

1. Test all solutions with pH test strips. The strong acid has the lowest pH, the weak acid has the next lowest pH, and the remaining solutions have pH 7.
2. Test the pH 7 solutions with a conductivity tester. The neutral ionic solution will test positive while the neutral molecular solution will test negative.

Making Connections

9. (a) Gastroesophageal reflux disease (GERD) is a condition that results in the movement of stomach or duodenal contents into the esophagus.
 (b) GERD can affect all age groups.
 (c) A thin spaghetti-like tube containing a tiny pH probe is passed down the throat to the esophagus where it measures acidity over a 24-h period. Data from the probe is recorded on a recording device. Hospitalization is usually required while the test is being done.
 (d) Treatments currently available include
 Monitoring diet: the following foods can aggravate acid reflux: deep-fried foods, whole milk, chocolate, creamy foods. The following foods can aggravate an already-inflamed lower esophagus: coffee and other caffeinated beverages, carbonated soft drinks, citric juices like orange and grapefruit juice.
 Surgery: the lower portion of the esophagus is surgically tightened, the end result being a one-way valve, which allows food to enter the stomach while preventing stomach contents from flowing upward.
10. (a) Prior to the 19th century, paper was handmade from linen or rags. To meet the skyrocketing demand for paper in the 19th century, chemists found that large quantities of paper could be made economically from wood pulp. During this process, alum (aluminum sulfate) was added to the paper to fill the microscopic holes in the paper. This procedure is known as sizing. Sizing prevents ink from bleeding across the paper. The acidity of the aluminum ion catalyzes the degradation of cellulose strands of the paper. Over a long period of time, acidic paper becomes extremely brittle. Acid-free paper is manufactured often with alkaline sizing agents such as alkyl ketene dimers (AKD).
 (b) Acid-free paper is used for important documents that must last a long time, such as birth certificates, marriage licences, photograph album pages, archival records, etc.
 (c) Acid-free paper lasts considerably longer and is stronger than regular acidic paper. It also requires less energy and fresh water to manufacture. Paper made from this process can be recycled more easily. One disadvantage of acid-free paper is that the sizing agents tend to form sticky deposits on the paper-making equipment, and sometimes on the paper itself.

8.2 WEAK ACIDS AND BASES

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Understanding Concepts

1. $[\text{H}^+_{(\text{aq})}] = 10^{-2.54}$
 $[\text{H}^+_{(\text{aq})}] = 2.88 \times 10^{-3} \text{ mol/L}$
 $[\text{H}^+_{(\text{aq})}] = \frac{p}{100} \times [\text{HC}_2\text{H}_3\text{O}_2]$

$$p = \frac{[\text{H}^+_{(\text{aq})}]}{[\text{HC}_2\text{H}_3\text{O}_2]} \times 100\%$$

$$= \frac{2.88 \times 10^{-3} \text{ mol/L}}{[0.46 \text{ mol/L}]} \times 100\%$$

 $p = 0.63\%$
2. $[\text{H}^+_{(\text{aq})}] = 10^{-2.00}$
 $[\text{H}^+_{(\text{aq})}] = 1.0 \times 10^{-2} \text{ mol/L}$
 $[\text{H}^+_{(\text{aq})}] = \frac{p}{100} \times [\text{HF}_{(\text{aq})}]$

$$p = \frac{[\text{H}_{(\text{aq})}^+]}{[\text{HF}_{(\text{aq})}]} \times 100\%$$

$$= \frac{1.0 \times 10^{-2} \text{ mol/L}}{[0.15 \text{ mol/L}]} \times 100\%$$

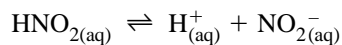
$$p = 6.7\%$$

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Understanding Concepts

3. 5.8%



$$K_a = \frac{[\text{H}_{(\text{aq})}^+][\text{NO}_{2(\text{aq})}^-]}{[\text{HNO}_{2(\text{aq})}]}$$

