNO_2. $K_a = 4.6 \times 10^{-4}$ **Find:** pH

Conceptual Plan: Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x. Determine $[H_3O^+]$ and pH.

Solution: $HNO_2(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + NO_2^-(aq)$

I 0.500 M 0.0 0.0

C $-x$ x x

E $0.500 - x$ x x

$$K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]} = \frac{(x)(x)}{(0.500 - x)} = 4.6 \times 10^{-4}$$

Assume x is small compared to 0.500.

$x^2 = (4.6 \times 10^{-4})(0.500)$ $x = 0.015\text{ M} = [H_3O^+]$

Check assumption: $\frac{0.015}{0.500} \times 100\% = 3.0\%$ assumption valid.

$\text{pH} = -\log(0.015) = 1.82$

(b) **Given:** 0.100 M HNO_2. $K_a = 4.6 \times 10^{-4}$ **Find:** pH

Conceptual Plan: Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x. Determine $[H_3O^+]$ and pH.

Solution: $HNO_2(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + NO_2^-(aq)$

I 0.100 M 0.0 0.0

C $-x$ x x

E $0.100 - x$ x x

$$K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]} = \frac{(x)(x)}{(0.100 - x)} = 4.6 \times 10^{-4}$$

Assume x is small compared to 0.100.

$x^2 = (4.6 \times 10^{-4})(0.100)$ $x = 0.0068\text{ M} = [H_3O^+]$

Check assumption: $\frac{0.0068}{0.100} \times 100\% = 6.8\%$ assumption not valid, solve using quadratic equation.

$x^2 = (4.6 \times 10^{-4})(0.100 - x)$ $x^2 + 4.6 \times 10^{-4}x - 4.6 \times 10^{-5} = 0$

$x = 0.00656$

$\text{pH} = -\log(0.00656) = 2.18$

(c) **Given:** 0.100 M HNO_2. $K_a = 4.6 \times 10^{-4}$ **Find:** pH

Conceptual Plan: Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x. Determine $[H_3O^+]$ and pH.

Solution: $HNO_2(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + NO_2^-(aq)$

I	0.0100 M	0.0	0.0
C	$-x$	x	x
E	$0.0100 - x$	x	x

$$K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]} = \frac{(x)(x)}{(0.0100 - x)} = 4.6 \times 10^{-4}$$

Assume x is small compared to 0.100.

$x^2 = (4.6 \times 10^{-4})(0.0100) \qquad x = 0.0021\text{ M} = [H_3O^+]$

Check assumption: $\frac{0.0021}{0.0100} \times 100\% = 21\%$ assumption not valid, solve with quadratic equation.

$x^2 = (4.6 \times 10^{-4})(0.0100 - x) \quad x^2 + 4.6 \times 10^{-4}x - 4.6 \times 10^{-6} = 0$

$x = 0.0019$

$pH = -\log(0.0019) = 2.72$

15.64 (a) **Given:** 0.250 M HF $K_a = 3.5 \times 10^{-4}$ **Find:** pH

Conceptual Plan: Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x. Determine $[H_3O^+]$ and pH.

Solution: $HF(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + F^-(aq)$

I	0.250 M	0.0	0.0
C	$-x$	x	x
E	$0.250 - x$	x	x

$$K_a = \frac{[H_3O^+][F^-]}{[HF]} = \frac{(x)(x)}{(0.250 - x)} = 3.5 \times 10^{-4}$$

Assume x is small compared to 0.250.

$x^2 = (3.5 \times 10^{-4})(0.250) \qquad x = 0.009\underline{3}5\text{ M} = [H_3O^+]$

Check assumption: $\frac{0.009\underline{3}5}{0.250} \times 100\% = 3.7\%$ assumption valid.

$pH = -\log(0.009\underline{3}5) = 2.03$

(b) **Given:** 0.0500 M HF $K_a = 3.5 \times 10^{-4}$ **Find:** pH

Conceptual Plan: Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x. Determine $[H_3O^+]$ and pH.

Solution: $HF(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + F^-(aq)$

I	0.0500 M	0.0	0.0
C	$-x$	x	x
E	$0.0500 - x$	x	x

$$K_a = \frac{[H_3O^+][F^-]}{[HF]} = \frac{(x)(x)}{(0.0500 - x)} = 3.5 \times 10^{-4}$$

Assume x is small compared to 0.0500.

$x^2 = (3.5 \times 10^{-4})(0.0500) \qquad x = 0.004\underline{1}8\text{ M} = [H_3O^+]$

Check assumption: $\frac{0.004\underline{1}8}{0.050} \times 100\% = 8.3\%$ assumption not valid, solve with quadratic equation.

$x^2 + 3.5 \times 10^{-4}x - 1.75 \times 10^{-5} = 0 \qquad x = 0.0040 = [H_3O^+]$

$pH = -\log(0.0040) = 2.40$

(c) **Given:** 0.0250 M HF $K_a = 3.5 \times 10^{-4}$ **Find:** pH

Conceptual Plan: Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x. Determine $[H_3O^+]$ and pH.

Solution: $HF(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + F^-(aq)$

I	0.0250 M	0.0	0.0
C	$-x$	x	x
E	$0.0250 - x$	x	x

$$K_a = \frac{[H_3O^+][F^-]}{[HF]} = \frac{(x)(x)}{(0.0250 - x)} = 3.5 \times 10^{-4}$$

Assume x is small compared to 0.0250.

$x^2 = (3.5 \times 10^{-4})(0.0250)$ $\quad x = 0.002\underline{9}5\ M = [H_3O^+]$

Check assumption: $\frac{0.002\underline{9}5}{0.0250} \times 100\% = 11.8\%$ assumption not valid, solve with quadratic equation.

$x^2 + 3.5 \times 10^{-4}x - 8.75 \times 10^{-6} = 0$ $\quad x = 0.002\underline{7}9 = [H_3O^+]$

$pH = -\log(0.002\underline{7}9) = 2.55$

15.65 **Given:** 15.0 mL glacial acetic, $d = 1.05$ g/mL, dilute to 1.50 L, $K_a = 1.8 \times 10^{-5}$ **Find:** pH

Conceptual Plan: mL acetic acid → g acetic acid → mol acetic acid → M and then write a balanced reaction.

$$\frac{1.05\ g}{mL} \qquad \frac{mol\ acetic\ acid}{60.05\ g} \qquad M = \frac{mol}{L}$$

Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x. Determine $[H_3O^+]$ and pH.

Solution: $15.0\ \cancel{mL} \times \frac{1.05\ \cancel{g}}{\cancel{mL}} \times \frac{1\ mol}{60.05\ \cancel{g}} \times \frac{1}{1.50\ L} = 0.17\underline{4}8\ M$

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

	$HC_2H_3O_2$	H_3O^+	$C_2H_3O_2^-$
I	0.1748 M	0.0	0.0
C	$-x$	x	x
E	$0.1748 - x$	x	x

$$K_a = \frac{[H_3O^+][C_2H_3O_2^-]}{[HC_2H_3O_2]} = \frac{(x)(x)}{(0.1748 - x)} = 1.8 \times 10^{-5}$$

Assume x is small compared to 0.1748.

$x^2 = (1.8 \times 10^{-5})(0.1748)$ $\quad x = 0.001\underline{7}7\ M = [H_3O^+]$

Check assumption: $\frac{0.001\underline{7}7}{0.1748} \times 1005 = 1.0\%$ assumption valid.

$pH = -\log(0.001\underline{7}7) = 2.75$

15.66 **Given:** 1.35% formic acid, $d = 1.01$ g/mL, $K_a = 1.8 \times 10^{-4}$ **Find:** pH

Conceptual Plan: % formic acid → g formic acid → mol and g soln → mL soln → L soln and then M

$$\frac{mol}{46.03\ g} \qquad \frac{1.01\ g\ soln}{mL\ soln} \qquad \frac{1000\ mL}{L\ soln}$$

Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x. Determine $[H_3O^+]$ and pH.

Solution: $\frac{1.35\ \cancel{g\ HCHO_2}}{100\ \cancel{g\ soln}} \times \frac{mol\ HCHO_2}{46.03\ \cancel{g}} \times \frac{1.01\ \cancel{g\ soln}}{\cancel{mL\ soln}} \times \frac{1000\ \cancel{mL\ soln}}{L\ soln} = 0.29\underline{6}2\ M$

$$HCHO_2(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + CHO_2^-(aq)$$

	$HCHO_2$	H_3O^+	CHO_2^-
I	0.2962 M	0.0	0.0
C	$-x$	x	x
E	$0.2962 - x$	x	x

$$K_a = \frac{[H_3O^+][CHO_2^-]}{[HCHO_2]} = \frac{(x)(x)}{(0.2962 - x)} = 1.8 \times 10^{-4}$$

Assume x is small compared to 0.2962.

