

ICE Table for the ionization of $\text{H}_2\text{S}_{(\text{aq})}$			
	$\text{H}_2\text{S}_{(\text{aq})} \rightleftharpoons$	$\text{H}^+_{(\text{aq})} +$	$\text{HS}^-_{(\text{aq})}$
Initial concentration (mol/L)	0.010	0.00	0.00
Change in concentration (mol/L)	-x	+x	+x
Equilibrium concentration (mol/L)	$0.010 - x$	x	x

$$K_a = \frac{[\text{H}^+_{(\text{aq})}][\text{HS}^-_{(\text{aq})}]}{[\text{H}_2\text{S}_{(\text{aq})}]}$$

$$\frac{x^2}{0.010 - x} = 1.1 \times 10^{-7}$$

Predicting whether  $0.010 - x \doteq 0.010$  ...

$$\frac{[\text{HA}]_{\text{initial}}}{K_a} = \frac{0.010 \text{ mol/L}}{1.1 \times 10^{-7}}$$

$$\frac{[\text{HA}]_{\text{initial}}}{K_a} = 9.0 \times 10^5$$

Since  $9.0 \times 10^5 > 100$ , we can assume that  $0.010 - x \doteq 0.010$ .

$$\frac{x^2}{(0.010)} \doteq 1.0 \times 10^{-2}$$

$$x^2 \doteq 1.0 \times 10^{-9}$$

$$x = 3.3 \times 10^{-5}$$

Validating the assumption ...

$$\frac{3.3 \times 10^{-5}}{0.010} \times 100\% = 0.33\%$$

Since  $0.33\% < 5.0\%$ , the assumption is valid.

$$[\text{H}^+_{(\text{aq})}] = 3.3 \times 10^{-5} \text{ mol/L}$$

$$\text{pH} = -\log[\text{H}^+_{(\text{aq})}]$$

$$= -\log[3.3 \times 10^{-5}]$$

$$\text{pH} = 4.48$$

The pH of a 0.010 mol/L solution of hydrosulfuric acid is 4.48.

## SECTION 8.2 QUESTIONS

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### Understanding Concepts

- A dilute solution of a strong acid and a solution of a weak acid may have similar pH values but differ in the degree of ionization that occurs. Strong acids are fully ionized while weak acids are only partially ionized.
- $\text{CN}^-_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \rightleftharpoons \text{HCN}_{(\text{aq})} + \text{OH}^-_{(\text{aq})}$

$$K_b = \frac{[\text{CN}^-_{(\text{aq})}][\text{OH}^-_{(\text{aq})}]}{[\text{HCN}_{(\text{aq})}]}$$

ICE Table for the Ionization of $\text{CN}^-_{(\text{aq})}$				
	$\text{CN}^-_{(\text{aq})} +$	$\text{H}_2\text{O}_{(\text{l})} \rightleftharpoons$	$\text{HCN}_{(\text{aq})} +$	$\text{OH}^-_{(\text{aq})}$
Initial concentration (mol/L)	0.18	—	0	0
Change in concentration (mol/L)	-x	—	+x	+x
Equilibrium concentration (mol/L)	$0.18 - x$	—	x	x

$$K_b = \frac{1.0 \times 10^{-14}}{K_a}$$

$$= \frac{1.0 \times 10^{-14}}{6.2 \times 10^{-10}}$$

$$K_b = 1.6 \times 10^{-5}$$

$$\frac{x^2}{0.18 - x} = 1.6 \times 10^{-5}$$

Predicting whether  $0.18 - x \approx 0.18$  ...

$$\frac{0.18}{1.6 \times 10^{-5}} = 1.1 \times 10^4$$

Since  $1.1 \times 10^4 > 100$ , we assume that  $0.18 - x \approx 0.18$ .

$$\frac{x^2}{(0.18)} \approx 1.6 \times 10^{-5}$$

$$x^2 \approx 2.9 \times 10^{-6}$$

$$x = 1.7 \times 10^{-3}$$

Validating the assumption ...

