

- 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  $\text{HNO}_3(aq) \rightarrow \text{H}^+(aq) + \text{NO}_3^-(aq)$   
(b) acid  $\text{NH}_4^+(aq) \rightarrow \text{H}^+(aq) + \text{NH}_3(aq)$   
(c) base  $\text{KOH}(aq) \rightarrow \text{K}^+(aq) + \text{OH}^-(aq)$   
(d) acid  $\text{HC}_2\text{H}_3\text{O}_2(aq) \rightarrow \text{H}^+(aq) + \text{C}_2\text{H}_3\text{O}_2^-(aq)$
- 15.34 (a) base  $\text{NaOH}(aq) \rightarrow \text{Na}^+(aq) + \text{OH}^-(aq)$   
(b) acid  $\text{H}_2\text{SO}_4(aq) \rightarrow 2\text{H}^+(aq) + \text{SO}_4^{2-}(aq)$   
(c) acid  $\text{HBr}(aq) \rightarrow \text{H}^+(aq) + \text{Br}^-(aq)$   
(d) base  $\text{Sr}(\text{OH})_2(aq) \rightarrow \text{Sr}^{2+}(aq) + 2\text{OH}^-(aq)$
- 15.35 (a) Since  $\text{H}_2\text{CO}_3$  donates a proton to  $\text{H}_2\text{O}$ , it is the acid. After  $\text{H}_2\text{CO}_3$  donates the proton, it becomes  $\text{HCO}_3^-$ , the conjugate base. Since  $\text{H}_2\text{O}$  accepts a proton, it is the base. After  $\text{H}_2\text{O}$  accepts the proton, it becomes  $\text{H}_3\text{O}^+$ , the conjugate acid.
- (b) Since  $\text{H}_2\text{O}$  donates a proton to  $\text{NH}_3$ , it is the acid. After  $\text{H}_2\text{O}$  donates the proton, it becomes  $\text{OH}^-$ , the conjugate base. Since  $\text{NH}_3$  accepts a proton, it is the base. After  $\text{NH}_3$  accepts the proton, it becomes  $\text{NH}_4^+$ , the conjugate acid.
- (c) Since  $\text{HNO}_3$  donates a proton to  $\text{H}_2\text{O}$ , it is the acid. After  $\text{HNO}_3$  donates the proton, it becomes  $\text{NO}_3^-$ , the conjugate base. Since  $\text{H}_2\text{O}$  accepts a proton, it is the base. After  $\text{H}_2\text{O}$  accepts the proton, it becomes  $\text{H}_3\text{O}^+$ , the conjugate acid.
- (d) Since  $\text{H}_2\text{O}$  donates a proton to  $\text{C}_5\text{H}_5\text{N}$ , it is the acid. After  $\text{H}_2\text{O}$  donates the proton, it becomes  $\text{OH}^-$ , the conjugate base. Since  $\text{C}_5\text{H}_5\text{N}$  accepts a proton, it is the base. After  $\text{C}_5\text{H}_5\text{N}$  accepts the proton, it becomes  $\text{C}_5\text{H}_5\text{NH}^+$ , the conjugate acid.
- 15.36 (a) Since  $\text{HI}$  donates a proton to  $\text{H}_2\text{O}$ , it is the acid. After  $\text{HI}$  donates the proton, it becomes  $\text{I}^-$ , the conjugate base. Since  $\text{H}_2\text{O}$  accepts a proton, it is the base. After  $\text{H}_2\text{O}$  accepts the proton, it becomes  $\text{H}_3\text{O}^+$ , the conjugate acid.
- (b) Since  $\text{H}_2\text{O}$  donates a proton to  $\text{CH}_3\text{NH}_2$ , it is the acid. After  $\text{H}_2\text{O}$  donates the proton, it becomes  $\text{OH}^-$ , the conjugate base. Since  $\text{CH}_3\text{NH}_2$  accepts a proton, it is the base. After  $\text{CH}_3\text{NH}_2$  accepts the proton, it becomes  $\text{CH}_3\text{NH}_3^+$ , the conjugate acid.
- (c) Since  $\text{H}_2\text{O}$  donates a proton to  $\text{CO}_3^{2-}$ , it is the acid. After  $\text{H}_2\text{O}$  donates the proton, it becomes  $\text{OH}^-$ , the conjugate base. Since  $\text{CO}_3^{2-}$  accepts a proton, it is the base. After  $\text{CO}_3^{2-}$  accepts the proton, it becomes  $\text{HCO}_3^-$ , the conjugate acid.

- (d) Since HBr donates a proton to  $\text{H}_2\text{O}$ , it is the acid. After HBr donates the proton, it becomes  $\text{Br}^-$ , the conjugate base. Since  $\text{H}_2\text{O}$  accepts a proton, it is the base. After  $\text{H}_2\text{O}$  accepts the proton, it becomes  $\text{H}_3\text{O}^+$ , the conjugate acid.

15.37

- (a)  $\text{Cl}^- \quad \text{HCl}(aq) + \text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{O}^+(aq) + \text{Cl}^-(aq)$   
 (b)  $\text{HSO}_3^- \quad \text{H}_2\text{SO}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{HSO}_3^-(aq)$   
 (c)  $\text{CHO}_2^- \quad \text{HCHO}_2(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{CHO}_2^-(aq)$   
 (d)  $\text{F}^- \quad \text{HF}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{F}^-(aq)$

15.38

- (a)  $\text{NH}_4^+ \quad \text{NH}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^+(aq) + \text{OH}^-(aq)$   
 (b)  $\text{HClO}_4 \quad \text{HClO}_4(aq) + \text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{O}^+(aq) + \text{ClO}_4^-(aq)$   
 (c)  $\text{H}_2\text{SO}_4 \quad \text{H}_2\text{SO}_4(aq) + \text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{O}^+(aq) + \text{HSO}_4^-(aq)$   
 (d)  $\text{HCO}_3^- \quad \text{HCO}_3^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{CO}_3^{2-}(aq)$

15.39

- $\text{H}_2\text{PO}_4^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{HPO}_4^{2-}(aq)$   
 $\text{H}_2\text{PO}_4^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{PO}_4(aq) + \text{OH}^-(aq)$

15.40

- $\text{HCO}_3^-(aq) + \text{HS}^-(aq) \rightleftharpoons \text{H}_2\text{S}(aq) + \text{CO}_3^{2-}(aq)$   
 $\text{HCO}_3^-(aq) + \text{HS}^-(aq) \rightleftharpoons \text{H}_2\text{CO}_3(aq) + \text{S}^{2-}(aq)$

### Acid Strength and $K_a$

15.41

- (a)  $\text{HNO}_3$  is a strong acid.  
 (b)  $\text{HCl}$  is a strong acid.  
 (c)  $\text{HBr}$  is a strong acid.  
 (d)  $\text{H}_2\text{SO}_3$  is a weak acid.  $\text{H}_2\text{SO}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{HSO}_3^-(aq)$   

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{HSO}_3^-]}{[\text{H}_2\text{SO}_3]}$$

15.42

- (a)  $\text{HF}$  is a weak acid.  $\text{HF}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{F}^-(aq)$   

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HF}]}$$
  
 (b)  $\text{HCHO}_2$  is a weak acid.  $\text{HCHO}_2(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{CHO}_2^-(aq)$   

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CHO}_2^-]}{[\text{HCHO}_2]}$$
  
 (c)  $\text{H}_2\text{SO}_4$  is a strong acid.  
 (d)  $\text{H}_2\text{CO}_3$  is a weak acid.  $\text{H}_2\text{CO}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{HCO}_3^-(aq)$   

$$K_{a1} = \frac{[\text{H}_3\text{O}^+][\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]}$$

15.43

- (a) contains no HA, 10  $\text{H}^+$ , and 10  $\text{A}^-$   
 (b) contains 3 HA, 3  $\text{H}^+$ , and 7  $\text{A}^-$   
 (c) contains 9 HA, 1  $\text{H}^+$ , and 1  $\text{A}^-$

So, solution a > solution b > solution c.

15.44 HCl is a strong acid,  $K_a(\text{HF}) = 3.5 \times 10^{-4}$ ,  $K_a(\text{HClO}) = 2.9 \times 10^{-8}$ ,  $K_a(\text{HC}_6\text{H}_5\text{O}) = 1.3 \times 10^{-10}$ .

The larger the value of  $K_a$  the stronger the acid and the greater the  $[\text{H}_3\text{O}^+]$ .

The order of decreasing  $[\text{H}_3\text{O}^+]$  is  $\text{HCl} > \text{HF} > \text{HClO} > \text{HC}_6\text{H}_5\text{O}$ .

15.45

- (a)  $\text{F}^-$  is a stronger base than  $\text{Cl}^-$ .  
 $\text{F}^-$  is the conjugate base of HF (a weak acid),  $\text{Cl}^-$  is the conjugate base of HCl (a strong acid), the weaker the acid, the stronger the conjugate base.
- (b)  $\text{NO}_2^-$  is a stronger base than  $\text{NO}_3^-$ .  
 $\text{NO}_2^-$  is the conjugate base of  $\text{HNO}_2$  (a weak acid),  $\text{NO}_3^-$  is the conjugate base of  $\text{HNO}_3$  (a strong acid), the weaker the acid, the stronger the conjugate base.
- (c)  $\text{ClO}^-$  is a stronger base than  $\text{F}^-$ .  
 $\text{F}^-$  is the conjugate base of HF ( $K_a = 3.5 \times 10^{-4}$ ),  $\text{ClO}^-$  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)  $\text{ClO}_2^-$  is a stronger base than  $\text{ClO}_4^-$ .  
 $\text{ClO}_2^-$  is the conjugate base of  $\text{HClO}_2$  (a weak acid),  $\text{ClO}_4^-$  is the conjugate base of  $\text{HClO}_4$  (a strong acid), the weaker the acid, the stronger the conjugate base.
- (b)  $\text{H}_2\text{O}$  is a stronger base than  $\text{Cl}^-$ .  
 $\text{H}_2\text{O}$  is the conjugate base of  $\text{H}_3\text{O}^+$ ,  $\text{Cl}^-$  is the conjugate base of HCl (a strong acid), the weaker the acid, the stronger the conjugate base.
- (c)  $\text{CN}^-$  is stronger base than  $\text{ClO}^-$ .  
 $\text{CN}^-$  is the conjugate base of HCN ( $K_a = 4.9 \times 10^{-10}$ ),  $\text{ClO}^-$  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:**  $K_w = 1.0 \times 10^{-14}$ ,  $[\text{H}_3\text{O}^+] = 1.2 \times 10^{-8} \text{ M}$  **Find:**  $[\text{OH}^-]$

**Conceptual Plan:**  $[\text{H}_3\text{O}^+] \rightarrow [\text{OH}^-]$

$$K_w = 1.0 \times 10^{-14} = [\text{H}_3\text{O}^+][\text{OH}^-]$$

**Solution:**

$$K_w = 1.0 \times 10^{-14} = (1.2 \times 10^{-8})[\text{OH}^-]$$

$$[\text{OH}^-] = 8.3 \times 10^{-7} \text{ M}$$

$[\text{OH}^-] > [\text{H}_3\text{O}^+]$  so the solution is basic.