$x^2 = (1.8 \times 10^{-4})(0.2962)$ $\quad x = 0.007\underline{3}0\ M = [H_3O^+]$

Check assumption: $\frac{0.007\underline{3}0}{0.2962} \times 100\% = 2.5\%$ assumption valid.

$pH = -\log(0.007\underline{3}0) = 2.14$

15.67 **Given:** 0.185 M HA, pH = 2.95 **Find:** K_a

Conceptual Plan: pH → $[H_3O^+]$ and then write a balanced reaction. Prepare an ICE table, calculate equilibrium concentrations, and then plug into the equilibrium expression to solve for K_a.

Solution: $[H_3O^+] = 10^{-2.95} = 0.00112\ M = [A^-]$

	$HA(aq) + H_2O(l) \rightleftharpoons$	$H_3O^+(aq) +$	$A^-(aq)$
I	0.185 M	0.0	0.0
C	$-x$	x	x
E	$0.185 - 0.00112$	0.00112	0.0011

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} = \frac{(0.00112)(0.00112)}{(0.185 - 0.00112)} = 6.8 \times 10^{-6}$$

15.68 **Given:** 0.115 M HA, pH = 3.29 **Find:** K_a

Conceptual Plan: pH → $[H_3O^+]$ and then write a balanced reaction. Prepare an ICE table, calculate equilibrium concentrations, and then plug into the equilibrium expression to solve for K_a.

Solution: $[H_3O^+] = 10^{-3.29} = 5.\underline{1}3 \times 10^{-4}\ M = [A^-]$

	$HA(aq) + H_2O(l) \rightleftharpoons$	$H_3O^+(aq) +$	$A^-(aq)$
I	0.185 M	0.0	0.0
C	$-x$	x	x
E	$0.115 - 5.\underline{1}3 \times 10^{-4}$	$5.\underline{1}3 \times 10^{-4}$	$5.\underline{1}3 \times 10^{-4}$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} = \frac{(5.\underline{1}3 \times 10^{-4})(5.\underline{1}3 \times 10^{-4})}{(0.115 - 5.\underline{1}3 \times 10^{-4})} = 2.3 \times 10^{-6}$$

15.69 **Given:** 0.125 M HCN $K_a = 4.9 \times 10^{-10}$ **Find:** % ionization

Conceptual Plan: Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, solve for x, and then x → % ionization.

$$\%\ \text{ionization} = \frac{x}{[HCN]_{\text{original}}} \times 100$$

Solution: $HCN(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CN^-(aq)$

	$HCN(aq)$	$H_3O^+(aq)$	$CN^-(aq)$
I	0.125 M	0.0	0.0
C	$-x$	x	x
E	$0.125 - x$	x	x

$$K_a = \frac{[H_3O^+][CN^-]}{[HCN]} = \frac{(x)(x)}{(0.125 - x)} = 4.9 \times 10^{-10}$$

Assume x is small compared to 0.125.

$x^2 = (4.9 \times 10^{-10})(0.125) \qquad x = 7.\underline{8}3 \times 10^{-6}$

$$\%\ \text{ionization} = \frac{7.\underline{8}3 \times 10^{-6}}{0.125} \times 100 = 0.0063\%\ \text{ionized}$$

15.70 **Given:** 0.225 M $HC_7H_5O_2$ $K_a = 6.5 \times 10^{-5}$ **Find:** % ionization

Conceptual Plan: Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, solve for x, and then x → % ionization.

$$\%\ \text{ionization} = \frac{x}{[HC_7H_5O_2]_{\text{original}}} \times 100$$

Solution: $HC_7H_5O_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + C_7H_5O_2^-(aq)$

	$HC_7H_5O_2(aq)$	$H_3O^+(aq)$	$C_7H_5O_2^-(aq)$
I	0.125 M	0.0	0.0
C	$-x$	x	x
E	$0.125 - x$	x	x

$$K_a = \frac{[H_3O^+][C_7H_5O_2^-]}{[HC_7H_5O_2]} = \frac{(x)(x)}{(0.225 - x)} = 6.5 \times 10^{-5}$$

Assume x is small compared to 0.225.

$x^2 = (6.5 \times 10^{-5})(0.1225) \qquad x = 0.003\underline{8}2$

$$\%\ \text{ionization} = \frac{0.003\underline{8}2}{0.225} \times 100 = 1.7\%\ \text{ionized}$$

$$K_a = \frac{[H_3O^+][CHO_2^-]}{[HCHO_2]} = \frac{(x)(x)}{(0.100 - x)} = 1.8 \times 10^{-4}$$

Assume x is small compared to 0.100.

$x^2 = (1.8 \times 10^{-4})(0.100) \qquad x = 0.004\underline{2}4$

$$\% \text{ ionization} = \frac{0.004\underline{2}4}{0.100} \times 100 = 4.2\% \text{ ionized}$$

(d) **Given:** 0.0500 M $HCHO_2$ $K_a = 1.8 \times 10^{-4}$ **Find:** % ionization

Conceptual Plan: Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, solve for x, and then $x \rightarrow$ % ionization.

$$\% \text{ ionization} = \frac{x}{[HCHO_2]_{\text{original}}} \times 100$$

Solution: $HCHO_2(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + CHO_2^-(aq)$

I	0.0500 M	0.0	0.0
C	$-x$	x	x
E	$0.0500 - x$	x	x

$$K_a = \frac{[H_3O^+][CHO_2^-]}{[HCHO_2]} = \frac{(x)(x)}{(0.0500 - x)} = 1.8 \times 10^{-4}$$

Assume x is small compared to 0.0500.

$x^2 = (1.8 \times 10^{-4})(0.0500) \qquad x = 0.0030\underline{0}$

$$\% \text{ ionization} = \frac{0.0030\underline{0}}{0.0500} \times 100\% = 6.0\% \text{ ionized}$$

x is small assumption is invalid since 6.0% is greater than the 5.0% limit.

$x^2 + 1.8 \times 10^{-4}x - 9.0 \times 10^{-6} = 0$

Solve for x using the quadratic equation $\qquad x = 0.0029\underline{1}$

$$\% \text{ ionization} = \frac{0.0029\underline{1}}{0.0500} \times 100 = 5.8\%$$

15.73 **Given:** 0.148 M HA 1.55% dissociation **Find:** K_a

Conceptual Plan: M $\rightarrow$ $[H_3O^+]$ $\rightarrow$ K_a and then write a balanced reaction, determine equilibrium concentration, and plug into the equilibrium expression.

Solution: (0.148 M HA)(0.0155) = $0.0022\underline{9}4$ $[H_3O^+] = [A^-]$

$HA(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + A^-(aq)$

I	0. 148 M	0.0	0.0
C	$-x$	x	x
E	0.148 −0.002294	0.002294	0.002294

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} = \frac{(0.0022\underline{9}4)(0.0022\underline{9}4)}{(0.148 - 0.0022\underline{9}4)} = 3.61 \times 10^{-5}$$

15.74 **Given:** 0.085 M HA 0.59% dissociation **Find:** K_a

Conceptual Plan: M $\rightarrow$ $[H_3O^+]$ $\rightarrow$ K_a and then write a balanced reaction, determine equilibrium concentration, and plug into the equilibrium expression.

Solution: (0.085 M HA)(0.0059) = $5.\underline{0}2 \times 10^{-4}[H_3O^+] = [A^-]$

$HA(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + A^-(aq)$

I	0. 085 M	0.0	0.0
C	$-x$	x	x
E	$0.085 - 5.02 \times 10^{-4}$	5.02×10^{-4}	5.02×10^{-4}

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} = \frac{(5.\underline{0}2 \times 10^{-4})(5.\underline{0}2 \times 10^{-4})}{(0.085 - 5.\underline{0}2 \times 10^{-4})} = 3.0 \times 10^{-6}$$

15.75 (a) **Given:** 0.250 M HF $K_a = 3.5 \times 10^{-4}$ **Find:** pH, % dissociation

Conceptual Plan: Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, solve for x, and then $x \rightarrow$ % ionization.

$$\% \text{ ionization} = \frac{x}{[HF]_{\text{original}}} \times 100$$

15.78 (a) **Given:** 0.075 M HNO_3 (strong acid), 0.175 M $HC_7H_5O_2$ (weak acid) **Find:** pH

Conceptual Plan: Since the mixture is a strong acid and a weak acid, the strong acid will dominate. Use the concentration of the strong acid to determine $[H_3O^+]$ and then pH.