$$\frac{1.7 \times 10^{-3}}{0.18} \times 100\% = 0.94\%$$

Since  $0.94\% < 5.0\%$ , the assumption is valid.

$$[\text{OH}^-_{(\text{aq})}] = 1.7 \times 10^{-3} \text{ mol/L}$$

$$\text{pOH} = -\log[\text{OH}^-_{(\text{aq})}]$$

$$= -\log[1.7 \times 10^{-3}]$$

$$\text{pOH} = 2.77$$

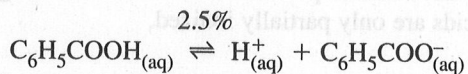
$$\text{pH} = 14 - \text{pOH}$$

$$= 14 - 2.77$$

$$\text{pH} = 11.23$$

The pH of a 0.18 mol/L cyanide solution is 11.23.

4.



$$K_a = \frac{[\text{H}^+_{(\text{aq})}][\text{C}_6\text{H}_5\text{COO}^-_{(\text{aq})}]}{[\text{C}_6\text{H}_5\text{COOH}_{(\text{aq})}]}$$

$$x = 0.100 \text{ mol/L} \times 0.0025$$

$$x = 2.5 \times 10^{-3}$$

ICE Table for the Ionization of $C_6H_5COOH_{(aq)}$			
	$C_6H_5COOH_{(aq)} \rightleftharpoons$	$H^+_{(aq)} +$	$C_6H_5COO^-_{(aq)}$
Initial concentration (mol/L)	0.100	0.00	0.00
Change in concentration (mol/L)	$-2.5 \times 10^{-3}$	$+2.5 \times 10^{-3}$	$+2.5 \times 10^{-3}$
Equilibrium concentration (mol/L)	0.098	$2.5 \times 10^{-3}$	$2.5 \times 10^{-3}$

$$K_a = \frac{[H^+_{(aq)}][C_6H_5COO^-_{(aq)}]}{[C_6H_5COOH_{(aq)}]}$$

$$= \frac{(2.5 \times 10^{-3})^2}{0.098}$$

$$K_a = 6.4 \times 10^{-5}$$

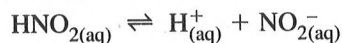
The  $K_a$  of benzoic acid is  $6.4 \times 10^{-5}$ .

5.  $pH = 2.1$

$$[H^+_{(aq)}] = 10^{-pH}$$

$$= 10^{-2.1}$$

$$[H^+_{(aq)}] = 8 \times 10^{-3} \text{ mol/L}$$



$$K_a = \frac{[H^+_{(aq)}][NO_{2(aq)}^-]}{[HNO_{2(aq)}]}$$

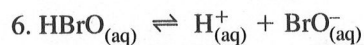
ICE Table for the Ionization of $HNO_{2(aq)}$			
	$HNO_{2(aq)} \rightleftharpoons$	$H^+_{(aq)} +$	$NO_{2(aq)}^-$
Initial concentration (mol/L)	0.100	0	0
Change in concentration (mol/L)	$-8 \times 10^{-3}$	$+8 \times 10^{-3}$	$+8 \times 10^{-3}$
Equilibrium concentration (mol/L)	0.09	$8 \times 10^{-3}$	$8 \times 10^{-3}$

$$K_a = \frac{[H^+_{(aq)}][NO_{2(aq)}^-]}{[HNO_{2(aq)}]}$$

$$= \frac{(8 \times 10^{-3})^2}{0.09}$$

$$K_a = 7 \times 10^{-4}$$

The  $K_a$  for nitrous acid is  $7 \times 10^{-4}$ .



$$K_a = \frac{[H^+_{(aq)}][BrO^-_{(aq)}]}{[HBrO_{(aq)}]}$$

$$K_a = 2.5 \times 10^{-9}$$

Since  $2.5 \times 10^{-9} > 100$ , we can assume that  $0.200 - x \doteq 0.200$ .