(b) **Given:**  $K_w = 1.0 \times 10^{-14}$ ,  $[\text{H}_3\text{O}^+] = 8.5 \times 10^{-5} \text{ M}$  **Find:**  $[\text{OH}^-]$

**Conceptual Plan:**  $[\text{H}_3\text{O}^+] \rightarrow [\text{OH}^-]$

$$K_w = 1.0 \times 10^{-14} = [\text{H}_3\text{O}^+][\text{OH}^-]$$

**Solution:**

$$K_w = 1.0 \times 10^{-14} = (8.5 \times 10^{-5})[\text{OH}^-]$$

$$[\text{OH}^-] = 1.2 \times 10^{-10} \text{ M}$$

$[\text{H}_3\text{O}^+] > [\text{OH}^-]$  so the solution is acidic.

(c) **Given:**  $K_w = 1.0 \times 10^{-14}$ ,  $[\text{H}_3\text{O}^+] = 3.5 \times 10^{-2} \text{ M}$  **Find:**  $[\text{OH}^-]$

**Conceptual Plan:**  $[\text{H}_3\text{O}^+] \rightarrow [\text{OH}^-]$

$$K_w = 1.0 \times 10^{-14} = [\text{H}_3\text{O}^+][\text{OH}^-]$$

**Solution:**

$$K_w = 1.0 \times 10^{-14} = (3.5 \times 10^{-2})[\text{OH}^-]$$

$$[\text{OH}^-] = 2.9 \times 10^{-13} \text{ M}$$

$[\text{H}_3\text{O}^+] > [\text{OH}^-]$  so the solution is acidic.

15.48 (a) **Given:**  $K_w = 1.0 \times 10^{-14}$ ,  $[\text{OH}^-] = 1.1 \times 10^{-9} \text{ M}$  **Find:**  $[\text{H}_3\text{O}^+]$

**Conceptual Plan:**  $[\text{OH}^-] \rightarrow [\text{H}_3\text{O}^+]$

$$K_w = 1.0 \times 10^{-14} = [\text{H}_3\text{O}^+][\text{OH}^-]$$



(c) Given: pH = 2.87 Find:  $[\text{H}_3\text{O}^+]$ ,  $[\text{OH}^-]$ Conceptual Plan: pH  $\rightarrow$   $[\text{H}_3\text{O}^+]$   $\rightarrow$   $[\text{OH}^-]$ 

$$\text{pH} = -\log[\text{H}_3\text{O}^+] \quad K_w = 1.0 \times 10^{-14} = [\text{H}_3\text{O}^+][\text{OH}^-]$$

Solution: pH =  $-\log[\text{H}_3\text{O}^+]$  2.87 =  $-\log[\text{H}_3\text{O}^+]$ 

$$-2.87 = \log[\text{H}_3\text{O}^+] \quad 10^{-2.87} = 10^{\log[\text{H}_3\text{O}^+]}$$

$$10^{-2.87} = [\text{H}_3\text{O}^+] \quad [\text{H}_3\text{O}^+] = 1.3 \times 10^{-3}$$

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

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

15.51

$$\text{pH} = -\log[\text{H}_3\text{O}^+] \quad K_w = 1.0 \times 10^{-14} = [\text{H}_3\text{O}^+][\text{OH}^-]$$

$[\text{H}_3\text{O}^+]$	$[\text{OH}^-]$	pH	Acidic or basic
$7.1 \times 10^{-4}$	$1.4 \times 10^{-11}$	3.15	acidic
$3.7 \times 10^{-9}$	$2.7 \times 10^{-6}$	8.43	basic
$8 \times 10^{-12}$	$1 \times 10^{-3}$	11.1	basic
$6.2 \times 10^{-4}$	$1.6 \times 10^{-11}$	3.20	acidic

$$[\text{H}_3\text{O}^+] = 10^{-3.15} = 7.1 \times 10^{-4} \quad [\text{OH}^-] = \frac{1.0 \times 10^{-14}}{7.1 \times 10^{-4}} = 1.4 \times 10^{-11}$$

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

$$[\text{H}_3\text{O}^+] = 10^{-11.1} = 8 \times 10^{-12} \quad [\text{OH}^-] = \frac{1.0 \times 10^{-14}}{8 \times 10^{-12}} = 1 \times 10^{-3}$$

$$[\text{H}_3\text{O}^+] = \frac{1.0 \times 10^{-14}}{1.6 \times 10^{-11}} = 6.2 \times 10^{-4} \quad \text{pH} = -\log(6.2 \times 10^{-4}) = 3.20$$

15.52

$$\text{pH} = -\log[\text{H}_3\text{O}^+] \quad K_w = 1.0 \times 10^{-14} = [\text{H}_3\text{O}^+][\text{OH}^-]$$

$[\text{H}_3\text{O}^+]$	$[\text{OH}^-]$	pH	Acidic or basic
$3.5 \times 10^{-3}$	$2.9 \times 10^{-12}$	2.46	acidic
$2.6 \times 10^{-8}$	$3.8 \times 10^{-7}$	7.58	basic
$1.8 \times 10^{-9}$	$5.6 \times 10^{-6}$	8.74	basic
$7.1 \times 10^{-8}$	$1.4 \times 10^{-7}$	7.15	basic

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

$$[\text{H}_3\text{O}^+] = \frac{1.0 \times 10^{-14}}{3.8 \times 10^{-7}} = 2.6 \times 10^{-8} \quad \text{pH} = -\log(2.6 \times 10^{-8}) = 7.58$$

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

$$[\text{H}_3\text{O}^+] = 10^{-7.15} = 7.1 \times 10^{-8} \quad [\text{OH}^-] = \frac{1.0 \times 10^{-14}}{7.1 \times 10^{-8}} = 1.4 \times 10^{-7}$$

15.53

Given:  $K_w = 2.4 \times 10^{-14}$  at  $37^\circ\text{C}$  Find:  $[\text{H}_3\text{O}^+]$ , pHConceptual Plan:  $K_w \rightarrow [\text{H}_3\text{O}^+] \rightarrow \text{pH}$ 

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] \quad \text{pH} = -\log[\text{H}_3\text{O}^+]$$

Solution:  $\text{H}_2\text{O}(l) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{OH}^-(aq)$ 

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

$$[\text{H}_3\text{O}^+] = [\text{OH}^-] = \sqrt{K_w} = \sqrt{2.4 \times 10^{-14}} = 1.5 \times 10^{-7}$$

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log(1.5 \times 10^{-7}) = 6.81$$

**Check:** The value of  $K_w$  increased indicating more products formed, so the  $[\text{H}_3\text{O}^+]$  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:**  $[\text{H}_3\text{O}^+]$ ,  $[\text{OH}^-]$ , pH  
**Conceptual Plan:**  $[\text{HCl}] \rightarrow [\text{H}_3\text{O}^+] \rightarrow \text{pH}$  and then  $[\text{H}_3\text{O}^+] \rightarrow [\text{OH}^-]$   
 $[\text{HCl}] \rightarrow [\text{H}_3\text{O}^+] \text{ pH} = -\log[\text{H}_3\text{O}^+] \quad [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$   
**Solution:** 0.25 M HCl = 0.25 M  $\text{H}_3\text{O}^+$   $\text{pH} = -\log(0.25) = 0.60$   
 $[\text{OH}^-] = 1.0 \times 10^{-14} / 0.25 \text{ M} = 4.0 \times 10^{-14}$   
**Check:** HCl is a strong acid with a relatively high concentration, so we expect the pH to be low and the  $[\text{OH}^-]$  to be small.
- (b) **Given:** 0.015 M  $\text{HNO}_3$  (strong acid) **Find:**  $[\text{H}_3\text{O}^+]$ ,  $[\text{OH}^-]$ , pH  
**Conceptual Plan:**  $[\text{HNO}_3] \rightarrow [\text{H}_3\text{O}^+] \rightarrow \text{pH}$  and then  $[\text{H}_3\text{O}^+] \rightarrow [\text{OH}^-]$   
 $[\text{HNO}_3] \rightarrow [\text{H}_3\text{O}^+] \text{ pH} = -\log[\text{H}_3\text{O}^+] \quad [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$   
**Solution:** 0.015 M  $\text{HNO}_3 = 0.015 \text{ M } \text{H}_3\text{O}^+$   $\text{pH} = -\log(0.015) = 1.82$   
 $[\text{OH}^-] = 1.0 \times 10^{-14} / 0.015 \text{ M} = 6.7 \times 10^{-13}$   
**Check:**  $\text{HNO}_3$  is a strong acid, so we expect the pH to be low and the  $[\text{OH}^-]$  to be small.
- (c) **Given:** 0.052 M HBr and 0.020 M  $\text{HNO}_3$  (strong acids) **Find:**  $[\text{H}_3\text{O}^+]$ ,  $[\text{OH}^-]$ , pH  
**Conceptual Plan:**  $[\text{HBr}] + [\text{HNO}_3] \rightarrow [\text{H}_3\text{O}^+] \rightarrow \text{pH}$  and then  $[\text{H}_3\text{O}^+] \rightarrow [\text{OH}^-]$   
 $[\text{HBr}] + [\text{HNO}_3] \rightarrow [\text{H}_3\text{O}^+] \text{ pH} = -\log[\text{H}_3\text{O}^+] \quad [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$   
**Solution:** 0.052 M HBr = 0.052 M  $\text{H}_3\text{O}^+$  and 0.020 M  $\text{HNO}_3 = 0.020 \text{ M } \text{H}_3\text{O}^+$   
Total  $\text{H}_3\text{O}^+ = 0.052 \text{ M} + 0.020 \text{ M} = 0.072 \text{ M}$   $\text{pH} = -\log(0.072) = 1.14$   
 $[\text{OH}^-] = 1.0 \times 10^{-14} / 0.072 \text{ M} = 1.4 \times 10^{-13}$   
**Check:** HBr and  $\text{HNO}_3$  are both strong acids and completely dissociate. This gives a relatively high concentration, so we expect the pH to be low and the  $[\text{OH}^-]$  to be small.
- (d) **Given:**  $\text{HNO}_3 = 0.655\%$  by mass,  $d_{\text{solution}} = 1.01 \text{ g/mL}$  **Find:**  $[\text{H}_3\text{O}^+]$ ,  $[\text{OH}^-]$ , pH  
**Conceptual Plan:**  
 $\% \text{ mass } \text{HNO}_3 \rightarrow \text{g } \text{HNO}_3 \rightarrow \text{mol } \text{HNO}_3 \text{ and then g soln} \rightarrow \text{mL soln} \rightarrow \text{L soln} \rightarrow \text{M } \text{HNO}_3$   
 $\frac{\%}{100} \times \frac{\text{mol } \text{HNO}_3}{63.018 \text{ g } \text{HNO}_3} \times \frac{1.01 \text{ g soln}}{\text{mL soln}} \times \frac{1000 \text{ mL soln}}{\text{L soln}} \times \frac{\text{mol } \text{HNO}_3}{\text{L soln}}$   
 $\rightarrow \text{M } \text{H}_3\text{O}^+ \rightarrow \text{pH}$  and then  $[\text{H}_3\text{O}^+] \rightarrow [\text{OH}^-]$   
 $[\text{HNO}_3] \rightarrow [\text{H}_3\text{O}^+] \text{ pH} = -\log[\text{H}_3\text{O}^+] \quad [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$   
**Solution:**  $\frac{0.655 \text{ g } \text{HNO}_3}{100 \text{ g soln}} \times \frac{1 \text{ mol } \text{HNO}_3}{63.018 \text{ g } \text{HNO}_3} \times \frac{1.01 \text{ g soln}}{\text{mL soln}} \times \frac{1000 \text{ mL soln}}{\text{L soln}} = 0.105 \text{ M } \text{HNO}_3$   
 $0.105 \text{ M } \text{HNO}_3 = 0.105 \text{ M } \text{H}_3\text{O}^+$   $\text{pH} = -\log(0.105) = 0.979$   
 $[\text{OH}^-] = 1.00 \times 10^{-14} / 0.105 \text{ M} = 9.52 \times 10^{-14}$   
**Check:**  $\text{HNO}_3$  is a strong acid and completely dissociates. This gives a relatively high concentration, so we expect the pH to be low and the  $[\text{OH}^-]$  to be small.
- 15.56 (a) **Given:** 0.048 M HI (strong acid) **Find:**  $[\text{H}_3\text{O}^+]$ ,  $[\text{OH}^-]$ , pH  
**Conceptual Plan:**  $[\text{HI}] \rightarrow [\text{H}_3\text{O}^+] \rightarrow \text{pH}$  and then  $[\text{H}_3\text{O}^+] \rightarrow [\text{OH}^-]$   
 $[\text{HI}] \rightarrow [\text{H}_3\text{O}^+] \text{ pH} = -\log[\text{H}_3\text{O}^+] \quad [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$   
**Solution:** 0.048 M HI = 0.048 M  $\text{H}_3\text{O}^+$   $\text{pH} = -\log(0.048) = 1.32$   
 $[\text{OH}^-] = 1.0 \times 10^{-14} / 0.048 \text{ M} = 2.1 \times 10^{-13}$   
**Check:** HI is a strong acid with a relatively high concentration, so we expect the pH to be low and the  $[\text{OH}^-]$  to be small.