Solution: 0.075 M HNO_3 = 0.075 M $[H_3O^+]$ pH = - log(0.075) = 1.12

(b) **Given:** 0.020 M HBr (strong acid), 0.015 M $HClO_4$ (strong acid) **Find:** pH

Conceptual Plan: Since the mixture is a strong acid and another strong acid, the $[H_3O^+]$ is the sum of the concentration of both acids. Determine pH.

Solution: 0.020 M HBr = 0.020 $[H_3O^+]$, 0.015 M $HClO_4$ = 0.015 M $[H_3O^+]$

$[H_3O^+]$ = 0.020 + 0.015 = 0.035 M pH = -log(0.035) = 1.46

(c) **Given:** 0.095M HF, $K_a = 3.5 \times 10^{-4}$; 0.225 M $HC_6H_5O_2$, $K_a = 1.3 \times 10^{-10}$ **Find:** pH

Conceptual Plan: Since the values of K are more than 10^1 apart, the acid with the larger K will dominate the reaction. Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x.

Solution: $HF(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + F^-(aq)$

I	0.095 M	0.0	0.0
C	$-x$	x	x
E	$0.095 - x$	x	x

$$K_a = \frac{[H_3O^+][F^-]}{[HF]} = \frac{(x)(x)}{(0.095 - x)} = 3.5 \times 10^{-4}$$

$x^2 = (3.5 \times 10^{-4})(0.095)$ $\quad x = 0.005\underline{7}7\text{ M} = [H_3O^+]$

$(0.005\underline{7}7/0.095) \times 100 = 6.1\%$. Assumption is invalid since 6.1% is greater than the 5.0% limit.

Solve for x using the quadratic formula;

$x^2 + 3.5 \times 10^{-4}x - 3.325 \times 10^{-5} = 0$

$x = 0.005\underline{5}9\text{ M} = [H_3O^+]$

$\text{pH} = -\log(0.005\underline{5}9) = 2.25$

(d) **Given:** 0.100 M $HCHO_2$, $K_a = 1.8 \times 10^{-4}$; 0.050 M HClO, $K_a = 2.9 \times 10^{-8}$ **Find:** pH

Conceptual Plan: Since the values of K are more than 10^1 apart, the acid with the larger K will dominate the reaction. Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x.

Solution: $HCHO_2(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + C_2H_3O_2^-(aq)$

I	0.100 M	0.0	0.0
C	$-x$	x	x
E	$0.100 - x$	x	x

$$K_a = \frac{[H_3O^+][CHO_2^-]}{[HCHO_2]} = \frac{(x)(x)}{(0.100 - x)} = 1.8 \times 10^{-4}$$

Assume x is small compared to 0.0500.

$x^2 = (1.8 \times 10^{-4})(0.100)$ $\quad x = 0.004\underline{2}4$

$\text{pH} = -\log(0.004\underline{2}4) = 2.37$

Base Solutions

15.79 (a) **Given:** 0.15 M NaOH **Find:** $[OH^-]$, $[H_3O^+]$, pH, pOH

Conceptual Plan: $\mathbf{[NaOH] \rightarrow [OH^-] \xrightarrow[K_w = [H_3O^+][OH^-]]{} [H_3O^+] \xrightarrow[\text{pH} = -\log[H_3O^+]]{} pH \xrightarrow[\text{pH} + \text{pOH} = 14]{} pOH}$

Solution: $[OH^-] = [NaOH] = 0.15M$

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{0.15M} = 6.7 \times 10^{-14}\text{ M}$$

$\text{pH} = -\log(6.7 \times 10^{-14}) = 13.17$

$\text{pOH} = 14.00 - 13.18 = 0.83$

(b) **Given:** 1.5×10^{-3} M $Ca(OH)_2$ **Find:** $[OH^-]$, $[H_3O^+]$, pH, pOH

Conceptual Plan: $[\mathrm{Ca(OH)_2}] \rightarrow [\mathrm{OH^-}] \xrightarrow[K_w = [H_3O^+][OH^-]]{} [\mathrm{H_3O^+}] \xrightarrow[pH = -\log[H_3O^+]]{} \mathrm{pH} \xrightarrow[pH + pOH = 14]{} \mathrm{pOH}$

Solution: $[OH^-] = 2[Ca(OH)_2] = 2(1.5 \times 10^{-3}) = 0.0030$ M

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{0.0030\ M} = 3.\underline{3}3 \times 10^{-12}\ M$$

$$pH = -\log(3.\underline{3}3 \times 10^{-12}) = 11.48$$

$$pOH = 14.00 - 11.48 = 2.52$$

(c) **Given:** 4.8×10^{-4} M $Sr(OH)_2$ **Find:** $[OH^-]$, $[H_3O^+]$, pH, pOH

Conceptual Plan: $[\mathrm{Sr(OH)_2}] \rightarrow [\mathrm{OH^-}] \xrightarrow[K_w = [H_3O^+][OH^-]]{} [\mathrm{H_3O^+}] \xrightarrow[pH = -\log[H_3O^+]]{} \mathrm{pH} \xrightarrow[pH + pOH = 14]{} \mathrm{pOH}$

Solution: $[OH^-] = [Sr(OH)_2] = 2(4.8 \times 10^{-4}) = 9.6 \times 10^{-4}$ M

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{9.6 \times 10^{-4}\ M} = 1.\underline{0}4 \times 10^{-11}\ M$$

$$pH = -\log(1.\underline{0}4 \times 10^{-11}) = 10.98$$

$$pOH = 14.00 - 10.98 = 3.02$$

(d) **Given:** 8.7×10^{-5} M KOH **Find:** $[OH^-]$, $[H_3O^+]$, pH, pOH

Conceptual Plan: $[\mathrm{KOH}] \rightarrow [\mathrm{OH^-}] \xrightarrow[K_w = [H_3O^+][OH^-]]{} [\mathrm{H_3O^+}] \xrightarrow[pH = -\log[H_3O^+]]{} \mathrm{pH} \xrightarrow[pH + pOH = 14]{} \mathrm{pOH}$

Solution: $[OH^-] = [KOH] = 8.7 \times 10^{-5}$ M

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{8.7 \times 10^{-5}\ M} = 1.\underline{1} \times 10^{-10}\ M$$

$$pH = -\log(1.\underline{1} \times 10^{-10}) = 9.94$$

$$pOH = 14.00 - 9.94 = 4.06$$

15.80 (a) **Given:** 8.77×10^{-3} M LiOH **Find:** $[OH^-]$, $[H_3O^+]$, pH, pOH

Conceptual Plan: $[\mathrm{LiOH}] \rightarrow [\mathrm{OH^-}] \xrightarrow[K_w = [H_3O^+][OH^-]]{} [\mathrm{H_3O^+}] \xrightarrow[pH = -\log[H_3O^+]]{} \mathrm{pH} \xrightarrow[pH + pOH = 14]{} \mathrm{pOH}$

Solution: $[OH^-] = [LiOH] = 8.77 \times 10^{-3}$ M

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{8.77 \times 10^{-3}\ M} = 1.1\underline{4}0 \times 10^{-12}\ M$$

$$pH = -\log(1.1\underline{4}0 \times 10^{-12}) = 11.943$$

$$pOH = 14.00 - 11.943 = 2.057$$

(b) **Given:** 0.0112 M $Ba(OH)_2$ **Find:** $[OH^-]$, $[H_3O^+]$, pH, pOH

Conceptual Plan: $[\mathrm{Ba(OH)_2}] \rightarrow [\mathrm{OH^-}] \xrightarrow[K_w = [H_3O^+][OH^-]]{} [\mathrm{H_3O^+}] \xrightarrow[pH = -\log[H_3O^+]]{} \mathrm{pH} \xrightarrow[pH + pOH = 14]{} \mathrm{pOH}$

Solution: $[OH^-] = 2[Ba(OH)_2] = 2(0.0112) = 0.0224$ M

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{0.0224\ M} = 4.4\underline{6}4 \times 10^{-13}\ M$$

$$pH = -\log(4.4\underline{6}4 \times 10^{-13}) = 12.350$$

$$pOH = 14.000 - 12.350 = 1.650$$

(c) **Given:** 1.9×10^{-4} M KOH **Find:** $[OH^-]$, $[H_3O^+]$, pH, pOH

Conceptual Plan: $[\mathrm{KOH}] \rightarrow [\mathrm{OH^-}] \xrightarrow[K_w = [H_3O^+][OH^-]]{} [\mathrm{H_3O^+}] \xrightarrow[pH = -\log[H_3O^+]]{} \mathrm{pH} \xrightarrow[pH + pOH = 14]{} \mathrm{pOH}$