Predict whether a simplifying assumption is justified:

$$\frac{[\text{HA}]_{\text{initial}}}{K_a} = \frac{0.200}{2.5 \times 10^{-9}}$$

$$\frac{[\text{HA}]_{\text{initial}}}{K_a} = 1.3 \times 10^9$$

The equilibrium expression becomes

$$\frac{x^2}{0.200} \doteq 2.5 \times 10^{-9}$$

which yields

$$x^2 \doteq 5.0 \times 10^{-10}$$

$$x \doteq 2.2 \times 10^{-5}$$

Validation of the approximation:

In general, the approximation is valid if

$$\frac{x}{[\text{HA}]_{\text{initial}}} \times 100\% \leq 5\%$$

$$[\text{HA}]_{\text{initial}} = 0.200 \text{ mol/L}$$

$$\frac{x}{[\text{HA}]_{\text{initial}}} \times 100\% = \frac{2.2 \times 10^{-5}}{0.200} \times 100\%$$

$$\frac{x}{[\text{HA}]_{\text{initial}}} \times 100\% = 0.011\%$$

Since  $0.011\% < 5\%$ , the assumption is valid.

Therefore,

$$[\text{H}^+_{(\text{aq})}] = 2.2 \times 10^{-5} \text{ mol/L}$$

$$\text{pH} = -\log 2.2 \times 10^{-5}$$

$$= 2.28$$

$$\text{pH} = 4.65$$

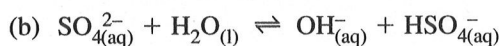
The pH of the hypobromous acid solution is 4.65.

7. (a)  $\text{HCN}_{(\text{aq})}$ ,  $\text{HCO}_2\text{H}_{(\text{aq})}$ ,  $\text{HF}_{(\text{aq})}$ ,  $\text{HNO}_3_{(\text{aq})}$   
 (b) Approximate pH of 1.0 mol/L solutions:

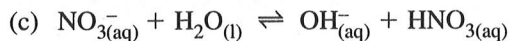
Acid	Approximate pH
$\text{HCN}_{(\text{aq})}$	5
$\text{HCO}_2\text{H}_{(\text{aq})}$	4
$\text{HF}_{(\text{aq})}$	2
$\text{HNO}_3_{(\text{aq})}$	0

8. (a) (strongest acid)  $\text{H}_3\text{PO}_4_{(\text{aq})}$ ,  $\text{HNO}_2_{(\text{aq})}$ ,  $\text{HC}_2\text{H}_3\text{O}_2_{(\text{aq})}$ ,  $\text{H}_2\text{S}_{(\text{aq})}$ ,  $\text{NH}_4^+_{(\text{aq})}$  (weakest acid)  
 (b) (highest pH)  $\text{NH}_4^+_{(\text{aq})}$ ,  $\text{H}_2\text{S}_{(\text{aq})}$ ,  $\text{HC}_2\text{H}_3\text{O}_2_{(\text{aq})}$ ,  $\text{HNO}_2_{(\text{aq})}$ ,  $\text{H}_3\text{PO}_4_{(\text{aq})}$  (lowest pH)  
 (c) (strongest base)  $\text{NH}_3_{(\text{aq})}$ ,  $\text{HS}^-_{(\text{aq})}$ ,  $\text{C}_2\text{H}_3\text{O}_2^-_{(\text{aq})}$ ,  $\text{NO}_2^-_{(\text{aq})}$ ,  $\text{H}_2\text{PO}_4^-_{(\text{aq})}$  (weakest base)
9. (a)  $\text{CN}^-_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \rightleftharpoons \text{OH}^-_{(\text{aq})} + \text{HCN}_{(\text{aq})}$