15.62 Given: 0.200 M formic acid.  $K_a = 1.8 \times 10^{-4}$  Find:  $[\text{H}_3\text{O}^+]$ , 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  $[\text{H}_3\text{O}^+]$  and pH.

**Solution:**  $\text{HCH}_2\text{O}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{CH}_2\text{O}^-(aq)$

I	0.200 M	0.0	0.0
C	- $x$	$x$	$x$
E	$0.200 - x$	$x$	$x$

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CH}_2\text{O}^-]}{[\text{HCH}_2\text{O}]} = \frac{(x)(x)}{(0.200 - x)} = 1.8 \times 10^{-4}$$

Assume  $x$  is small compared to 0.200.

$$x^2 = (1.8 \times 10^{-4})(0.200) \quad x = 6.0 \times 10^{-3} \text{ M} = [\text{H}_3\text{O}^+]$$

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  $\text{HNO}_2$ .  $K_a = 4.6 \times 10^{-4}$  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  $[\text{H}_3\text{O}^+]$  and pH.

**Solution:**  $\text{HNO}_2(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{NO}_2^-(aq)$

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

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{NO}_2^-]}{[\text{HNO}_2]} = \frac{(x)(x)}{(0.500 - x)} = 4.6 \times 10^{-4}$$

Assume  $x$  is small compared to 0.500.

$$x^2 = (4.6 \times 10^{-4})(0.500) \quad x = 0.015 \text{ M} = [\text{H}_3\text{O}^+]$$

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  $\text{HNO}_2$ .  $K_a = 4.6 \times 10^{-4}$  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  $[\text{H}_3\text{O}^+]$  and pH.

**Solution:**  $\text{HNO}_2(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{NO}_2^-(aq)$

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

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{NO}_2^-]}{[\text{HNO}_2]} = \frac{(x)(x)}{(0.100 - x)} = 4.6 \times 10^{-4}$$

Assume  $x$  is small compared to 0.100.

$$x^2 = (4.6 \times 10^{-4})(0.100) \quad x = 0.0068 \text{ M} = [\text{H}_3\text{O}^+]$$

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

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

$$x = 0.00656$$

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

(c) Given: 0.100 M  $\text{HNO}_2$ .  $K_a = 4.6 \times 10^{-4}$  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  $[\text{H}_3\text{O}^+]$  and pH.





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

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{NO}_2^-]}{[\text{HNO}_2]} = \frac{(x)(x)}{(0.0100 - x)} = 4.6 \times 10^{-4}$$

Assume  $x$  is small compared to 0.100.

$$x^2 = (4.6 \times 10^{-4})(0.0100) \quad x = 0.0021 \text{ M} = [\text{H}_3\text{O}^+]$$

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

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

$$x = 0.0019$$

$$\text{pH} = -\log(0.0019) = 2.72$$

- 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  $[\text{H}_3\text{O}^+]$  and pH.



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

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HF}]} = \frac{(x)(x)}{(0.250 - x)} = 3.5 \times 10^{-4}$$

Assume  $x$  is small compared to 0.250.

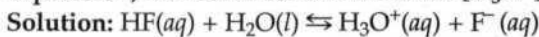
$$x^2 = (3.5 \times 10^{-4})(0.250) \quad x = 0.00935 \text{ M} = [\text{H}_3\text{O}^+]$$

Check assumption:  $\frac{0.00935}{0.250} \times 100\% = 3.7\%$  assumption valid.

$$\text{pH} = -\log(0.00935) = 2.03$$

- (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  $[\text{H}_3\text{O}^+]$  and pH.



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

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HF}]} = \frac{(x)(x)}{(0.0500 - x)} = 3.5 \times 10^{-4}$$

Assume  $x$  is small compared to 0.0500.

$$x^2 = (3.5 \times 10^{-4})(0.0500) \quad x = 0.00418 \text{ M} = [\text{H}_3\text{O}^+]$$

Check assumption:  $\frac{0.00418}{0.050} \times 100\% = 8.3\%$  assumption not valid, solve with quadratic equation.

$$x^2 + 3.5 \times 10^{-4}x - 1.75 \times 10^{-5} = 0 \quad x = 0.0040 = [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log(0.0040) = 2.40$$

- (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  $[\text{H}_3\text{O}^+]$  and pH.



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

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HF}]} = \frac{(x)(x)}{(0.0250 - x)} = 3.5 \times 10^{-4}$$

Assume  $x$  is small compared to 0.0250.

$$x^2 = (3.5 \times 10^{-4})(0.0250) \quad x = 0.00295 \text{ M} = [\text{H}_3\text{O}^+]$$

Check assumption:  $\frac{0.00295}{0.0250} \times 100\% = 11.8\%$  assumption not valid, solve with quadratic equation.

$$x^2 + 3.5 \times 10^{-4}x - 8.75 \times 10^{-6} = 0 \quad x = 0.00279 = [\text{H}_3\text{O}^+]$$

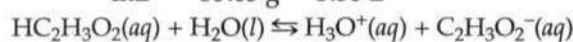
$$\text{pH} = -\log(0.00279) = 2.55$$

- 15.65 **Given:** 15.0 mL glacial acetic,  $d = 1.05 \text{ g/mL}$ , dilute to 1.50 L,  $K_a = 1.8 \times 10^{-5}$  **Find:** pH  
**Conceptual Plan:** mL acetic acid  $\rightarrow$  g acetic acid  $\rightarrow$  mol acetic acid  $\rightarrow$  M and then write a balanced reaction.

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

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  $[\text{H}_3\text{O}^+]$  and pH.

$$\text{Solution: } 15.0 \text{ mL} \times \frac{1.05 \text{ g}}{\text{mL}} \times \frac{1 \text{ mol}}{60.05 \text{ g}} \times \frac{1}{1.50 \text{ L}} = 0.1748 \text{ M}$$



I	0.1748 M	0.0	0.0
C	- $x$	$x$	$x$
E	$0.1748 - x$	$x$	$x$

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]} = \frac{(x)(x)}{(0.1748 - x)} = 1.8 \times 10^{-5}$$

Assume  $x$  is small compared to 0.1748.

$$x^2 = (1.8 \times 10^{-5})(0.1748) \quad x = 0.00177 \text{ M} = [\text{H}_3\text{O}^+]$$

Check assumption:  $\frac{0.00177}{0.1748} \times 100\% = 1.0\%$  assumption valid.

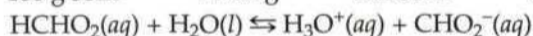
$$\text{pH} = -\log(0.00177) = 2.75$$

- 15.66 **Given:** 1.35% formic acid,  $d = 1.01 \text{ g/mL}$ ,  $K_a = 1.8 \times 10^{-4}$  **Find:** pH  
**Conceptual Plan:** % formic acid  $\rightarrow$  g formic acid  $\rightarrow$  mol and g soln  $\rightarrow$  mL soln  $\rightarrow$  L soln and then M

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

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  $[\text{H}_3\text{O}^+]$  and pH.

$$\text{Solution: } \frac{1.35 \text{ g HCHO}_2}{100 \text{ g soln}} \times \frac{\text{mol HCHO}_2}{46.03 \text{ g}} \times \frac{1.01 \text{ g soln}}{\text{mL soln}} \times \frac{1000 \text{ mL soln}}{\text{L soln}} = 0.2962 \text{ M}$$



I	0.2962 M	0.0	0.0
C	- $x$	$x$	$x$
E	$0.2962 - x$	$x$	$x$

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CHO}_2^-]}{[\text{HCHO}_2]} = \frac{(x)(x)}{(0.2962 - x)} = 1.8 \times 10^{-4}$$

Assume  $x$  is small compared to 0.2962.

$$x^2 = (1.8 \times 10^{-4})(0.2962) \quad x = 0.00730 \text{ M} = [\text{H}_3\text{O}^+]$$

Check assumption:  $\frac{0.00730}{0.2962} \times 100\% = 2.5\%$  assumption valid.

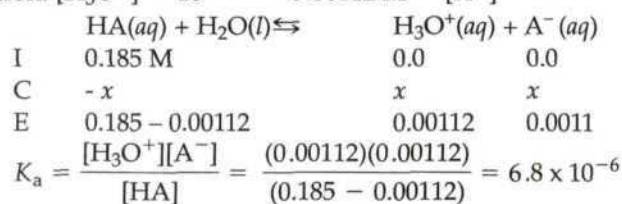
$$\text{pH} = -\log(0.00730) = 2.14$$

15.67

- Given:** 0.185 M HA,  $\text{pH} = 2.95$  **Find:**  $K_a$   
**Conceptual Plan:** pH  $\rightarrow$   $[\text{H}_3\text{O}^+]$  and then write a balanced reaction. Prepare an ICE table, calculate equilibrium concentrations, and then plug into the equilibrium expression to solve for  $K_a$ .