Solution: $[OH^-] = [KOH] = 1.9 \times 10^{-4}$ M

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{1.9 \times 10^{-4}\ M} = 5.\underline{2}6 \times 10^{-11}\ M$$

$$pH = -\log(5.\underline{2}6 \times 10^{-11}) = 10.28$$

$$pOH = 14.00 - 10.28 = 3.72$$

(d) **Given:** 5.0×10^{-4} M $Ca(OH)_2$ **Find:** $[OH^-]$, $[H_3O^+]$, pH, pOH
Conceptual Plan: $\mathbf{[Ca(OH)_2] \rightarrow [OH^-] \rightarrow [H_3O^+] \rightarrow pH \rightarrow pOH}$
$K_w = [H_3O^+][OH^-]$ $pH = -\log[H_3O^+]$ $pH + pOH = 14$
Solution: $[OH^-] = [Ca(OH)_2] = 2(5.0 \times 10^{-4}) = 0.0010$ M

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{0.0010\ M} = 1.\underline{0}0 \times 10^{-11}\ M$$
$$pH = -\log(1.\underline{0}0 \times 10^{-11}) = 11.00$$
$$pOH = 14.00 - 11.00 = 3.00$$

15.81 **Given:** 3.85% KOH by mass, $d = 1.01$ g/mL **Find:** pH
Conceptual Plan:
% mass → g KOH → mol KOH and mass soln → mL soln → L soln → M KOH → $[OH^-]$
$\frac{1\ \text{mol KOH}}{56.01\ \text{g KOH}}$ $\frac{1.01\ \text{g soln}}{\text{mL soln}}$ $\frac{1000\ \text{mL soln}}{\text{L soln}}$ $\frac{\text{mol KOH}}{\text{L soln}}$
→ pOH → pH
$pOH = -\log[OH^-]$ $pH + pOH = 14$

Solution: $$\frac{3.85\ \cancel{\text{g KOH}}}{100.0\ \cancel{\text{g soln}}} \times \frac{1\ \text{mol KOH}}{56.01\ \cancel{\text{g KOH}}} \times \frac{1.01\ \cancel{\text{g soln}}}{\cancel{\text{mL soln}}} \times \frac{1000\ \cancel{\text{mL soln}}}{\text{L soln}} = 0.69\underline{4}2\ \text{M KOH}$$

$[OH^-] = [KOH] = 0.69\underline{4}2$ M $pOH = -\log(0.69\underline{4}2) = 0.159$ $pH = 14.000 - 0.159 = 13.841$

15.82 **Given:** 1.55% NaOH by mass, $d = 1.01$ g/mL **Find:** pH
Conceptual Plan:
% mass → g NaOH → mol NaOH and mass soln → mL soln → L soln → M NaOH → $[OH^-]$
$\frac{1\ \text{mol NaOH}}{40.01\ \text{g NaOH}}$ $\frac{1.01\ \text{g soln}}{\text{mL soln}}$ $\frac{1000\ \text{mL soln}}{\text{L soln}}$ $\frac{\text{mol NaOH}}{\text{L soln}}$
→ pOH → pH
$pOH = -\log[OH^-]$ $pH + pOH = 14$

Solution: $$\frac{1.55\ \cancel{\text{g NaOH}}}{100.0\ \cancel{\text{g soln}}} \times \frac{1\ \text{mol NaOH}}{40.01\ \cancel{\text{g KOH}}} \times \frac{1.01\ \cancel{\text{g soln}}}{\cancel{\text{mL soln}}} \times \frac{1000\ \cancel{\text{mL soln}}}{\text{L soln}} = 0.39\underline{1}3\ \text{M NaOH}$$

$[OH^-] = [NaOH] = 0.39\underline{1}3$ M $pOH = -\log(0.39\underline{1}3) = 0.407$ $pH = 14.000 - 0.407 = 13.593$

15.83 **Given:** 3.55 L, pH = 12.4; 0.855 M KOH **Find:** Vol
Conceptual Plan: **pH → $[H_3O^+]$ → $[OH^-]$ and then $V_1M_1 = V_2M_2$**
$[H_3O^+] = 10^{-pH}$ $1.0 \times 10^{-14} = [H_3O^+][OH^-]$ $V_1M_1 = V_2M_2$
Solution: $[H_3O^+] = 10^{-12.4} = 3.98 \times 10^{-13}$ $1.0 \times 10^{-14} = 3.98 \times 10^{-13}[OH^-]$
$[OH^-] = 0.025\underline{1}2$
$V_1M_1 = V_2M_2$ $V_1(0.855\ M) = (3.55\ L)(0.0251\ M)$ $V_1 = 0.104$ L

15.84 **Given:** 5.00 L, pH = 10.8; 15.0% NaOH, $d = 1.116$ g/ mL **Find:** Vol
Conceptual Plan:
% mass → g NaOH → mol NaOH and mass soln → mL soln → L soln → M NaOH → $[OH^-]$
$\frac{1\ \text{mol NaOH}}{40.01\ \text{g NaOH}}$ $\frac{1.01\ \text{g soln}}{\text{mL soln}}$ $\frac{1000\ \text{mL soln}}{\text{Lsoln}}$ $\frac{\text{mol NaOH}}{\text{L soln}}$
and then: $V_1M_1 \rightarrow V_2M_2$
$V_1M_1 = V_2M_2$

Solution: $$\frac{15.0\ \cancel{\text{g NaOH}}}{100.0\ \cancel{\text{g soln}}} \times \frac{1\ \text{mol NaOH}}{40.01\ \cancel{\text{g KOH}}} \times \frac{1.116\ \cancel{\text{g soln}}}{\cancel{\text{mL soln}}} \times \frac{1000\ \cancel{\text{mL soln}}}{\text{L soln}} = 4.1\underline{8}4\ \text{M NaOH}$$

$[OH^-] = [NaOH] = 4.1\underline{8}4$ M
$[H_3O^+] = 10^{-10.8} = 1.58 \times 10^{-11}$ $1.0 \times 10^{-14} = (1.58 \times 10^{-11})[OH^-]$
$[OH^-] = 6.31 \times 10^{-4}$ M
$V_1(4.18\ M) = (5.00\ L)(6.31 \times 10^{-4}\ M)$ $V_1 = 7.55 \times 10^{-4}\ L = 0.8$ mL

15.85 (a) $NH_3(aq) + H_2O(l) \leftrightarrows NH_4^+(aq) + OH^-(aq)$ $K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$

(b) $HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$ $\quad K_b = \dfrac{[H_2CO_3][OH^-]}{[HCO_3^-]}$

(c) $CH_3NH_2(aq) + H_2O(l) \rightleftharpoons CH_3NH_3^+(aq) + OH^-(aq)$ $\quad K_b = \dfrac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]}$

15.86 (a) $CO_3^{2-}(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + OH^-(aq)$ $\quad K_b = \dfrac{[HCO_3^-][OH^-]}{[CO_3^{2-}]}$

(b) $C_6H_5NH_2(aq) + H_2O(l) \rightleftharpoons C_6H_5NH_3^+(aq) + OH^-(aq)$ $\quad K_b = \dfrac{[C_6H_5NH_3^+][OH^-]}{[C_6H_5NH_2]}$

(c) $C_2H_5NH_2(aq) + H_2O(l) \rightleftharpoons C_2H_5NH_3^+(aq) + OH^-(aq)$ $\quad K_b = \dfrac{[C_2H_5NH_3^+][OH^-]}{[C_2H_5NH_2]}$

15.87 **Given:** 0.15 M NH_3 $K_b = 1.76 \times 10^{-5}$ **Find:** $[OH^-]$, pH, pOH

Conceptual Plan: Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x.

$x = [OH^-] \rightarrow pOH \rightarrow pH$

pOH = -log[OH⁻] pH + pOH = 14

Solution: $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

I	0.15 M	0.0	0.0
C	$-x$	x	x
E	$0.15 - x$	x	x

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = \frac{(x)(x)}{(0.15 - x)} = 1.76 \times 10^{-5}$$

Assume x is small.