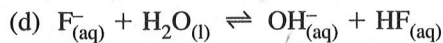
$$K_b = \frac{[\text{HCN}_{(\text{aq})}][\text{OH}^-_{(\text{aq})}]}{[\text{CN}^-_{(\text{aq})}]}$$



$$K_b = \frac{[\text{HSO}_{4(aq)}^-][\text{OH}_{(aq)}^-]}{[\text{SO}_{4(aq)}^{2-}]}$$



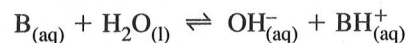
$$K_b = \frac{[\text{HNO}_{3(aq)}][\text{OH}_{(aq)}^-]}{[\text{NO}_{3(aq)}^-]}$$



$$K_b = \frac{[\text{HF}_{(aq)}][\text{OH}_{(aq)}^-]}{[\text{F}_{(aq)}^-]}$$

10. (a) (weakest base) morphine, erythromycin, atropine (strongest base)

(b) **pH of morphine ...**



$$K_b = \frac{[\text{BH}_{(aq)}^+][\text{OH}_{(aq)}^-]}{[\text{B}_{(aq)}]}$$

$$K_b = 7.9 \times 10^{-7}$$

ICE Table for the Ionization of the Base Atropine $\text{B}_{(aq)}$				
	$\text{B}_{(aq)}$	$\text{H}_2\text{O}_{(l)} \rightleftharpoons$	$\text{OH}_{(aq)}^- +$	$\text{BH}_{(aq)}^+$
Initial concentration (mol/L)	0.1	–	0	0
Change in concentration (mol/L)	–x	–	+x	+x
Equilibrium concentration (mol/L)	0.1 – x	–	x	x

Since  $0.3\% < 5\%$ , the assumption is valid.

Therefore,

$$[\text{OH}_{(aq)}^-] = 2.8 \times 10^{-4} \text{ mol/L}$$

$$\text{pOH} = -\log 2.8 \times 10^{-4}$$

$$\text{pOH} = 3.6$$

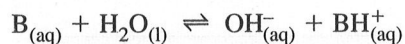
$$\text{pH} = 14 - \text{pOH}$$

$$= 14.0 - 3.6$$

$$\text{pH} = 10.4$$

The pH of a 0.1 mol/L solution of morphine is 10.4.

**pH of erythromycin ...**



$$K_b = \frac{[\text{BH}_{(aq)}^+][\text{OH}_{(aq)}^-]}{[\text{B}_{(aq)}]}$$

$$K_b = 6.3 \times 10^{-6}$$

ICE Table for the Ionization of the Base Erythromycin $\text{B}_{(aq)}$				
	$\text{B}_{(aq)}$	$\text{H}_2\text{O}_{(l)} \rightleftharpoons$	$\text{OH}_{(aq)}^- +$	$\text{BH}_{(aq)}^+$
Initial concentration (mol/L)	0.1	–	0.0	0.0
Change in concentration (mol/L)	–x	–	+x	+x
Equilibrium concentration (mol/L)	0.1 – x	–	x	x

Since  $0.8\% < 5\%$ , the assumption is valid.

Therefore,

$$[\text{OH}^-_{(\text{aq})}] = 8.0 \times 10^{-4} \text{ mol/L}$$

$$\text{pOH} = -\log 8.0 \times 10^{-4}$$

$$\text{pOH} = 3.1$$

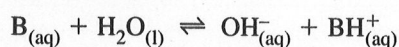
$$\text{pH} = 14 - \text{pOH}$$

$$= 14.0 - 3.1$$

$$\text{pH} = 10.9$$

The pH of a 0.1 mol/L solution of erythromycin is 10.9.