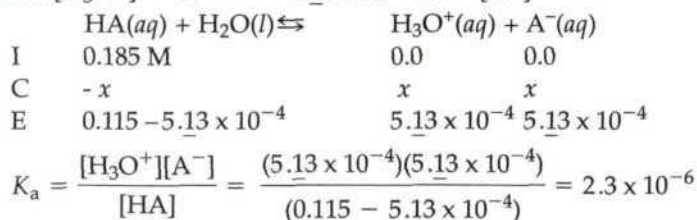
**Solution:**  $[\text{H}_3\text{O}^+] = 10^{-2.95} = 0.00112 \text{ M} = [\text{A}^-]$



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

**Conceptual Plan:** pH  $\rightarrow$   $[\text{H}_3\text{O}^+]$  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:**  $[\text{H}_3\text{O}^+] = 10^{-3.29} = 5.13 \times 10^{-4} \text{ M} = [\text{A}^-]$



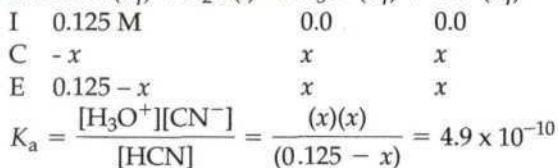
15.69

**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  $x \rightarrow$  % ionization.

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

**Solution:**  $\text{HCN}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{CN}^-(aq)$



Assume  $x$  is small compared to 0.125.

$$x^2 = (4.9 \times 10^{-10})(0.125) \quad x = 7.83 \times 10^{-6}$$

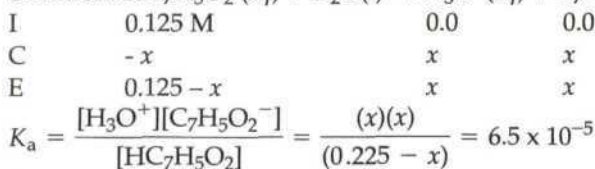
$$\% \text{ ionization} = \frac{7.83 \times 10^{-6}}{0.125} \times 100 = 0.0063\% \text{ ionized}$$

15.70 **Given:** 0.225 M  $\text{HC}_7\text{H}_5\text{O}_2$   $K_a = 6.5 \times 10^{-5}$  **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  $x \rightarrow$  % ionization.

$$\% \text{ ionization} = \frac{x}{[\text{HC}_7\text{H}_5\text{O}_2]_{\text{original}}} \times 100$$

**Solution:**  $\text{HC}_7\text{H}_5\text{O}_2(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{C}_7\text{H}_5\text{O}_2^-(aq)$



Assume  $x$  is small compared to 0.225.

$$x^2 = (6.5 \times 10^{-5})(0.1225) \quad x = 0.00382$$

$$\% \text{ ionization} = \frac{0.00382}{0.225} \times 100 = 1.7\% \text{ ionized}$$

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CHO}_2^-]}{[\text{HCHO}_2]} = \frac{(x)(x)}{(0.100 - x)} = 1.8 \times 10^{-4}$$

Assume  $x$  is small compared to 0.100.

$$x^2 = (1.8 \times 10^{-4})(0.100) \quad x = 0.00424$$

$$\% \text{ ionization} = \frac{0.00424}{0.100} \times 100 = 4.2\% \text{ ionized}$$

(d) **Given:** 0.0500 M  $\text{HCHO}_2$   $K_a = 1.8 \times 10^{-4}$  **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.

$$\% \text{ ionization} = \frac{x}{[\text{HCHO}_2]_{\text{original}}} \times 100$$

**Solution:**  $\text{HCHO}_2(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{CHO}_2^-(aq)$

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

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CHO}_2^-]}{[\text{HCHO}_2]} = \frac{(x)(x)}{(0.0500 - x)} = 1.8 \times 10^{-4}$$

Assume  $x$  is small compared to 0.0500.

$$x^2 = (1.8 \times 10^{-4})(0.0500) \quad x = 0.00300$$

$$\% \text{ ionization} = \frac{0.00300}{0.0500} \times 100\% = 6.0\% \text{ ionized}$$

$x$  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  $x = 0.00291$

$$\% \text{ ionization} = \frac{0.00291}{0.0500} \times 100 = 5.8\%$$

15.73

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

**Conceptual Plan:**  $\text{M} \rightarrow [\text{H}_3\text{O}^+] \rightarrow K_a$  and then write a balanced reaction, determine equilibrium concentration, and plug into the equilibrium expression.

**Solution:**  $(0.148 \text{ M HA})(0.0155) = 0.002294$   $[\text{H}_3\text{O}^+] = [\text{A}^-]$

$\text{HA}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{A}^-(aq)$

I	0.148 M	0.0	0.0
C	- $x$	$x$	$x$
E	0.148 - 0.002294	0.002294	0.002294

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} = \frac{(0.002294)(0.002294)}{(0.148 - 0.002294)} = 3.61 \times 10^{-5}$$

15.74

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

**Conceptual Plan:**  $\text{M} \rightarrow [\text{H}_3\text{O}^+] \rightarrow K_a$  and then write a balanced reaction, determine equilibrium concentration, and plug into the equilibrium expression.

**Solution:**  $(0.085 \text{ M HA})(0.0059) = 5.02 \times 10^{-4}$   $[\text{H}_3\text{O}^+] = [\text{A}^-]$

$\text{HA}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{A}^-(aq)$

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

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} = \frac{(5.02 \times 10^{-4})(5.02 \times 10^{-4})}{(0.085 - 5.02 \times 10^{-4})} = 3.0 \times 10^{-6}$$

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.

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

- 15.78 (a) **Given:** 0.075 M HNO<sub>3</sub> (strong acid), 0.175 M HC<sub>7</sub>H<sub>5</sub>O<sub>2</sub> (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<sub>3</sub>O<sup>+</sup>] and then pH.  
**Solution:** 0.075 M HNO<sub>3</sub> = 0.075 M [H<sub>3</sub>O<sup>+</sup>] pH = -log(0.075) = 1.12
- (b) **Given:** 0.020 M HBr (strong acid), 0.015 M HClO<sub>4</sub> (strong acid) **Find:** pH  
**Conceptual Plan:** Since the mixture is a strong acid and another strong acid, the [H<sub>3</sub>O<sup>+</sup>] is the sum of the concentration of both acids. Determine pH.  
**Solution:** 0.020 M HBr = 0.020 [H<sub>3</sub>O<sup>+</sup>], 0.015 M HClO<sub>4</sub> = 0.015 M [H<sub>3</sub>O<sup>+</sup>]  
[H<sub>3</sub>O<sup>+</sup>] = 0.020 + 0.015 = 0.035 M pH = -log(0.035) = 1.46
- (c) **Given:** 0.095M HF, K<sub>a</sub> = 3.5 × 10<sup>-4</sup>; 0.225 M HC<sub>6</sub>H<sub>5</sub>O<sub>2</sub>, K<sub>a</sub> = 1.3 × 10<sup>-10</sup> **Find:** pH  
**Conceptual Plan:** Since the values of K are more than 10<sup>1</sup> 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<sub>2</sub>O(l) ⇌ H<sub>3</sub>O<sup>+</sup>(aq) + F<sup>-</sup>(aq)
- |   |           |     |     |
|---|-----------|-----|-----|
| I | 0.095 M   | 0.0 | 0.0 |
| C | -x        | x   | x   |
| E | 0.095 - x | x   | x   |
- $$K_a = \frac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HF}]} = \frac{(x)(x)}{(0.095 - x)} = 3.5 \times 10^{-4}$$
- $$x^2 = (3.5 \times 10^{-4})(0.095) \quad x = 0.00577 \text{ M} = [\text{H}_3\text{O}^+]$$
- $$(0.00577/0.095) \times 100 = 6.1\%. \text{ Assumption is invalid since } 6.1\% \text{ is greater than the } 5.0\% \text{ limit.}$$
- Solve for x using the quadratic formula;  
 $x^2 + 3.5 \times 10^{-4}x - 3.325 \times 10^{-5} = 0$   
 $x = 0.00559 \text{ M} = [\text{H}_3\text{O}^+]$   
pH = -log(0.00559) = 2.25
- (d) **Given:** 0.100 M HCHO<sub>2</sub>, K<sub>a</sub> = 1.8 × 10<sup>-4</sup>; 0.050 M HClO, K<sub>a</sub> = 2.9 × 10<sup>-8</sup> **Find:** pH  
**Conceptual Plan:** Since the values of K are more than 10<sup>1</sup> 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<sub>2</sub>(aq) + H<sub>2</sub>O(l) ⇌ H<sub>3</sub>O<sup>+</sup>(aq) + C<sub>2</sub>H<sub>3</sub>O<sub>2</sub><sup>-</sup>(aq)
- |   |           |     |     |
|---|-----------|-----|-----|
| I | 0.100 M   | 0.0 | 0.0 |
| C | -x        | x   | x   |
| E | 0.100 - x | x   | x   |
- $$K_a = \frac{[\text{H}_3\text{O}^+][\text{CHO}_2^-]}{[\text{HCHO}_2]} = \frac{(x)(x)}{(0.100 - x)} = 1.8 \times 10^{-4}$$
- Assume x is small compared to 0.0500.  
 $x^2 = (1.8 \times 10^{-4})(0.100) \quad x = 0.00424$   
pH = -log(0.00424) = 2.37

## Base Solutions

- 15.79 (a) **Given:** 0.15 M NaOH **Find:** [OH<sup>-</sup>], [H<sub>3</sub>O<sup>+</sup>], pH, pOH  
**Conceptual Plan:** [NaOH] → [OH<sup>-</sup>] → [H<sub>3</sub>O<sup>+</sup>] → pH → pOH  
 $K_w = [\text{H}_3\text{O}^+][\text{OH}^-] \quad \text{pH} = -\log[\text{H}_3\text{O}^+] \quad \text{pH} + \text{pOH} = 14$   
**Solution:** [OH<sup>-</sup>] = [NaOH] = 0.15M  

$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{0.15\text{M}} = 6.7 \times 10^{-14} \text{ M}$$
  
pH = -log(6.7 × 10<sup>-14</sup>) = 13.17  
pOH = 14.00 - 13.18 = 0.83



- (b) **Given:**  $1.5 \times 10^{-3}$  M  $\text{Ca(OH)}_2$  **Find:**  $[\text{OH}^-]$ ,  $[\text{H}_3\text{O}^+]$ , pH, pOH  
**Conceptual Plan:**  $[\text{Ca(OH)}_2] \rightarrow [\text{OH}^-] \rightarrow [\text{H}_3\text{O}^+] \rightarrow \text{pH} \rightarrow \text{pOH}$   
 $K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$      $\text{pH} = -\log[\text{H}_3\text{O}^+]$      $\text{pH} + \text{pOH} = 14$   
**Solution:**  $[\text{OH}^-] = 2[\text{Ca(OH)}_2] = 2(1.5 \times 10^{-3}) = 0.0030$  M  

$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{0.0030 \text{ M}} = 3.33 \times 10^{-12} \text{ M}$$

$$\text{pH} = -\log(3.33 \times 10^{-12}) = 11.48$$

$$\text{pOH} = 14.00 - 11.48 = 2.52$$
- (c) **Given:**  $4.8 \times 10^{-4}$  M  $\text{Sr(OH)}_2$  **Find:**  $[\text{OH}^-]$ ,  $[\text{H}_3\text{O}^+]$ , pH, pOH  
**Conceptual Plan:**  $[\text{Sr(OH)}_2] \rightarrow [\text{OH}^-] \rightarrow [\text{H}_3\text{O}^+] \rightarrow \text{pH} \rightarrow \text{pOH}$   
 $K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$      $\text{pH} = -\log[\text{H}_3\text{O}^+]$      $\text{pH} + \text{pOH} = 14$   
**Solution:**  $[\text{OH}^-] = [\text{Sr(OH)}_2] = 2(4.8 \times 10^{-4}) = 9.6 \times 10^{-4}$  M  