$x^2 = (1.76 \times 10^{-5})(0.15)$ $\quad x = [OH^-] = 0.001\underline{6}2\ M$

$pOH = -\log(0.001\underline{6}2) = 2.79$

$pH = 14.00 - 2.79 = 11.21$

15.88 **Given:** 0.125 M CO_3^{2-} $K_b = 1.8 \times 10^{-4}$ **Find:** $[OH^-]$, pH, pOH

Conceptual Plan: Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x.

$x = [OH^-] \rightarrow pOH \rightarrow pH$

pOH = -log[OH⁻] pH + pOH = 14

Solution: $CO_3^{2-}(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + OH^-(aq)$

I	0.125 M	0.0	0.0
C	$-x$	x	x
E	$0.125 - x$	x	x

$$K_b = \frac{[HCO_3^-][OH^-]}{[CO_3^-]} = \frac{(x)(x)}{(0.125 - x)} = 1.8 \times 10^{-4}$$

Assume x is small.

$x^2 = (1.8 \times 10^{-4})(0.125)$ $\quad x = [OH^-] = 0.004\underline{7}4\ M$

$$\frac{0.004\underline{7}4}{0.125} \times 100\% = 3.8\%\text{; assumption is valid}$$

$pOH = -\log(0.004\underline{7}4) = 2.32$

$pH = 14.00 - 2.32 = 11.68$

15.89 **Given:** $pK_b = 10.4$, 455 mg/L caffeine **Find:** pH

Conceptual Plan: $pK_b \rightarrow K_b$ and then mg/L → g/L → mol/L and then write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x.

$x = [OH^-] \rightarrow pOH \rightarrow pH$

pOH = -log[OH⁻] pH + pOH = 14

Solution: $K_b = 10^{-10.4} = \underline{3}.98 \times 10^{-11}$

$$\frac{455\ \cancel{\text{mg caffeine}}}{\text{L soln}} \times \frac{\cancel{\text{g caffeine}}}{1000\ \cancel{\text{mg caffeine}}} \times \frac{1\ \text{mol caffeine}}{194.19\ \cancel{\text{g}}} = 0.0023\underline{4}3\ \text{M caffeine}$$

$$\begin{array}{llll}
 & C_8H_{10}N_4O_2(aq) + H_2O(l) \leftrightarrows & HC_8H_{10}N_4O_2^+(aq) + & OH^-(aq) \\
\text{I} & 0.0023\underline{4}2 & 0.0 & 0.0 \\
\text{C} & -x & x & x \\
\text{E} & 0.0023\underline{4}2 - x & x & x
\end{array}$$

$$K_b = \frac{[HC_8H_{10}N_4O_2^+][OH^-]}{[C_8H_{10}N_4O_2]} = \frac{(x)(x)}{(0.0023\underline{4}2 - x)} = \underline{3}.98 \times 10^{-11}$$

Assume x is small.

$x^2 = (\underline{3}.98 \times 10^{-11})(0.0023\underline{4}2) \quad x = [OH^-] = \underline{3}.05 \times 10^{-7}\ \text{M}$

$\dfrac{\underline{3}.05 \times 10^{-7}\ \text{M}}{0.0023\underline{4}2} \times 100\% = 0.013\%$; assumption is valid

$\text{pOH} = -\log(\underline{3}.05 \times 10^{-7}) = 6.5$

$\text{pH} = 14.00 - 6.5 = 7.5$

15.90 **Given:** pK_b = 4.2, 225 mg/L amphetamine **Find:** pH

Conceptual Plan: pK_b → K_b and then mg/L → g/L → mol/L and then write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x.

x = [OH⁻] → pOH → pH

$\text{pOH} = -\log[OH^-]\ \ \text{pH} + \text{pOH} = 14$

Solution: $K_b = 10^{-4.2} = \underline{6}.31 \times 10^{-5}$

$$\frac{225\ \cancel{\text{mg amphetamine}}}{\text{L soln}} \times \frac{\cancel{\text{g amphetamine}}}{1000\ \cancel{\text{mg amphetamine}}} \times \frac{1\ \text{mol amphetamine}}{135.21\ \cancel{\text{g}}} = 0.0016\underline{6}4\ \text{M amphetamine}$$

$$\begin{array}{llll}
 & C_9H_{13}N(aq) + H_2O(l) \leftrightarrows & C_9H_{13}NH^+(aq) + & OH^-(aq) \\
\text{I} & 0.0016\underline{6}4\ \text{M} & 0.0 & 0.0 \\
\text{C} & -x & x & x \\
\text{E} & 0.0016\underline{6}4 - x & x & x
\end{array}$$

$$K_b = \frac{[C_9H_{13}NH^+][OH^-]}{[C_9H_{13}N]} = \frac{(x)(x)}{(0.0016\underline{6}4 - x)} = \underline{6}.31 \times 10^{-5}$$

Assume x is small.

$x^2 = (\underline{6}.31 \times 10^{-5})(0.0016\underline{6}4) \quad x = [OH^-] = \underline{3}.24 \times 10^{-4}\ \text{M}$

$\dfrac{\underline{3}.24 \times 10^{-4}}{0.0016\underline{6}4} \times 100\% = 19.4\%$; assumption not valid, solve with quadratic equation.

$x^2 + \underline{6}.31 \times 10^{-5} x - 1.05 \times 10^{-7} = 0$

$x = [OH^-] = \underline{2}.94 \times 10^{-4}\ \text{M}$

$\text{pOH} = -\log(\underline{2}.94 \times 10^{-4}) = 3.5$

$\text{pH} = 14.00 - 3.5 = 10.5$

15.91 **Given:** 0.150 M morphine, pH = 10.5 **Find:** K_b

Conceptual Plan:

pH → pOH → [OH⁻] and then write a balanced equation, prepare an ICE table, and determine

$\text{pH} = \text{pOH} = 14\ \ \text{pOH} = -\log[OH^-]$

equilibrium concentrations ⟶ K_b.

Solution: pOH = 14.0 − 10.5 = 3.5 $[OH^-] = 10^{-3.5} = \underline{3}.16 \times 10^{-4} = [\text{Hmorphine}^+]$

$$\begin{array}{llll}
 & \text{morphine}(aq) + H_2O(l) \leftrightarrows & \text{Hmorphine}^+(aq) + & OH^-(aq) \\
\text{I} & 0.150\ \text{M} & 0.0 & 0.0 \\
\text{C} & -x & x & x \\
\text{E} & 0.150 - x & 3.16 \times 10^{-4} & 3.16 \times 10^{-4}
\end{array}$$

$$K_b = \frac{[\text{Hmorphine}^+][OH^-]}{[\text{morphine}]} = \frac{(3.16 \times 10^{-4})(3.16 \times 10^{-4})}{(0.150 - 3.16 \times 10^{-4})} = \underline{6}.68 \times 10^{-7} = 7 \times 10^{-7}$$

15.92 **Given:** 0.135 M base, pH = 11.23 **Find:** K_b

Conceptual Plan:

pH → pOH → [OH⁻] and then write a balanced equation, prepare an ICE table, and determine

pH = pOH = 14 pOH = -log[OH⁻]

equilibrium concentrations ⟶ K_b.

Solution: pOH = 14.00 – 11.23 = 2.77 $[OH^-] = 10^{-2.77} = 1.\underline{6}98 \times 10^{-3} = [HB^+]$

	$B(aq) + H_2O(l) \leftrightarrows$	$HB^+(aq) +$	$OH^-(aq)$
I	0.135	0.0	0.0
C	$-x$	x	x
E	$0.135 - x$	$1.\underline{6}98 \times 10^{-3}$	$1.\underline{6}98 \times 10^{-3}$

$$K_b = \frac{[HB^+][OH^-]}{[B]} = \frac{(1.\underline{6}98 \times 10^{-3})(1.\underline{6}98 \times 10^{-3})}{(0.135 - 1.\underline{6}98 \times 10^{-3})} = 2.\underline{1}6 \times 10^{-5} = 2.2 \times 10^{-5}$$

Acid – Base Properties of Ions and Salts

15.93 (a) pH neutral: Br^- is the conjugate base of a strong acid; therefore, it is pH neutral.

(b) weak base: ClO^- is the conjugate base of a weak acid; therefore, it is a weak base.
$ClO^-(aq) + H_2O(l) \leftrightarrows HClO(aq) + OH^-(aq)$

(c) weak base: CN^- is the conjugate base of a weak acid; therefore, it is a weak base.
$CN^-(aq) + H_2O(l) \leftrightarrows HCN(aq) + OH^-(aq)$

(d) pH neutral: Cl^- is the conjugate base of a strong acid; therefore, it is pH neutral.