**pH of atropine ...**



$$K_b = \frac{[\text{BH}^+_{(\text{aq})}][\text{OH}^-_{(\text{aq})}]}{[\text{B}_{(\text{aq})}]}$$

$$K_b = 3.2 \times 10^{-5}$$

ICE Table for the ionization of the Base Atropine $\text{B}_{(\text{aq})}$				
	$\text{B}_{(\text{aq})}$	$\text{H}_2\text{O}_{(\text{l})} \rightleftharpoons$	$\text{OH}^-_{(\text{aq})}$	$\text{BH}^+_{(\text{aq})}$
Initial concentration (mol/L)	0.1	–	0.0	0.0
Change in concentration (mol/L)	–x	–	+x	+x
Equilibrium concentration (mol/L)	$0.1 - x$	–	x	x

Since  $2\% < 5\%$ , the assumption is valid.

Therefore,

$$[\text{OH}^-_{(\text{aq})}] = 2 \times 10^{-3} \text{ mol/L}$$

$$\text{pOH} = -\log 2 \times 10^{-3}$$

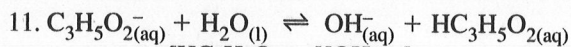
$$\text{pOH} = 2.7$$

$$\text{pH} = 14 - \text{pOH}$$

$$= 14.0 - 2.7$$

$$\text{pH} = 11.3$$

The pH of a 0.1 mol/L solution of atropine is 11.3.



$$K_b = \frac{[\text{HC}_3\text{H}_5\text{O}_2_{(\text{aq})}][\text{OH}^-_{(\text{aq})}]}{[\text{C}_3\text{H}_5\text{O}_2^-_{(\text{aq})}]}$$

ICE Table for the ionization of $\text{C}_3\text{H}_5\text{O}_2^-_{(\text{aq})}$				
	$\text{C}_3\text{H}_5\text{O}_2^-_{(\text{aq})}$	$\text{H}_2\text{O}_{(\text{l})} \rightleftharpoons$	$\text{OH}^-_{(\text{aq})}$	$\text{HC}_3\text{H}_5\text{O}_2_{(\text{aq})}$
Initial concentration (mol/L)	0.157	–	0.000	0.000
Change in concentration (mol/L)	$-1.1 \times 10^{-5}$	–	$+1.1 \times 10^{-5}$	$+1.1 \times 10^{-5}$
Equilibrium concentration (mol/L)	0.157	–	$1.1 \times 10^{-5}$	$1.1 \times 10^{-5}$

$$K_b = \frac{[\text{HC}_3\text{H}_5\text{O}_2_{(\text{aq})}][\text{OH}^-_{(\text{aq})}]}{[\text{C}_3\text{H}_5\text{O}_2^-_{(\text{aq})}]}$$

$$= \frac{(1.1 \times 10^{-5})^2}{0.157}$$

$$K_b = 7.7 \times 10^{-10}$$

The base ionization constant of the propanoate ion is  $7.7 \times 10^{-10}$ .

12.  $K_w = 1.0 \times 10^{-14}$

$$K_a = 7.2 \times 10^{-4}$$

$$K_a K_b = K_w$$

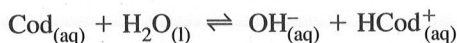
$$K_b = \frac{K_w}{K_a}$$

$$= \frac{1.0 \times 10^{-14}}{7.2 \times 10^{-4}}$$

$$K_b = 1.4 \times 10^{-11}$$

The base dissociation constant for the nitrite ion is  $1.4 \times 10^{-11}$ .

13. pH of codeine ...



$$K_b = \frac{[\text{HCod}^+_{(\text{aq})}][\text{OH}^-_{(\text{aq})}]}{[\text{Cod}_{(\text{aq})}]}$$

$$K_b = 1.73 \times 10^{-6}$$

ICE Table for the Base $\text{Cod}_{(\text{aq})}$				
	$\text{Cod}_{(\text{aq})}$	$\text{H}_2\text{O}_{(\text{l})}$	$\text{OH}^-_{(\text{aq})}$	$\text{HCod}^+_{(\text{aq})}$
Initial concentration (mol/L)	0.020	—	0.000	0.000
Change in concentration (mol/L)	-x	—	+x	+x
Equilibrium concentration (mol/L)	0.020 - x	—	x	x

Since  $0.93\% < 5\%$ , the assumption is valid.