$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{9.6 \times 10^{-4} \text{ M}} = 1.04 \times 10^{-11} \text{ M}$$

$$\text{pH} = -\log(1.04 \times 10^{-11}) = 10.98$$

$$\text{pOH} = 14.00 - 10.98 = 3.02$$
- (d) **Given:**  $8.7 \times 10^{-5}$  M  $\text{KOH}$  **Find:**  $[\text{OH}^-]$ ,  $[\text{H}_3\text{O}^+]$ , pH, pOH  
**Conceptual Plan:**  $[\text{KOH}] \rightarrow [\text{OH}^-] \rightarrow [\text{H}_3\text{O}^+] \rightarrow \text{pH} \rightarrow \text{pOH}$   
 $K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$      $\text{pH} = -\log[\text{H}_3\text{O}^+]$      $\text{pH} + \text{pOH} = 14$   
**Solution:**  $[\text{OH}^-] = [\text{KOH}] = 8.7 \times 10^{-5}$  M  

$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{8.7 \times 10^{-5} \text{ M}} = 1.1 \times 10^{-10} \text{ M}$$

$$\text{pH} = -\log(1.1 \times 10^{-10}) = 9.94$$

$$\text{pOH} = 14.00 - 9.94 = 4.06$$
- 15.80 (a) **Given:**  $8.77 \times 10^{-3}$  M  $\text{LiOH}$  **Find:**  $[\text{OH}^-]$ ,  $[\text{H}_3\text{O}^+]$ , pH, pOH  
**Conceptual Plan:**  $[\text{LiOH}] \rightarrow [\text{OH}^-] \rightarrow [\text{H}_3\text{O}^+] \rightarrow \text{pH} \rightarrow \text{pOH}$   
 $K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$      $\text{pH} = -\log[\text{H}_3\text{O}^+]$      $\text{pH} + \text{pOH} = 14$   
**Solution:**  $[\text{OH}^-] = [\text{LiOH}] = 8.77 \times 10^{-3}$  M  

$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{8.77 \times 10^{-3} \text{ M}} = 1.140 \times 10^{-12} \text{ M}$$

$$\text{pH} = -\log(1.140 \times 10^{-12}) = 11.943$$

$$\text{pOH} = 14.00 - 11.943 = 2.057$$
- (b) **Given:** 0.0112 M  $\text{Ba(OH)}_2$  **Find:**  $[\text{OH}^-]$ ,  $[\text{H}_3\text{O}^+]$ , pH, pOH  
**Conceptual Plan:**  $[\text{Ba(OH)}_2] \rightarrow [\text{OH}^-] \rightarrow [\text{H}_3\text{O}^+] \rightarrow \text{pH} \rightarrow \text{pOH}$   
 $K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$      $\text{pH} = -\log[\text{H}_3\text{O}^+]$      $\text{pH} + \text{pOH} = 14$   
**Solution:**  $[\text{OH}^-] = 2[\text{Ba(OH)}_2] = 2(0.0112) = 0.0224$  M  

$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{0.0224 \text{ M}} = 4.464 \times 10^{-13} \text{ M}$$

$$\text{pH} = -\log(4.464 \times 10^{-13}) = 12.350$$

$$\text{pOH} = 14.000 - 12.350 = 1.650$$
- (c) **Given:**  $1.9 \times 10^{-4}$  M  $\text{KOH}$  **Find:**  $[\text{OH}^-]$ ,  $[\text{H}_3\text{O}^+]$ , pH, pOH  
**Conceptual Plan:**  $[\text{KOH}] \rightarrow [\text{OH}^-] \rightarrow [\text{H}_3\text{O}^+] \rightarrow \text{pH} \rightarrow \text{pOH}$   
 $K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$      $\text{pH} = -\log[\text{H}_3\text{O}^+]$      $\text{pH} + \text{pOH} = 14$   
**Solution:**  $[\text{OH}^-] = [\text{KOH}] = 1.9 \times 10^{-4}$  M  

$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{1.9 \times 10^{-4} \text{ M}} = 5.26 \times 10^{-11} \text{ M}$$

$$\text{pH} = -\log(5.26 \times 10^{-11}) = 10.28$$

$$\text{pOH} = 14.00 - 10.28 = 3.72$$

- (d) **Given:**  $5.0 \times 10^{-4}$  M  $\text{Ca}(\text{OH})_2$  **Find:**  $[\text{OH}^-]$ ,  $[\text{H}_3\text{O}^+]$ , pH, pOH  
**Conceptual Plan:**  $[\text{Ca}(\text{OH})_2] \rightarrow [\text{OH}^-] \rightarrow [\text{H}_3\text{O}^+] \rightarrow \text{pH} \rightarrow \text{pOH}$   
 $K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$     $\text{pH} = -\log[\text{H}_3\text{O}^+]$     $\text{pH} + \text{pOH} = 14$   
**Solution:**  $[\text{OH}^-] = [\text{Ca}(\text{OH})_2] = 2(5.0 \times 10^{-4}) = 0.0010$  M  
 $[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{0.0010 \text{ M}} = 1.00 \times 10^{-11}$  M  
 $\text{pH} = -\log(1.00 \times 10^{-11}) = 11.00$   
 $\text{pOH} = 14.00 - 11.00 = 3.00$

- 15.81 **Given:** 3.85% KOH by mass,  $d = 1.01$  g/mL **Find:** pH

**Conceptual Plan:**

% mass  $\rightarrow$  g KOH  $\rightarrow$  mol KOH and mass soln  $\rightarrow$  mL soln  $\rightarrow$  L soln  $\rightarrow$  M KOH  $\rightarrow$   $[\text{OH}^-]$

$\rightarrow$  pOH  $\rightarrow$  pH

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

$$\text{Solution: } \frac{3.85 \text{ g KOH}}{100.0 \text{ g soln}} \times \frac{1 \text{ mol KOH}}{56.01 \text{ g KOH}} \times \frac{1.01 \text{ g soln}}{\text{mL soln}} \times \frac{1000 \text{ mL soln}}{\text{L soln}} = 0.6942 \text{ M KOH}$$

$$[\text{OH}^-] = [\text{KOH}] = 0.6942 \text{ M} \quad \text{pOH} = -\log(0.6942) = 0.159 \quad \text{pH} = 14.000 - 0.159 = 13.841$$

- 15.82 **Given:** 1.55% NaOH by mass,  $d = 1.01$  g/mL **Find:** pH

**Conceptual Plan:**

% mass  $\rightarrow$  g NaOH  $\rightarrow$  mol NaOH and mass soln  $\rightarrow$  mL soln  $\rightarrow$  L soln  $\rightarrow$  M NaOH  $\rightarrow$   $[\text{OH}^-]$

$\rightarrow$  pOH  $\rightarrow$  pH

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

$$\text{Solution: } \frac{1.55 \text{ g NaOH}}{100.0 \text{ g soln}} \times \frac{1 \text{ mol NaOH}}{40.01 \text{ g NaOH}} \times \frac{1.01 \text{ g soln}}{\text{mL soln}} \times \frac{1000 \text{ mL soln}}{\text{L soln}} = 0.3913 \text{ M NaOH}$$

$$[\text{OH}^-] = [\text{NaOH}] = 0.3913 \text{ M} \quad \text{pOH} = -\log(0.3913) = 0.407 \quad \text{pH} = 14.000 - 0.407 = 13.593$$

- 15.83 **Given:** 3.55 L, pH = 12.4; 0.855 M KOH **Find:** Vol

**Conceptual Plan:** pH  $\rightarrow$   $[\text{H}_3\text{O}^+]$   $\rightarrow$   $[\text{OH}^-]$  and then  $V_1M_1 = V_2M_2$

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} \quad 1.0 \times 10^{-14} = [\text{H}_3\text{O}^+][\text{OH}^-] \quad V_1M_1 = V_2M_2$$

$$\text{Solution: } [\text{H}_3\text{O}^+] = 10^{-12.4} = 3.98 \times 10^{-13} \quad 1.0 \times 10^{-14} = 3.98 \times 10^{-13}[\text{OH}^-]$$

$$[\text{OH}^-] = 0.02512$$

$$V_1M_1 = V_2M_2 \quad V_1(0.855 \text{ M}) = (3.55 \text{ L})(0.0251 \text{ M}) \quad V_1 = 0.104 \text{ L}$$

- 15.84 **Given:** 5.00 L, pH = 10.8; 15.0% NaOH,  $d = 1.116$  g/mL **Find:** Vol

**Conceptual Plan:**

% mass  $\rightarrow$  g NaOH  $\rightarrow$  mol NaOH and mass soln  $\rightarrow$  mL soln  $\rightarrow$  L soln  $\rightarrow$  M NaOH  $\rightarrow$   $[\text{OH}^-]$

$$\frac{1 \text{ mol NaOH}}{40.01 \text{ g NaOH}} \quad \frac{1.01 \text{ g soln}}{\text{mL soln}} \quad \frac{1000 \text{ mL soln}}{\text{L soln}} \quad \frac{\text{mol NaOH}}{\text{L soln}}$$

and then:  $V_1M_1 \rightarrow V_2M_2$

$$V_1M_1 = V_2M_2$$

$$\text{Solution: } \frac{15.0 \text{ g NaOH}}{100.0 \text{ g soln}} \times \frac{1 \text{ mol NaOH}}{40.01 \text{ g NaOH}} \times \frac{1.116 \text{ g soln}}{\text{mL soln}} \times \frac{1000 \text{ mL soln}}{\text{L soln}} = 4.184 \text{ M NaOH}$$

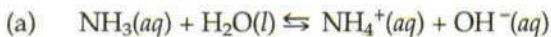
$$[\text{OH}^-] = [\text{NaOH}] = 4.184 \text{ M}$$

$$[\text{H}_3\text{O}^+] = 10^{-10.8} = 1.58 \times 10^{-11} \quad 1.0 \times 10^{-14} = (1.58 \times 10^{-11})[\text{OH}^-]$$

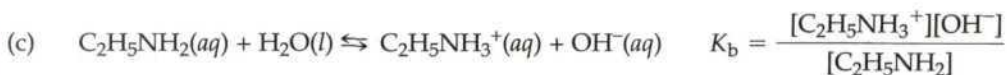
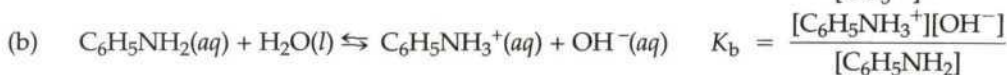
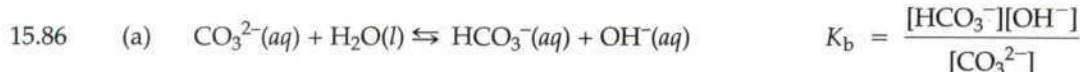
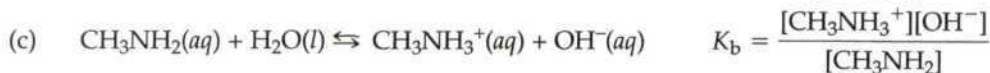
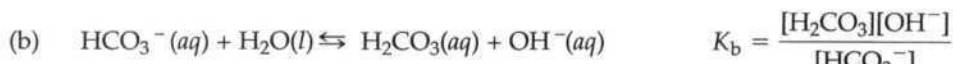
$$[\text{OH}^-] = 6.31 \times 10^{-4} \text{ M}$$

$$V_1(4.18 \text{ M}) = (5.00 \text{ L})(6.31 \times 10^{-4} \text{ M}) \quad V_1 = 7.55 \times 10^{-4} \text{ L} = 0.8 \text{ mL}$$

15.85



$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$



15.87 Given: 0.15 M  $\text{NH}_3$   $K_b = 1.76 \times 10^{-5}$  Find:  $[\text{OH}^-]$ , 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  $x$ .