15.94 (a) weak base: $C_7H_5O_2^-$ is the conjugate base of a weak acid; therefore, it is a weak base.
$C_7H_5O_2^-(aq) + H_2O(l) \leftrightarrows HC_7H_5O_2(aq) + OH^-(aq)$

(b) pH neutral: I^- is the conjugate base of a strong acid; therefore, it is pH neutral.

(c) pH neutral: NO_3^- is the conjugate base of a strong acid; therefore, it is pH neutral.

(d) weak base: F^- is the conjugate base of a weak acid; therefore, it is a weak base.
$F^-(aq) + H_2O(l) \leftrightarrows HF(aq) + OH^-(aq)$

15.95 **Given:** $[F^-] = 0.140$ M, $K_a(HF) = 3.5 \times 10^{-4}$ **Find:** $[OH^-]$, pOH

Conceptual Plan: Determine K_b. Write a balanced reaction. Prepare an ICE table, represent the change

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

with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x. Determine [OH⁻] → pOH → pH.

pOH = -log[OH⁻] pH + pOH = 14

Solution: $F^-(aq) + H_2O(l) \leftrightarrows HF(aq) + OH^-(aq)$

	F^-	HF	OH^-
I	0.140 M	0.0	0.0
C	$-x$	x	x
E	$0.140 - x$	x	x

$$K_b = \frac{K_w}{K_a} = \frac{1 \times 10^{-14}}{3.5 \times 10^{-4}} = \frac{(x)(x)}{(0.140 - x)}$$

Assume x is small.

$x = 2.0 \times 10^{-6} = [OH^-]$ pOH $= -\log(2.0 \times 10^{-6}) = 5.70$

pH = 14.00 − 5.70 = 8.30

15.96 **Given:** $[HCO_3^-] = 0.250$ M, $K_a(H_2CO_3) = 4.3 \times 10^{-7}$ **Find:** $[OH^-]$, pOH

Conceptual Plan: Determine K_b. Write a balanced reaction. Prepare an ICE table, represent the change

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

with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x. Determine [OH⁻] → pOH → pH.

pOH = -log[OH⁻] pH + pOH = 14

Solution: $HCO_3^-\,(aq) + H_2O(l) \leftrightarrows H_2CO_3\,(aq) + OH^-\,(aq)$

I	0.250	0.0	0.0
C	$-x$	x	x
E	$0.250 - x$	x	x

$$K_b = \frac{K_w}{K_a} = \frac{1 \times 10^{-14}}{4.3 \times 10^{-7}} = \frac{(x)(x)}{(0.250 - x)}$$

Assume x is small.

$x = 7.6 \times 10^{-5} = [OH^-]$ $pOH = -\log(7.6 \times 10^{-5}) = 4.12$

$pH = 14.00 - 4.12 = 9.88$

15.97 (a) weak acid: NH_4^+ is the conjugate acid of a weak base; therefore, it is a weak acid.
$NH_4^+(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + NH_3(aq)$

(b) pH neutral: Na^+ is the counterion of a strong base; therefore, it is pH neutral.

(c) weak acid: The Co^{3+} cation is a small, highly charged metal cation; therefore, it is a weak acid.

$Co(H_2O)_6^{3+}(aq) + H_2O(l) \leftrightarrows Co(H_2O)_5(OH)^{2+}(aq) + H_3O^+(aq)$

(d) weak acid: $CH_2NH_3^+$ is the conjugate acid of a weak base; therefore, it is a weak acid.
$CH_2NH_3^+(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + CH_2NH_2(aq)$

15.98 (a) pH neutral: Sr^{2+} is the counterion of a strong base; therefore, it is pH neutral.

(b) weak acid: The Mn^{3+} cation is a small, highly charged metal cation; therefore, it is a weak acid.
$Mn(H_2O)_6^{3+}(aq) + H_2O(l) \leftrightarrows Mn(H_2O)_5(OH)^{2+}(aq) + H_3O^+(aq)$

(c) weak acid: $C_5H_5NH_3^+$ is the conjugate acid of a weak base; therefore, it is a weak acid.
$C_5H_5NH_3^+\,(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + C_5H_5NH_2(aq)$

(d) pH neutral: Li^+ is the counterion of a strong base; therefore, it is pH neutral.

15.99 (a) acidic: $FeCl_3$ Fe^{3+} is a small, highly charged metal cation and is therefore acidic. Cl^- is the conjugate base of a strong acid; therefore, it is pH neutral.

(b) basic: NaF Na^+ is the counterion of a strong base; therefore, it is pH neutral. F^- is the conjugate base of a weak acid; therefore, it is basic.

(c) pH neutral: $CaBr_2$ Ca^{2+} is the counterion of a strong base; therefore, it is pH neutral. Br^- is the conjugate base of a strong acid; therefore, it is pH neutral.

(d) acidic: NH_4Br NH_4^+ is the conjugate acid of a weak base; therefore, it is acidic. Br^- is the conjugate base of a strong acid; therefore, it is pH neutral.

(e) acidic: $C_6H_5NH_3NO_2$ $C_6H_5NH_3^+$ is the conjugate acid of a weak base; therefore, it is a weak acid. NO_2^- is the conjugate base of a weak acid; therefore, it is a weak base. To determine pH, compare K values.

$$K_a\,(C_6H_5NH_3^+) = \frac{1.0 \times 10^{-14}}{3.9 \times 10^{-10}} = 2.6 \times 10^{-5} \quad K_b(NO_2^-) = \frac{1.0 \times 10^{-14}}{4.6 \times 10^{-4}} = 2.2 \times 10^{-11}$$

$K_a > K_b$; therefore, the solution is acidic.

15.100 (a) acidic: $Al(NO_3)_3$ Al^{3+} is a small, highly charged metal cation and is therefore acidic. NO_3^- is the conjugate base of a strong acid; therefore, it is pH neutral.

(b) acidic: $C_2H_5NH_3NO_3$ $C_2H_5NH_3^+$ is the conjugate acid of a weak base; therefore, it is acidic. NO_3^- is the conjugate base of a strong acid; therefore, it is pH neutral.

(c) basic: K_2CO_3 K^+ is the counterion of a strong base; therefore, it is pH neutral. CO_3^{2-} is the conjugate base of a weak acid; therefore, it is basic.

(d) pH neutral: RbI Rb^+ is the counterion of a strong base; therefore, it is pH neutral. I^- is the conjugate base of a strong acid; therefore, it is pH neutral.

(e) basic NH_4ClO NH_4^+ is the conjugate acid of a weak base; therefore, it is a weak acid. ClO^- is the conjugate base of a weak acid; therefore, it is a weak base. To determine pH, compare K values.

$$K_a(NH_4^+) = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10} \qquad K_b(ClO^-) = \frac{1.0 \times 10^{-14}}{2.9 \times 10^{-8}} = 3.4 \times 10^{-7}$$

$K_b > K_a$; therefore, the solution is basic.

15.101 Identify each species and determine the acid, base, or neutral.

NaCl pH neutral: Na^+ is the counterion of a strong base; therefore, it is pH neutral. Cl^- is the conjugate base of a strong acid; therefore, it is pH neutral.

NH_4Cl acidic: NH_4^+ is the conjugate acid of a weak base; therefore, it is acidic. Cl^- is the conjugate base of a strong acid; therefore, it is pH neutral.

$NaHCO_3$ basic: Na^+ is the counterion of a strong base; therefore, it is pH neutral. HCO_3^- is the conjugate base of a weak acid; therefore, it is basic.

NH_4ClO_2 acidic: NH_4ClO_2 NH_4^+ is the conjugate acid of a weak base; therefore, it is a weak acid. ClO_2^- is the conjugate base of a weak acid; therefore, it is a weak base. $K_a(NH_4^+) = 5.6 \times 10^{-10}$ $K_b(ClO_2^-) = 9.1 \times 10^{-13}$

NaOH strong base

Increasing acidity: $NaOH < NaHCO_3 < NaCl < NH_4ClO_2 < NH_4Cl$

15.102 Identify each species and determine the acid, base, or neutral.

CH_3NH_3Br acidic: $CH_3NH_3^+$ is the conjugate acid of a weak base; therefore, it is acidic. Br^- is the conjugate base of a strong acid; therefore, it is pH neutral.

KOH strong base

KBr pH neutral: K^+ is the counterion of a strong base; therefore, it is pH neutral. Br^- is the conjugate base of a strong acid; therefore, it is pH neutral.