Therefore,

$$[\text{OH}^-_{(\text{aq})}] = 1.9 \times 10^{-4} \text{ mol/L}$$

$$\text{pOH} = -\log 1.9 \times 10^{-4}$$

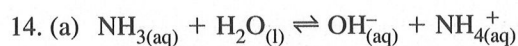
$$\text{pOH} = 3.73$$

$$\text{pH} = 14 - \text{pOH}$$

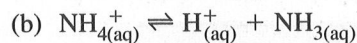
$$= 14.0 - 3.73$$

$$\text{pH} = 10.27$$

The pH of a 0.020 mol/L solution of codeine is 10.27.



$$K_b = \frac{[\text{NH}_4^+_{(\text{aq})}][\text{OH}^-_{(\text{aq})}]}{[\text{NH}_3_{(\text{aq})}]}$$



$$K_a = \frac{[\text{H}^+_{(\text{aq})}][\text{NH}_3_{(\text{aq})}]}{[\text{NH}_4^+_{(\text{aq})}]}$$

Handwritten calculations:

$$3.46 \times 10^{-8} = x^2$$

$$1.86 \times 10^{-4} = x$$

$$[\text{OH}^-] = 1.9 \times 10^{-4}$$

$$\text{pOH} = -\log 1.9 \times 10^{-4}$$

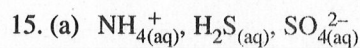
$$= 3.73$$

$$\text{pH} = 14 - 3.73 = 10.27$$

$$(c) \frac{[\text{H}_{(aq)}^+][\text{NH}_{3(aq)}]}{[\text{NH}_{4(aq)}^+]} \times \frac{[\text{NH}_{4(aq)}^+][\text{OH}_{(aq)}^-]}{[\text{NH}_{3(aq)}]} = [\text{H}_{(aq)}^+][\text{OH}_{(aq)}^-] = K_w$$

Numerically...

$$[5.80 \times 10^{-10}][1.72 \times 10^{-5}] = 1.00 \times 10^{-14}$$



(b)  $\text{NH}_3 : K_b = 1.7 \times 10^{-5}$

$\text{HS}_{(aq)}^- : K_b = 9.1 \times 10^{-8}$

$\text{SO}_{4(aq)}^{2-} : K_b = 1.0 \times 10^{-12}$

16. pH = 10.10

pOH = 14 - pH

= 14.0 - 10.10

pOH = 3.90

$[\text{OH}_{(aq)}^-] = 10^{-\text{pOH}}$

=  $10^{-3.90}$

$[\text{OH}_{(aq)}^-] = 1.26 \times 10^{-4} \text{ mol/L}$



$$K_b = \frac{[\text{HMor}_{(aq)}^+][\text{OH}_{(aq)}^-]}{[\text{Mor}_{(aq)}]}$$

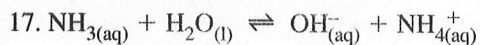
ICE Table for the Ionization of Morphine $\text{Mor}_{(aq)}$				
	$\text{Mor}_{(aq)}$	$+\text{H}_2\text{O}_{(l)} \rightleftharpoons$	$\text{OH}_{(aq)}^- +$	$\text{HMor}_{(aq)}^+$
Initial concentration (mol/L)	0.010	-	0.000	0.000
Change in concentration (mol/L)	$-1.26 \times 10^{-4}$	-	$+1.26 \times 10^{-4}$	$+1.26 \times 10^{-4}$
Equilibrium concentration (mol/L)	$9.87 \times 10^{-3}$	-	$1.26 \times 10^{-4}$	$1.26 \times 10^{-4}$

$$K_b = \frac{[\text{HMor}_{(aq)}^+][\text{OH}_{(aq)}^-]}{[\text{Mor}_{(aq)}]}$$

$$= \frac{(1.26 \times 10^{-4})^2}{9.87 \times 10^{-3}}$$

$$K_b = 1.6 \times 10^{-6}$$

The  $K_b$  for morphine is  $1.6 \times 10^{-6}$ .