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

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

Solution:  $\text{NH}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^+(aq) + \text{OH}^-(aq)$

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

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = \frac{(x)(x)}{(0.15 - x)} = 1.76 \times 10^{-5}$$

Assume  $x$  is small.

$$x^2 = (1.76 \times 10^{-5})(0.15) \quad x = [\text{OH}^-] = 0.00162 \text{ M}$$

$$\text{pOH} = -\log(0.00162) = 2.79$$

$$\text{pH} = 14.00 - 2.79 = 11.21$$

15.88 Given: 0.125 M  $\text{CO}_3^{2-}$   $K_b = 1.8 \times 10^{-4}$  Find:  $[\text{OH}^-]$ , 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  $x$ .

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

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

Solution:  $\text{CO}_3^{2-}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HCO}_3^-(aq) + \text{OH}^-(aq)$

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

$$K_b = \frac{[\text{HCO}_3^-][\text{OH}^-]}{[\text{CO}_3^{2-}]} = \frac{(x)(x)}{(0.125 - x)} = 1.8 \times 10^{-4}$$

Assume  $x$  is small.

$$x^2 = (1.8 \times 10^{-4})(0.125) \quad x = [\text{OH}^-] = 0.00474 \text{ M}$$

$$\frac{0.00474}{0.125} \times 100\% = 3.8\%; \text{ assumption is valid}$$

$$\text{pOH} = -\log(0.00474) = 2.32$$

$$\text{pH} = 14.00 - 2.32 = 11.68$$

15.89 Given:  $\text{p}K_b = 10.4$ , 455 mg/L caffeine Find: pH

Conceptual Plan:  $\text{p}K_b \rightarrow K_b$  and then  $\text{mg/L} \rightarrow \text{g/L} \rightarrow \text{mol/L}$  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  $x$ .

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

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



**Solution:**  $K_b = 10^{-10.4} = 3.98 \times 10^{-11}$

$$\frac{455 \text{ mg caffeine}}{\text{L soln}} \times \frac{\text{g caffeine}}{1000 \text{ mg caffeine}} \times \frac{1 \text{ mol caffeine}}{194.19 \text{ g}} = 0.002343 \text{ M caffeine}$$

	$\text{C}_8\text{H}_{10}\text{N}_4\text{O}_2(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HC}_8\text{H}_{10}\text{N}_4\text{O}_2^+(aq) + \text{OH}^-(aq)$		
I	0.002342	0.0	0.0
C	-x	x	x
E	0.002342 - x	x	x

$$K_b = \frac{[\text{HC}_8\text{H}_{10}\text{N}_4\text{O}_2^+][\text{OH}^-]}{[\text{C}_8\text{H}_{10}\text{N}_4\text{O}_2]} = \frac{(x)(x)}{(0.002342 - x)} = 3.98 \times 10^{-11}$$

Assume  $x$  is small.

$$x^2 = (3.98 \times 10^{-11})(0.002342) \quad x = [\text{OH}^-] = 3.05 \times 10^{-7} \text{ M}$$

$$\frac{3.05 \times 10^{-7} \text{ M}}{0.002342} \times 100\% = 0.013\%; \text{ assumption is valid}$$

$$\text{pOH} = -\log(3.05 \times 10^{-7}) = 6.5$$

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

15.90

**Given:**  $\text{p}K_b = 4.2$ , 225 mg/L amphetamine **Find:** pH

**Conceptual Plan:**  $\text{p}K_b \rightarrow K_b$  and then  $\text{mg/L} \rightarrow \text{g/L} \rightarrow \text{mol/L}$  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  $x$ .

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

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

**Solution:**  $K_b = 10^{-4.2} = 6.31 \times 10^{-5}$

$$\frac{225 \text{ mg amphetamine}}{\text{L soln}} \times \frac{\text{g amphetamine}}{1000 \text{ mg amphetamine}} \times \frac{1 \text{ mol amphetamine}}{135.21 \text{ g}} = 0.001664 \text{ M amphetamine}$$



I	0.001664 M	0.0	0.0
C	-x	x	x
E	0.001664 - x	x	x

$$K_b = \frac{[\text{C}_9\text{H}_{13}\text{NH}^+][\text{OH}^-]}{[\text{C}_9\text{H}_{13}\text{N}]} = \frac{(x)(x)}{(0.001664 - x)} = 6.31 \times 10^{-5}$$

Assume  $x$  is small.

$$x^2 = (6.31 \times 10^{-5})(0.001664) \quad x = [\text{OH}^-] = 3.24 \times 10^{-4} \text{ M}$$

$$\frac{3.24 \times 10^{-4}}{0.001664} \times 100\% = 19.4\%; \text{ assumption not valid, solve with quadratic equation.}$$

$$x^2 + 6.31 \times 10^{-5}x - 1.05 \times 10^{-7} = 0$$

$$x = [\text{OH}^-] = 2.94 \times 10^{-4} \text{ M}$$

$$\text{pOH} = -\log(2.94 \times 10^{-4}) = 3.5$$

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

15.91

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

**Conceptual Plan:**

$\text{pH} \rightarrow \text{pOH} \rightarrow [\text{OH}^-]$  and then write a balanced equation, prepare an ICE table, and determine

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

**equilibrium concentrations**  $\rightarrow K_b$ .

**Solution:**  $\text{pOH} = 14.0 - 10.5 = 3.5$   $[\text{OH}^-] = 10^{-3.5} = 3.16 \times 10^{-4} = [\text{Hmorphine}^+]$



I	0.150 M	0.0	0.0
C	-x	x	x
E	0.150 - x	$3.16 \times 10^{-4}$	$3.16 \times 10^{-4}$

$$K_b = \frac{[\text{Hmorphine}^+][\text{OH}^-]}{[\text{morphine}]} = \frac{(3.16 \times 10^{-4})(3.16 \times 10^{-4})}{(0.150 - 3.16 \times 10^{-4})} = 6.68 \times 10^{-7} = 7 \times 10^{-7}$$

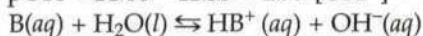
15.92

Given: 0.135 M base, pH = 11.23 Find:  $K_b$ 

Conceptual Plan:

pH  $\rightarrow$  pOH  $\rightarrow$   $[\text{OH}^-]$  and then write a balanced equation, prepare an ICE table, and determine

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

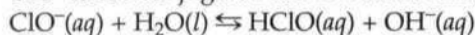
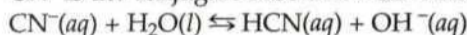
equilibrium concentrations  $\rightarrow K_b$ .Solution:  $\text{pOH} = 14.00 - 11.23 = 2.77$   $[\text{OH}^-] = 10^{-2.77} = 1.698 \times 10^{-3} = [\text{HB}^+]$ 

I	0.135	0.0	0.0
C	-x	x	x
E	0.135 - x	$1.698 \times 10^{-3}$	$1.698 \times 10^{-3}$

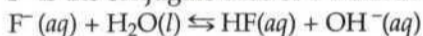
$$K_b = \frac{[\text{HB}^+][\text{OH}^-]}{[\text{B}]} = \frac{(1.698 \times 10^{-3})(1.698 \times 10^{-3})}{(0.135 - 1.698 \times 10^{-3})} = 2.16 \times 10^{-5} = 2.2 \times 10^{-5}$$

## Acid - Base Properties of Ions and Salts

15.93

(a) pH neutral:  $\text{Br}^-$  is the conjugate base of a strong acid; therefore, it is pH neutral.(b) weak base:  $\text{ClO}^-$  is the conjugate base of a weak acid; therefore, it is a weak base.(c) weak base:  $\text{CN}^-$  is the conjugate base of a weak acid; therefore, it is a weak base.(d) pH neutral:  $\text{Cl}^-$  is the conjugate base of a strong acid; therefore, it is pH neutral.

15.94

(a) weak base:  $\text{C}_7\text{H}_5\text{O}_2^-$  is the conjugate base of a weak acid; therefore, it is a weak base.(b) pH neutral:  $\text{I}^-$  is the conjugate base of a strong acid; therefore, it is pH neutral.(c) pH neutral:  $\text{NO}_3^-$  is the conjugate base of a strong acid; therefore, it is pH neutral.(d) weak base:  $\text{F}^-$  is the conjugate base of a weak acid; therefore, it is a weak base.

15.95

Given:  $[\text{F}^-] = 0.140 \text{ M}$ ,  $K_a(\text{HF}) = 3.5 \times 10^{-4}$  Find:  $[\text{OH}^-]$ , pOHConceptual 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  $[\text{OH}^-] \rightarrow \text{pOH} \rightarrow \text{pH}$ .

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

Solution:  $\text{F}^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HF}(aq) + \text{OH}^-(aq)$ 

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.

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

$$\text{pH} = 14.00 - 5.70 = 8.30$$

15.96

Given:  $[\text{HCO}_3^-] = 0.250 \text{ M}$ ,  $K_a(\text{H}_2\text{CO}_3) = 4.3 \times 10^{-7}$  Find:  $[\text{OH}^-]$ , pOHConceptual 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  $[\text{OH}^-] \rightarrow \text{pOH} \rightarrow \text{pH}$ .

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

**Solution:**  $\text{HCO}_3^- (aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_2\text{CO}_3 (aq) + \text{OH}^- (aq)$

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.