KCN basic: K^+ is the counterion of a strong base; therefore, it is pH neutral. CN^- is the conjugate base of a weak acid; therefore, it is basic.

$C_5H_5NHNO_2$ acidic: $C_5H_5NH^+$ is the conjugate acid of a weak base; therefore, it is acidic. NO_2^- is the conjugate base of a weak acid; therefore, it is basic. $K_a(C_5H_5NH^+) = 5.9 \times 10^{-6}$; $K_b(NO_2^-) = 2.2 \times 10^{-11}$

Increasing basicity: $CH_3NH_3Br < C_5H_5NHNO_2 < KBr < KCN < KOH$

15.103 (a) **Given:** 0.10 M NH_4Cl **Find:** pH

Conceptual Plan: Identify each species and determine which will contribute to pH. Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x. Determine $[H_3O^+] \rightarrow$ pH.

Solution: NH_4^+ is the conjugate acid of a weak base; therefore, it is acidic. Cl^- is the conjugate base of a strong acid; therefore, it is pH neutral.

	$NH_4^+(aq) + H_2O(l) \leftrightarrows$	$NH_3(aq) +$	$H_3O^+(aq)$
I	0.10	0.0	0.0
C	$-x$	x	x
E	$0.10 - x$	x	x

$$K_a = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.5\underline{6} \times 10^{-10} = \frac{(x)(x)}{(0.10 - x)}$$

Assume x is small.

$x = 7.\underline{4}5 \times 10^{-6} = [H_3O^+]$ pH $= -\log(7.\underline{4}5 \times 10^{-6}) = 5.13$

(b) **Given:** 0.10 M $NaC_2H_3O_2$ **Find:** pH

Conceptual Plan: Identify each species and determine which will contribute to pH. Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x. Determine $[OH^-]$ → pOH → pH.

$pOH = -\log[OH^-]$ $pH + pOH = 14$

Solution: Na^+ is the counterion of a strong base; therefore, it is pH neutral. $C_2H_3O_2^-$ is the conjugate base of a weak acid; therefore, it is basic.

$$C_2H_3O_2^-(aq) + H_2O(l) \leftrightarrows H\,C_2H_3O_2\,(aq) + OH^-\,(aq)$$

I	0.10	0.0	0.0
C	$-x$	x	x
E	$0.10 - x$	x	x

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.\underline{5}6 \times 10^{-10} = \frac{(x)(x)}{(0.10 - x)}$$

Assume x is small.

$x = 7.\underline{4}5 \times 10^{-6} = [OH^-]$ $pOH = -\log(7.\underline{4}5 \times 10^{-6}) = 5.13$

$pH = 14.00 - 5.13 = 8.87$

(c) **Given:** 0.10 M NaCl **Find:** pH

Conceptual Plan: Identify each species and determine which will contribute to pH.

Solution: Na^+ is the counterion of a strong base; therefore, it is pH neutral. Cl^- is the conjugate base of a strong acid; therefore, it is pH neutral.

pH = 7.0

15.104 (a) **Given:** 0.20 M $NaCHO_2$ **Find:** pH

Conceptual Plan: Identify each species and determine which will contribute to pH. Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x. Determine $[OH^-]$ → pOH → pH.

$pOH = -\log[OH^-]$ $pH + pOH = 14$

Solution: Na^+ is the counterion of a strong base; therefore, is pH neutral. CHO_2^- is the conjugate base of a weak acid; therefore, it is basic.

$$CHO_2^-(aq) + H_2O(l) \leftrightarrows HCHO_2(aq) + OH^-\,(aq)$$

I	0.20 M	0.0	0.0
C	$-x$	x	x
E	$0.20 - x$	x	x

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-4}} = 5.\underline{5}6 \times 10^{-11} = \frac{(x)(x)}{(0.20 - x)}$$

Assume x is small.

$x = 3.\underline{3}3 \times 10^{-6} = [OH^-]$ $pOH = -\log(3.\underline{3}3 \times 10^{-6}) = 5.48$

$pH = 14.00 - 5.48 = 8.52$

(b) **Given:** 0.20 M CH_3NH_3I **Find:** pH

Conceptual Plan: Identify each species and determine which will contribute to pH. Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x. Determine $[H_3O^+]$ → pH.

Solution: $CH_3NH_4^+$ is the conjugate acid of a weak base; therefore, it is acidic. Cl^- is the conjugate base of a strong acid; therefore, it is pH neutral.

$$CH_3NH_3^+(aq) + H_2O(l) \leftrightarrows CH_3NH_2\,(aq) + H_3O^+\,(aq)$$

I	0.20 M	0.0	0.0
C	$-x$	x	x
E	$0.20 - x$	x	x

$$K_a = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{4.4 \times 10^{-4}} = 2.\underline{2}7 \times 10^{-11} = \frac{(x)(x)}{(0.20 - x)}$$

Polyprotic Acids

15.107 $H_3PO_4(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + H_2PO_4^-(aq)$ $\quad K_{a_1} = \dfrac{[H_3O^+][H_2PO_4^-]}{[H_3PO_4]}$

$H_2PO_4^-(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + HPO_4^{2-}(aq)$ $\quad K_{a_2} = \dfrac{[H_3O^+][HPO_4^{2-}]}{[H_2PO_4^-]}$

$HPO_4^{2-}(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + PO_4^{3-}(aq)$ $\quad K_{a_3} = \dfrac{[H_3O^+][PO_4^{3-}]}{[HPO_4^{2-}]}$

15.108 $H_2CO_3(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + HCO_3^-(aq)$ $\quad K_{a_1} = \dfrac{[H_3O^+][HCO_3^-]}{[H_2CO_3]}$

$HCO_3^-(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + CO_3^{2-}(aq)$ $\quad K_{a_2} = \dfrac{[H_3O^+][CO_3^{2-}]}{[HCO_3^-]}$

15.109 (a) **Given:** 0.350 M H_3PO_4 $\quad K_{a_1} = 7.5 \times 10^{-3}$, $K_{a_2} = 6.2 \times 10^{-8}$ **Find:** $[H_3O^+]$, pH

Conceptual Plan: K_{a_1} is much larger than K_{a_2}, so use K_{a_1} to calculate $[H_3O^+]$. Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x.

Solution: $H_3PO_4(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + H_2PO_4^-(aq)$

I	0.350 M	0.0	0.0
C	$-x$	x	x
E	$0.350 - x$	x	x

$$K_a = \frac{[H_3O^+][H_2PO_4^-]}{[H_3PO_4]} = \frac{(x)(x)}{(0.350 - x)} = 7.3 \times 10^{-3}$$

Assume x is small compared to 0.350.

$x^2 = (7.3 \times 10^{-3})(0.350)$ $\quad x = 0.0505\ \text{M} = [H_3O^+]$

Check assumption: $\dfrac{0.0505}{0.350} \times 100\% = 14.4\%$ assumption not valid, solve with quadratic equation.

$x^2 + 7.5 \times 10^{-3}x - 0.002625 = 0$ $\quad x = 0.04\underline{7}62 = [H_3O^+]$

$\text{pH} = -\log(0.04\underline{7}62) = 1.32$

(b) **Given:** 0.350 M $H_2C_2O_4$ $\quad K_{a_1} = 6.0 \times 10^{-2}$, $K_{a_2} = 6.0 \times 10^{-5}$ **Find:** $[H_3O^+]$, pH

Conceptual Plan: K_{a_1} is much larger than K_{a_2}, so use K_{a_1} to calculate $[H_3O^+]$. Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x.

Solution: $H_2C_2O_4(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + HC_2O_4^-(aq)$

I	0.350 M	0.0	0.0
C	$-x$	x	x
E	$0.350 - x$	x	x

$$K_a = \frac{[H_3O^+][HC_2O_4^-]}{[H_2C_2O_4]} = \frac{(x)(x)}{(0.350 - x)} = 6.0 \times 10^{-2}$$

$x^2 + 6.0 \times 10^{-2}x - 0.021 = 0$ $\quad x = 0.1\underline{1}79 = 0.12\ \text{M}\ [H_3O^+]$

$\text{pH} = -\log(0.1\underline{1}79) = 0.93$

15.110 (a) **Given:** 0.125 M H_2CO_3 $\quad K_{a_1} = 4.3 \times 10^{-7}$, $K_{a_2} = 5.6 \times 10^{-11}$ **Find:** $[H_3O^+]$, pH

Conceptual Plan: K_{a_1} much larger than K_{a_2}, so use K_{a_1} to calculate $[H_3O^+]$. Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x.