$$K_b = \frac{[\text{NH}_{4(aq)}^+][\text{OH}_{(aq)}^-]}{[\text{NH}_{3(aq)}]}$$

$$K_b = 1.77 \times 10^{-5}$$

ICE Table for the Ionization of Ammonia				
	$\text{NH}_{3(aq)} +$	$\text{H}_2\text{O}_{(l)} \rightleftharpoons$	$\text{OH}_{(aq)}^- +$	$\text{NH}_{4(aq)}^+$
Initial concentration (mol/L)	0.100	-	0.000	0.000
Change in concentration (mol/L)	-x	-	+x	+x
Equilibrium concentration (mol/L)	$0.100 - x$	-	x	x



Since  $1.33\% < 5\%$ , the assumption is valid.

Therefore,

$$[\text{OH}^-_{(\text{aq})}] = 1.33 \times 10^{-3} \text{ mol/L}$$

$$\text{pOH} = -\log 1.33 \times 10^{-3}$$

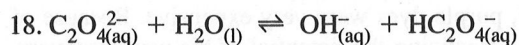
$$\text{pOH} = 2.877$$

$$\text{pH} = 14 - \text{pOH}$$

$$= 14.0 - 2.877$$

$$\text{pH} = 11.123$$

The pH of a 0.100 mol/L solution of ammonia is 11.123.



$$K_b = \frac{[\text{HC}_2\text{O}_4^-_{(\text{aq})}][\text{OH}^-_{(\text{aq})}]}{[\text{C}_2\text{O}_4^{2-}_{(\text{aq})}]}$$

$$K_b = 1.7 \times 10^{-10}$$

ICE Table for the Hydrolysis of Oxalate				
	$\text{C}_2\text{O}_4^{2-}_{(\text{aq})} +$	$\text{H}_2\text{O}_{(\text{l})} \rightleftharpoons$	$\text{OH}^-_{(\text{aq})} +$	$\text{HC}_2\text{O}_4^-_{(\text{aq})}$
Initial concentration (mol/L)	0.0500	–	0.000	0.000
Change in concentration (mol/L)	–x	–	+x	+x
Equilibrium concentration (mol/L)	$0.0500 - x$	–	x	x

Since  $5.83 \times 10^{-3} \% < 5\%$ , the assumption is valid.

Therefore,

$$[\text{OH}^-_{(\text{aq})}] = 2.9 \times 10^{-6} \text{ mol/L}$$

$$\text{pOH} = -\log 2.9 \times 10^{-6}$$

$$\text{pOH} = 5.54$$

$$\text{pH} = 14 - \text{pOH}$$

$$= 14.0 - 5.54$$

$$\text{pH} = 8.46$$

The pH of a 0.0500 mol/L sodium oxalate solution is 8.460.

### Applying Inquiry Skills

19. Strong bases have a higher pH than weaker bases. Weak bases that are molecular produce solutions that do not conduct electricity well.