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

$$\text{pH} = 14.00 - 4.12 = 9.88$$

15.97

- (a) weak acid:  $\text{NH}_4^+$  is the conjugate acid of a weak base; therefore, it is a weak acid.  
 $\text{NH}_4^+(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{NH}_3(aq)$
- (b) pH neutral:  $\text{Na}^+$  is the counterion of a strong base; therefore, it is pH neutral.
- (c) weak acid: The  $\text{Co}^{3+}$  cation is a small, highly charged metal cation; therefore, it is a weak acid.  
 $\text{Co}(\text{H}_2\text{O})_6^{3+}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{Co}(\text{H}_2\text{O})_5(\text{OH})^{2+}(aq) + \text{H}_3\text{O}^+(aq)$
- (d) weak acid:  $\text{CH}_2\text{NH}_3^+$  is the conjugate acid of a weak base; therefore, it is a weak acid.  
 $\text{CH}_2\text{NH}_3^+(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{CH}_2\text{NH}_2(aq)$

15.98

- (a) pH neutral:  $\text{Sr}^{2+}$  is the counterion of a strong base; therefore, it is pH neutral.
- (b) weak acid: The  $\text{Mn}^{3+}$  cation is a small, highly charged metal cation; therefore, it is a weak acid.  
 $\text{Mn}(\text{H}_2\text{O})_6^{3+}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{Mn}(\text{H}_2\text{O})_5(\text{OH})^{2+}(aq) + \text{H}_3\text{O}^+(aq)$
- (c) weak acid:  $\text{C}_5\text{H}_5\text{NH}_3^+$  is the conjugate acid of a weak base; therefore, it is a weak acid.  
 $\text{C}_5\text{H}_5\text{NH}_3^+(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{C}_5\text{H}_5\text{NH}_2(aq)$

15.99

- (d) pH neutral:  $\text{Li}^+$  is the counterion of a strong base; therefore, it is pH neutral.
- (a) acidic:  $\text{FeCl}_3$   $\text{Fe}^{3+}$  is a small, highly charged metal cation and is therefore acidic.  $\text{Cl}^-$  is the conjugate base of a strong acid; therefore, it is pH neutral.
- (b) basic:  $\text{NaF}$   $\text{Na}^+$  is the counterion of a strong base; therefore, it is pH neutral.  $\text{F}^-$  is the conjugate base of a weak acid; therefore, it is basic.
- (c) pH neutral:  $\text{CaBr}_2$   $\text{Ca}^{2+}$  is the counterion of a strong base; therefore, it is pH neutral.  $\text{Br}^-$  is the conjugate base of a strong acid; therefore, it is pH neutral.
- (d) acidic:  $\text{NH}_4\text{Br}$   $\text{NH}_4^+$  is the conjugate acid of a weak base; therefore, it is acidic.  $\text{Br}^-$  is the conjugate base of a strong acid; therefore, it is pH neutral.
- (e) acidic:  $\text{C}_6\text{H}_5\text{NH}_3\text{NO}_2$   $\text{C}_6\text{H}_5\text{NH}_3^+$  is the conjugate acid of a weak base; therefore, it is a weak acid.  $\text{NO}_2^-$  is the conjugate base of a weak acid; therefore, it is a weak base. To determine pH, compare  $K$  values.

$$K_a(\text{C}_6\text{H}_5\text{NH}_3^+) = \frac{1.0 \times 10^{-14}}{3.9 \times 10^{-10}} = 2.6 \times 10^{-5} \quad K_b(\text{NO}_2^-) = \frac{1.0 \times 10^{-14}}{4.6 \times 10^{-4}} = 2.2 \times 10^{-11}$$

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

15.100

- (a) acidic:  $\text{Al}(\text{NO}_3)_3$   $\text{Al}^{3+}$  is a small, highly charged metal cation and is therefore acidic.  $\text{NO}_3^-$  is the conjugate base of a strong acid; therefore, it is pH neutral.
- (b) acidic:  $\text{C}_2\text{H}_5\text{NH}_3\text{NO}_3$   $\text{C}_2\text{H}_5\text{NH}_3^+$  is the conjugate acid of a weak base; therefore, it is acidic.  $\text{NO}_3^-$  is the conjugate base of a strong acid; therefore, it is pH neutral.
- (c) basic:  $\text{K}_2\text{CO}_3$   $\text{K}^+$  is the counterion of a strong base; therefore, it is pH neutral.  $\text{CO}_3^{2-}$  is the conjugate base of a weak acid; therefore, it is basic.



- (d) pH neutral:  $\text{RbI}$   $\text{Rb}^+$  is the counterion of a strong base; therefore, it is pH neutral.  $\text{I}^-$  is the conjugate base of a strong acid; therefore, it is pH neutral.
- (e) basic  $\text{NH}_4\text{ClO}$   $\text{NH}_4^+$  is the conjugate acid of a weak base; therefore, it is a weak acid.  $\text{ClO}^-$  is the conjugate base of a weak acid; therefore, it is a weak base. To determine pH, compare  $K$  values.
- $$K_a(\text{NH}_4^+) = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10} \quad K_b(\text{ClO}^-) = \frac{1.0 \times 10^{-14}}{2.9 \times 10^{-8}} = 3.4 \times 10^{-7}$$
- $K_b > K_a$ ; therefore, the solution is basic.

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

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

$\text{NH}_4\text{Cl}$  acidic:  $\text{NH}_4^+$  is the conjugate acid of a weak base; therefore, it is acidic.  $\text{Cl}^-$  is the conjugate base of a strong acid; therefore, it is pH neutral.

$\text{NaHCO}_3$  basic:  $\text{Na}^+$  is the counterion of a strong base; therefore, it is pH neutral.  $\text{HCO}_3^-$  is the conjugate base of a weak acid; therefore, it is basic.

$\text{NH}_4\text{ClO}_2$  acidic:  $\text{NH}_4\text{ClO}_2$   $\text{NH}_4^+$  is the conjugate acid of a weak base; therefore, it is a weak acid.  $\text{ClO}_2^-$  is the conjugate base of a weak acid; therefore, it is a weak base.  $K_a(\text{NH}_4^+) = 5.6 \times 10^{-10}$   $K_b(\text{ClO}_2^-) = 9.1 \times 10^{-13}$

$\text{NaOH}$  strong base

Increasing acidity:  $\text{NaOH} < \text{NaHCO}_3 < \text{NaCl} < \text{NH}_4\text{ClO}_2 < \text{NH}_4\text{Cl}$

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

$\text{CH}_3\text{NH}_3\text{Br}$  acidic:  $\text{CH}_3\text{NH}_3^+$  is the conjugate acid of a weak base; therefore, it is acidic.  $\text{Br}^-$  is the conjugate base of a strong acid; therefore, it is pH neutral.

$\text{KOH}$  strong base

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

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

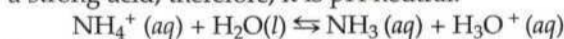
$\text{C}_5\text{H}_5\text{NHNO}_2$  acidic:  $\text{C}_5\text{H}_5\text{NH}^+$  is the conjugate acid of a weak base; therefore, it is acidic.  $\text{NO}_2^-$  is the conjugate base of a weak acid; therefore, it is basic.  $K_a(\text{C}_5\text{H}_5\text{NH}^+) = 5.9 \times 10^{-6}$ ;  $K_b(\text{NO}_2^-) = 2.2 \times 10^{-11}$

Increasing basicity:  $\text{CH}_3\text{NH}_3\text{Br} < \text{C}_5\text{H}_5\text{NHNO}_2 < \text{KBr} < \text{KCN} < \text{KOH}$

15.103 (a) Given: 0.10 M  $\text{NH}_4\text{Cl}$  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  $[\text{H}_3\text{O}^+] \rightarrow \text{pH}$ .

**Solution:**  $\text{NH}_4^+$  is the conjugate acid of a weak base; therefore, it is acidic.  $\text{Cl}^-$  is the conjugate base of a strong acid; therefore, it is pH neutral.



I 0.10 0.0 0.0

C  $-x$   $x$   $x$

E  $0.10 - x$   $x$   $x$

$$K_a = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.56 \times 10^{-10} = \frac{(x)(x)}{(0.10 - x)}$$

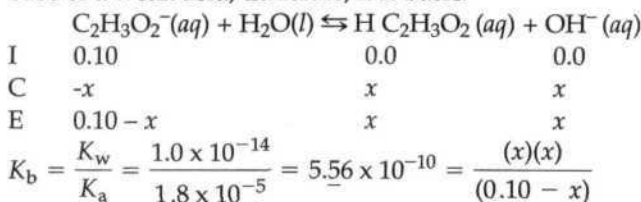
Assume  $x$  is small.

$$x = 7.45 \times 10^{-6} = [\text{H}_3\text{O}^+] \quad \text{pH} = -\log(7.45 \times 10^{-6}) = 5.13$$

- (b) Given: 0.10 M  $\text{NaC}_2\text{H}_3\text{O}_2$  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  $[\text{OH}^-] \rightarrow \text{pOH} \rightarrow \text{pH}$ .

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

**Solution:**  $\text{Na}^+$  is the counterion of a strong base; therefore, it is pH neutral.  $\text{C}_2\text{H}_3\text{O}_2^-$  is the conjugate base of a weak acid; therefore, it is basic.



Assume  $x$  is small.

$$x = 7.45 \times 10^{-6} = [\text{OH}^-] \quad \text{pOH} = -\log(7.45 \times 10^{-6}) = 5.13$$

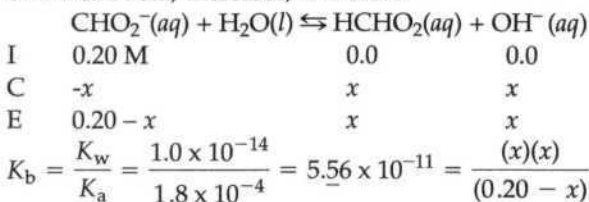
$$\text{pH} = 14.00 - 5.13 = 8.87$$

- (c) Given: 0.10 M  $\text{NaCl}$  Find: pH  
**Conceptual Plan:** Identify each species and determine which will contribute to pH.  
**Solution:**  $\text{Na}^+$  is the counterion of a strong base; therefore, it is pH neutral.  $\text{Cl}^-$  is the conjugate base of a strong acid; therefore, it is pH neutral.  
 pH = 7.0

- 15.104 (a) Given: 0.20 M  $\text{NaCHO}_2$  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  $[\text{OH}^-] \rightarrow \text{pOH} \rightarrow \text{pH}$ .

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

**Solution:**  $\text{Na}^+$  is the counterion of a strong base; therefore, it is pH neutral.  $\text{CHO}_2^-$  is the conjugate base of a weak acid; therefore, it is basic.



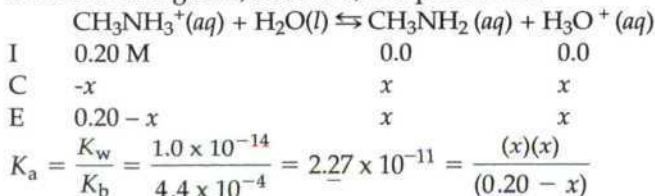
Assume  $x$  is small.

$$x = 3.33 \times 10^{-6} = [\text{OH}^-] \quad \text{pOH} = -\log(3.33 \times 10^{-6}) = 5.48$$

$$\text{pH} = 14.00 - 5.48 = 8.52$$

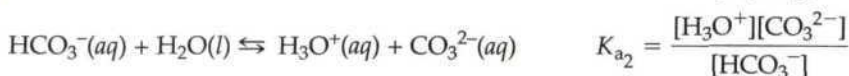
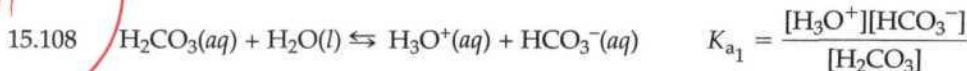
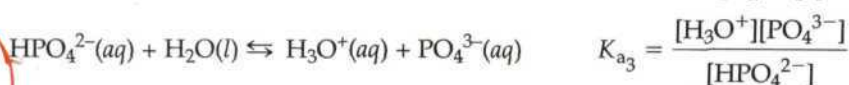
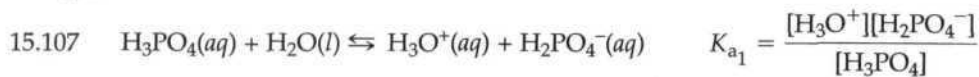
- (b) Given: 0.20 M  $\text{CH}_3\text{NH}_3\text{I}$  Find: pH  
**Conceptual Plan:** Identify each species and determine which will contribute to pH. Write a balanced reaction. Prepare an ICE table, represent the change with  $x$ , sum the table, determine the equilibrium values, put the equilibrium values in the equilibrium expression, and solve for  $x$ . Determine  $[\text{H}_3\text{O}^+] \rightarrow \text{pH}$ .