Solution: $H_2CO_3(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + HCO_3^-(aq)$

I	0.125 M	0.0	0.0
C	$-x$	x	x
E	$0.125 - x$	x	x

$$K_a = \frac{[H_3O^+][HCO_3^-]}{[H_2CO_3]} = \frac{(x)(x)}{(0.125 - x)} = 4.3 \times 10^{-7}$$

Assume x is small.

$x^2 = (4.3 \times 10^{-7})(0.125) \quad x = 2.\underline{3}2 \times 10^{-4} = [H_3O^+]$

$\frac{2.\underline{3}2 \times 10^{-4}}{0.125} \times 100\% = 0.19\%$; assumption valid

$pH = -\log(2.\underline{3}2 \times 10^{-4}) = 3.63$

(b) **Given:** 0.125 M $H_3C_6H_5O_3$ $K_{a_1} = 7.4 \times 10^{-4}$, $K_{a_2} = 1.7 \times 10^{-5}$, $K_{a_3} = 4.0 \times 10^{-7}$ **Find:** $[H_3O^+]$, pH

Conceptual Plan: K_{a_1} and K_{a_2} are only 10^{-1} apart, so use both to calculate $[H_3O^+]$. Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x.

Solution: $H_3C_6H_5O_3(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + H_2C_6H_5O_3^-(aq)$

I	0.125 M	0.0	0.0
C	$-x$	x	x
E	$0.125 - x$	x	x

$$K_{a_1} = \frac{[H_3O^+][H_2C_6H_5O_3^-]}{[H_3C_6H_5O_3]} = \frac{(x)(x)}{(0.125 - x)} = 7.4 \times 10^{-4}$$

Assume x is small.

$x^2 = (7.4 \times 10^{-4})(0.125) \quad x = 9.\underline{6}2 \times 10^{-3} = [H_3O^+]$

$\frac{9.\underline{6}2 \times 10^{-3}}{0.125} \times 100\% = 7.8\%$; assumption not valid, solve with quadratic equation.

$x^2 + 7.4 \times 10^{-4} x - 9.25 \times 10^{-5} = 0 \quad x = 0.009\underline{2}55 = [H_3O^+] = [H_2C_6H_5O_3^-]$

and then:

$H_2C_6H_5O_3^-(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + HC_6H_5O_3^{2-}(aq)$

I	0.009255 M	0.009255	0.0
C	$-y$	y	y
E	$0.009255 - y$	$0.009255 + y$	y

$$K_a = \frac{[H_3O^+][HC_6H_5O_3^{2-}]}{[H_2C_6H_5O_3^-]} = \frac{(0.009255 + y)(y)}{(0.009255 - y)} = 1.7 \times 10^{-5}$$

Assume y is small. Then, $y = 1.7 \times 10^{-5}$

$\frac{1.7 \times 10^{-5}}{0.009255} \times 100\% = 1.8\%$; assumption valid

$[H_3O^+] = 1.7 \times 10^{-5}$ (from second ionization)

$[H_3O^+] = 0.009\underline{2}55 + 1.7 \times 10^{-5} = 0.009\underline{2}7$ M $\quad pH = -\log(0.009\underline{2}7) = 2.03$

15.111 **Given:** 0.500 M H_2SO_3 $K_{a_1} = 1.6 \times 10^{-2}$, $K_{a_2} = 6.4 \times 10^{-8}$ **Find:** concentration all species

Conceptual Plan: K_{a_1} is much larger than K_{a_2} so, use K_{a_1} to calculate $[H_3O^+]$. Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x.

Solution: $H_2SO_3(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + HSO_3^-(aq)$

I	0.500 M	0.0	0.0
C	$-x$	x	x
E	$0.500 - x$	x	x

$$K_a = \frac{[H_3O^+][HSO_3^-]}{[H_2SO_3]} = \frac{(x)(x)}{(0.500 - x)} = 1.6 \times 10^{-2}$$

$x^2 + 1.6 \times 10^{-2} x - 0.0080 = 0 \qquad x = 0.08\underline{1}8 = 0.082 \text{ M } [H_3O^+] = [HSO_3^-]$

Use the values from reaction 1 in reaction 2.

$HSO_3^-(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + SO_3^{2-}(aq)$

I	0.0818 M	0.0818	0.0
C	$-y$	y	y
E	$0.0818 - y$	$0.0818 + y$	y

$$K_a = \frac{[H_3O^+][SO_3^{2-}]}{[HSO_3^-]} = \frac{(0.0818 + y)(y)}{(0.0818 - y)} = 6.4 \times 10^{-8}$$

Assume y is small $y = 6.4 \times 10^{-8}$.

$[H_2SO_3] = 0.500 - 0.08\underline{1}8 = 0.418$ M

$[HSO_3^-] = x = 0.08\underline{1}8 = 0.082$ M

$[SO_3^{2-}] = y = 6.4 \times 10^{-8}$ M

$[H_3O^+] = x + y = 0.08\underline{1}8 = 0.082$ M

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0 \times 10^{-14}}{0.0818} = 1.2 \times 10^{-13}\text{ M}$$

15.112 **Given:** 0.155 M H_2CO_3 $K_{a_1} = 4.3 \times 10^{-7}$, $K_{a_2} = 5.6 \times 10^{-11}$ **Find:** concentration all species

Conceptual Plan: K_{a_1} is much larger than K_{a_2} so, use K_{a_1} to calculate $[H_3O^+]$ and $[HCO_3^-]$. Use K_{a_2} to find $[CO_3^-]$. Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x.

Solution: $H_2CO_3(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + HCO_3^-(aq)$

	H_2CO_3	H_3O^+	HCO_3^-
I	0.125 M	0.0	0.0
C	$-x$	x	x
E	$0.125 - x$	x	x

$$K_a = \frac{[H_3O^+][HCO_3^-]}{[H_2CO_3]} = \frac{(x)(x)}{(0.155 - x)} = 4.3 \times 10^{-7}$$

Assume x is small.

$x^2 = (4.3\times10^{-7})(0.155)$ $x = 2.58 \times 10^{-4}\text{ M} = [H_3O^+] = [HCO_3^-]$

Reaction 2

$HCO_3^-(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + CO_3^{2-}(aq)$

Since K_{a_2} is small, $y = [CO_3^{2-}] = K_{a_2}$

$[H_2CO_3] = 0.155 - 2.\underline{5}8 \times 10^{-4} = 0.1547 = 0.155$ M

$[HCO_3^-] = x = 2.\underline{5}8 \times 10^{-4} = 2.6 \times 10^{-4}$ M

$[CO_3^{2-}] = y = 5.6 \times 10^{-11}$ M

$[H_3O^+] = x = 2.\underline{5}8 \times 10^{-4} = 2.6 \times 10^{-4}$ M

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0 \times 10^{-14}}{2.\underline{5}8 \times 10^{-4}} = 3.88 \times 10^{-11}\text{ M} = 3.9 \times 10^{-11}\text{ M}$$

15.113 (a) **Given:** $[H_2SO_4] = 0.50$ M $K_{a_2} = 0.012$ **Find:** $[H_3O^+]$, pH

Conceptual Plan: The first ionization step is strong. Use K_{a_2} and reaction 2. Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x.

Solution: $H_2SO_4(aq) + H_2O(l) \rightarrow H_3O^+(aq) + HSO_4^-(aq)$ strong

0.50 M

$[H_3O^+] = [HSO_4^-] = 0.50$ M

$HSO_4^-(aq) + H_2O(l) \leftrightarrows H_3O^+(aq) + SO_4^{2-}(aq)$

	HSO_4^-	H_3O^+	SO_4^{2-}
I	0.50 M	0.50	0.0
C	$-x$	x	x
E	$0.500 - x$	$0.50 + x$	x

$$K_a = \frac{[H_3O^+][SO_4^{2-}]}{[HSO_4^-]} = \frac{(0.50 + x)(x)}{(0.50 - x)} = 0.012$$

$x^2 + 0.512\,x - 0.006 = 0$ $x = 0.01\underline{1}5 = [H_3O^+]$ from second ionization step

$[H_3O^+] = 0.50 + 0.012 = 0.51$ M

pH = - log(0.51) = 0.29

(b) **Given:** $[H_2SO_4] = 0.10$ M $K_{a_2} = 0.012$ **Find:** $[H_3O^+]$, pH

Conceptual Plan: The first ionization step is strong. Use K_{a_2} and reaction 2. Write a balanced reaction. Prepare an ICE table, represent the change with x, sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for x.