### 20. Analysis

Solution 1:  $\text{HBr}_{(\text{aq})}$

Solution 2:  $\text{C}_{12}\text{H}_{22}\text{O}_{11(\text{aq})}$

Solution 3:  $\text{HC}_2\text{H}_3\text{O}_2_{(\text{aq})}$

Solution 4:  $\text{NaCl}_{(\text{aq})}$

### Making Connections

21. (a) pH = 10.10

$$\text{pOH} = 14 - \text{pH}$$

$$= 14.0 - 8.81$$

$$\text{pOH} = 5.19$$

$$[\text{OH}^-_{(\text{aq})}] = 10^{-\text{pOH}}$$

$$= 10^{-5.19}$$

$$[\text{OH}^-_{(\text{aq})}] = 6.5 \times 10^{-6} \text{ mol/L}$$

- (b) (Answers may vary. Answers might include the following information.) Aniline dyes are dyes derived from coal tar products. William Perkin accidentally discovered Mauvein, the first synthetic dye, while he was a 17-year-old student at the Royal College of Chemistry in London in 1856. Perkin's chemistry professor assigned him the task of synthesizing quinine, a drug to treat malaria. Perkin's first attempt ended in failure. Instead of quinine, Perkin had produced a seemingly useless black paste. Being the curious chemist, Perkin investigated the properties of this new material and found that it produced a brilliant purple colour when dissolved in alcohol. He also discovered that this dye had adhered to fabrics well, making it ideal as a fabric dye. In the nineteenth century, most fabric dyes came from natural sources. Purple, for example, came from the shell of a rare mollusk. Since hundreds of shells needed to be pulverized to produce a small amount of dye, purple dyes were very expensive. Because of its scarcity, purple was a popular colour among Europe's nobility. Realizing the potential of his new discovery, Perkin promptly left the College, patented the process for manufacturing aniline, and established a manufacturing facility in 1857.

ICE Table for the Ionization of Aniline				
	$\text{C}_6\text{H}_5\text{NH}_2_{(\text{aq})} +$	$\text{H}_2\text{O}_{(\text{l})} \rightleftharpoons$	$\text{OH}^-_{(\text{aq})} +$	$\text{HC}_6\text{H}_5\text{NH}_2^+_{(\text{aq})}$
Initial concentration (mol/L)	0.010	–	0.000	0.000
Change in concentration (mol/L)	$-6.5 \times 10^{-6}$	–	$+6.5 \times 10^{-6}$	$+6.5 \times 10^{-6}$
Equilibrium concentration (mol/L)	0.010	–	$6.5 \times 10^{-6}$	$6.5 \times 10^{-6}$

$$K_b = \frac{[\text{HC}_6\text{H}_5\text{NH}_2^+_{(\text{aq})}][\text{OH}^-_{(\text{aq})}]}{[\text{C}_6\text{H}_5\text{NH}_2_{(\text{aq})}]}$$

$$= \frac{(6.5 \times 10^{-6})^2}{0.10}$$

$$K_b = 4.2 \times 10^{-10}$$

The  $K_b$  for aniline is  $4.2 \times 10^{-10}$ .

22. (a) The high concentration of hydrogen ions in stomach fluid pushes the Aspirin equilibrium to the left, resulting in most of the Aspirin being in its un-ionized form.
- (b) Since this region of the stomach is of lower acidity, the forward reaction of the Aspirin equilibrium predominates to produce more hydrogen ions. This acidity is the cause of the irritation.
- (c) (Answers may vary, but may include the following information.) The enteric coating delays the release of Aspirin long enough for the tablet to pass through the stomach and into the intestine. This alleviates the potential irritation that Aspirin can cause.
23. (a) Let HLac represent the chemical formula for lactic acid.



$$K_a = \frac{[\text{H}^+_{(\text{aq})}][\text{Lac}^-_{(\text{aq})}]}{[\text{HLac}_{(\text{aq})}]}$$

$$K_a = 7.98 \times 10^{-5}$$

ICE Table for the Ionization of $\text{HC}_3\text{H}_5\text{O}_3_{(\text{aq})}$			
	$\text{HLac}_{(\text{aq})} \rightleftharpoons$	$\text{H}^+_{(\text{aq})} +$	$\text{Lac}^-_{(\text{aq})}$
Initial concentration (mol/L)	$5.6 \times 10^{-3}$	0	0
Change in concentration (mol/L)	$-x$	$+x$	$+x$
Equilibrium concentration (mol/L)	$5.6 \times 10^{-3} - x$	$x$	$x$