**Solution:**  $\text{CH}_3\text{NH}_3^+$  is the conjugate acid of a weak base; therefore, it is acidic.  $\text{I}^-$  is the conjugate base of a strong acid; therefore, it is pH neutral.

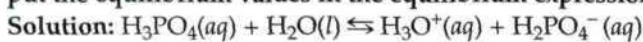




## Polyprotic Acids



- 15.109 (a) **Given:** 0.350 M  $\text{H}_3\text{PO}_4$   $K_{a1} = 7.5 \times 10^{-3}$ ,  $K_{a2} = 6.2 \times 10^{-8}$  **Find:**  $[\text{H}_3\text{O}^+]$ , pH  
**Conceptual Plan:**  $K_{a1}$  is much larger than  $K_{a2}$ , so use  $K_{a1}$  to calculate  $[\text{H}_3\text{O}^+]$ . 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  $x$ .



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

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{H}_2\text{PO}_4^-]}{[\text{H}_3\text{PO}_4]} = \frac{(x)(x)}{(0.350 - x)} = 7.3 \times 10^{-3}$$

Assume  $x$  is small compared to 0.350.

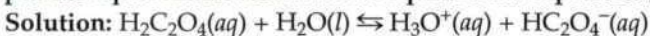
$$x^2 = (7.3 \times 10^{-3})(0.350) \quad x = 0.0505 \text{ M} = [\text{H}_3\text{O}^+]$$

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

$$x^2 + 7.5 \times 10^{-3}x - 0.002625 = 0 \quad x = 0.04762 = [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log(0.04762) = 1.32$$

- (b) **Given:** 0.350 M  $\text{H}_2\text{C}_2\text{O}_4$   $K_{a1} = 6.0 \times 10^{-2}$ ,  $K_{a2} = 6.0 \times 10^{-5}$  **Find:**  $[\text{H}_3\text{O}^+]$ , pH  
**Conceptual Plan:**  $K_{a1}$  is much larger than  $K_{a2}$ , so use  $K_{a1}$  to calculate  $[\text{H}_3\text{O}^+]$ . 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  $x$ .



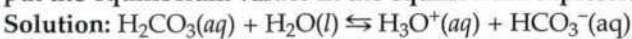
I	0.350 M	0.0	0.0
C	-x	x	x
E	0.350 - x	x	x

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{HC}_2\text{O}_4^-]}{[\text{H}_2\text{C}_2\text{O}_4]} = \frac{(x)(x)}{(0.350 - x)} = 6.0 \times 10^{-2}$$

$$x^2 + 6.0 \times 10^{-2}x - 0.021 = 0 \quad x = 0.1179 = 0.12 \text{ M } [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log(0.1179) = 0.93$$

- 15.110 (a) **Given:** 0.125 M  $\text{H}_2\text{CO}_3$   $K_{a1} = 4.3 \times 10^{-7}$ ,  $K_{a2} = 5.6 \times 10^{-11}$  **Find:**  $[\text{H}_3\text{O}^+]$ , pH  
**Conceptual Plan:**  $K_{a1}$  much larger than  $K_{a2}$ , so use  $K_{a1}$  to calculate  $[\text{H}_3\text{O}^+]$ . 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  $x$ .



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



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]} = \frac{(x)(x)}{(0.125 - x)} = 4.3 \times 10^{-7}$$

Assume  $x$  is small.

$$x^2 = (4.3 \times 10^{-7})(0.125) \quad x = 2.32 \times 10^{-4} = [\text{H}_3\text{O}^+]$$

$$\frac{2.32 \times 10^{-4}}{0.125} \times 100\% = 0.19\%; \text{ assumption valid}$$

$$\text{pH} = -\log(2.32 \times 10^{-4}) = 3.63$$

- (b) **Given:** 0.125 M  $\text{H}_3\text{C}_6\text{H}_5\text{O}_3$ ,  $K_{a1} = 7.4 \times 10^{-4}$ ,  $K_{a2} = 1.7 \times 10^{-5}$ ,  $K_{a3} = 4.0 \times 10^{-7}$  **Find:**  $[\text{H}_3\text{O}^+]$ , pH  
**Conceptual Plan:**  $K_{a1}$  and  $K_{a2}$  are only  $10^{-1}$  apart, so use both to calculate  $[\text{H}_3\text{O}^+]$ . 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:**  $\text{H}_3\text{C}_6\text{H}_5\text{O}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{H}_2\text{C}_6\text{H}_5\text{O}_3^-(aq)$

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

$$K_{a1} = \frac{[\text{H}_3\text{O}^+][\text{H}_2\text{C}_6\text{H}_5\text{O}_3^-]}{[\text{H}_3\text{C}_6\text{H}_5\text{O}_3]} = \frac{(x)(x)}{(0.125 - x)} = 7.4 \times 10^{-4}$$

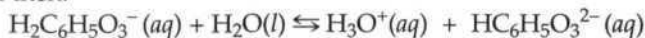
Assume  $x$  is small.

$$x^2 = (7.4 \times 10^{-4})(0.125) \quad x = 9.62 \times 10^{-3} = [\text{H}_3\text{O}^+]$$

$$\frac{9.62 \times 10^{-3}}{0.125} \times 100\% = 7.8\%; \text{ assumption not valid, solve with quadratic equation.}$$

$$x^2 + 7.4 \times 10^{-4}x - 9.25 \times 10^{-5} = 0 \quad x = 0.009255 = [\text{H}_3\text{O}^+] = [\text{H}_2\text{C}_6\text{H}_5\text{O}_3^-]$$

and then:



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

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{HC}_6\text{H}_5\text{O}_3^{2-}]}{[\text{H}_2\text{C}_6\text{H}_5\text{O}_3^-]} = \frac{(0.009255 + y)(y)}{(0.009255 - y)} = 1.7 \times 10^{-5}$$

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

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

$$[\text{H}_3\text{O}^+] = 1.7 \times 10^{-5} \text{ (from second ionization)}$$

$$[\text{H}_3\text{O}^+] = 0.009255 + 1.7 \times 10^{-5} = 0.00927 \text{ M} \quad \text{pH} = -\log(0.00927) = 2.03$$

15.111

**Given:** 0.500 M  $\text{H}_2\text{SO}_3$ ,  $K_{a1} = 1.6 \times 10^{-2}$ ,  $K_{a2} = 6.4 \times 10^{-8}$  **Find:** concentration all species

**Conceptual Plan:**  $K_{a1}$  is much larger than  $K_{a2}$  so, use  $K_{a1}$  to calculate  $[\text{H}_3\text{O}^+]$ . 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  $x$ .

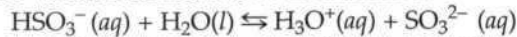
**Solution:**  $\text{H}_2\text{SO}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{HSO}_3^-(aq)$

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

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{HSO}_3^-]}{[\text{H}_2\text{SO}_3]} = \frac{(x)(x)}{(0.500 - x)} = 1.6 \times 10^{-2}$$

$$x^2 + 1.6 \times 10^{-2}x - 0.0080 = 0 \quad x = 0.0818 = 0.082 \text{ M } [\text{H}_3\text{O}^+] = [\text{HSO}_3^-]$$

Use the values from reaction 1 in reaction 2.



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

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{SO}_3^{2-}]}{[\text{HSO}_3^-]} = \frac{(0.0818 + y)(y)}{(0.0818 - y)} = 6.4 \times 10^{-8}$$

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

$$[\text{H}_2\text{SO}_3] = 0.500 - 0.0818 = 0.418 \text{ M}$$

$$[\text{HSO}_3^-] = x = 0.0818 = 0.082 \text{ M}$$

$$[\text{SO}_3^{2-}] = y = 6.4 \times 10^{-8} \text{ M}$$

$$[\text{H}_3\text{O}^+] = x + y = 0.0818 = 0.082 \text{ M}$$

$$[\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]} = \frac{1.0 \times 10^{-14}}{0.0818} = 1.2 \times 10^{-13} \text{ M}$$

- 15.112 **Given:** 0.155 M  $\text{H}_2\text{CO}_3$   $K_{a_1} = 4.3 \times 10^{-7}$ ,  $K_{a_2} = 5.6 \times 10^{-11}$  **Find:** concentration all species  
**Conceptual Plan:**  $K_{a_1}$  is much larger than  $K_{a_2}$  so, use  $K_{a_1}$  to calculate  $[\text{H}_3\text{O}^+]$  and  $[\text{HCO}_3^-]$ . Use  $K_{a_2}$  to find  $[\text{CO}_3^{2-}]$ . 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:**  $\text{H}_2\text{CO}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{HCO}_3^-(\text{aq})$

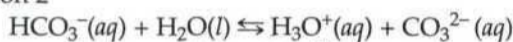
I	0.125 M	0.0	0.0
C	-x	x	x
E	0.125 - x	x	x

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]} = \frac{(x)(x)}{(0.155 - x)} = 4.3 \times 10^{-7}$$

Assume  $x$  is small.

$$x^2 = (4.3 \times 10^{-7})(0.155) \quad x = 2.58 \times 10^{-4} \text{ M} = [\text{H}_3\text{O}^+] = [\text{HCO}_3^-]$$

Reaction 2



Since  $K_{a_2}$  is small,  $y = [\text{CO}_3^{2-}] = K_{a_2}$

$$[\text{H}_2\text{CO}_3] = 0.155 - 2.58 \times 10^{-4} = 0.1547 = 0.155 \text{ M}$$

$$[\text{HCO}_3^-] = x = 2.58 \times 10^{-4} = 2.6 \times 10^{-4} \text{ M}$$

$$[\text{CO}_3^{2-}] = y = 5.6 \times 10^{-11} \text{ M}$$

$$[\text{H}_3\text{O}^+] = x = 2.58 \times 10^{-4} = 2.6 \times 10^{-4} \text{ M}$$

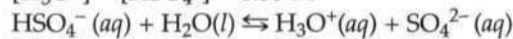
$$[\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]} = \frac{1.0 \times 10^{-14}}{2.58 \times 10^{-4}} = 3.88 \times 10^{-11} \text{ M} = 3.9 \times 10^{-11} \text{ M}$$

- 15.113 (a) **Given:**  $[\text{H}_2\text{SO}_4] = 0.50 \text{ M}$   $K_{a_2} = 0.012$  **Find:**  $[\text{H}_3\text{O}^+]$ , 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:**  $\text{H}_2\text{SO}_4(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{HSO}_4^-(\text{aq})$  strong

$$0.50 \text{ M}$$

$$[\text{H}_3\text{O}^+] = [\text{HSO}_4^-] = 0.50 \text{ M}$$



I	0.50 M	0.50	0.0
C	-x	x	x
E	0.500 - x	0.50 + x	x

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{SO}_4^{2-}]}{[\text{HSO}_4^-]} = \frac{(0.50 + x)(x)}{(0.50 - x)} = 0.012$$

$$x^2 + 0.512x - 0.006 = 0 \quad x = 0.0115 = [\text{H}_3\text{O}^+] \text{ from second ionization step}$$

$$[\text{H}_3\text{O}^+] = 0.50 + 0.012 = 0.51 \text{ M}$$

$$\text{pH} = -\log(0.51) = 0.29$$

- (b) **Given:**  $[\text{H}_2\text{SO}_4] = 0.10 \text{ M}$   $K_{a_2} = 0.012$  **Find:**  $[\text{H}_3\text{O}^+]$ , 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  $x$ .