$$x = 200 \text{ mol/L} \times 0.058$$

$$x = 0.0116 \text{ mol/L} \quad (\text{extra digits carried})$$

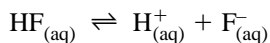
ICE Table for the Ionization of $\text{HNO}_{2(\text{aq})}$			
	$\text{HNO}_{2(\text{aq})} \rightleftharpoons$	$\text{H}_{(\text{aq})}^+ +$	$\text{NO}_{2(\text{aq})}^-$
Initial concentration (mol/L)	0.200		
Change in concentration (mol/L)	- 0.0116	+0.0116	+0.0116
Equilibrium concentration (mol/L)	0.188 mol/L	0.0116	0.0116

$$K_a = \frac{[\text{H}_{(\text{aq})}^+][\text{NO}_{2(\text{aq})}^-]}{[\text{HNO}_{2(\text{aq})}]}$$

$$= \frac{(0.0116)^2}{0.188}$$

$$K_a = 7.1 \times 10^{-4}$$

4. (a) 7.8%



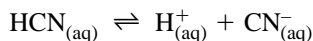
$$K_a = \frac{[\text{H}_{(\text{aq})}^+][\text{F}_{(\text{aq})}^-]}{[\text{HF}_{(\text{aq})}]}$$

$$x = 0.100 \text{ mol/L} \times 0.078$$

$$x = 0.0078 \text{ mol/L}$$

$$[\text{H}_{(\text{aq})}^+] = 7.8 \times 10^{-3} \text{ mol/L}$$

(b) 0.0078%



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

$$x = 0.100 \text{ mol/L} \times 7.8 \times 10^{-4}$$

$$x = 7.8 \times 10^{-6} \text{ mol/L}$$

$$[\text{H}_{(\text{aq})}^+] = 7.8 \times 10^{-6} \text{ mol/L}$$

(c) The hydrofluoric solution (a) is more acidic.

$$\begin{aligned}
 5. \quad [\text{H}^+_{(\text{aq})}] &= \frac{p}{100} \times [\text{HA}_{(\text{aq})}] \\
 p &= \frac{[\text{H}^+_{(\text{aq})}]}{[\text{HA}_{(\text{aq})}]} \times 100\% \\
 &= \frac{1.16 \times 10^{-3} \text{ mol/L}}{[0.100 \text{ mol/L}]} \times 100\% \\
 p &= 1.16\%
 \end{aligned}$$

The percent ionization of propanoic acid is 1.16%.

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Understanding Concepts

$$6. (a) \quad K_w = 1.0 \times 10^{-14}$$

$$K_a = 2.9 \times 10^{-8}$$

$$K_a K_b = K_w$$

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

$$= \frac{1.0 \times 10^{-14}}{2.9 \times 10^{-8}}$$

$$K_b = 3.4 \times 10^{-7}$$

The base dissociation constant for the hypochlorite ion is 3.4×10^{-7} .

$$(b) \quad K_w = 1.0 \times 10^{-14}$$

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

$$K_a K_b = K_w$$

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

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

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

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

$$(c) \quad K_w = 1.0 \times 10^{-14}$$

$$K_a = 6.3 \times 10^{-5}$$

$$K_a K_b = K_w$$

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

$$= \frac{1.0 \times 10^{-14}}{6.3 \times 10^{-5}}$$

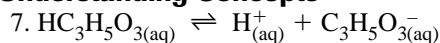
$$K_b = 1.6 \times 10^{-10}$$

The base dissociation constant for the benzoate ion is 1.6×10^{-10} .

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$$K_a = \frac{[\text{H}_{(\text{aq})}^+][\text{C}_3\text{H}_5\text{O}_{3(\text{aq})}^-]}{[\text{HC}_3\text{H}_5\text{O}_{3(\text{aq})}]}$$

$$K_a = 1.4 \times 10^{-4}$$

Predict whether a simplifying assumption is justified:

$$\frac{[\text{HA}]_{\text{initial}}}{K_a} = \frac{0.0010 \text{ mol/L}}{1.4 \times 10^{-4}}$$

$$\frac{[\text{HA}]_{\text{initial}}}{K_a} = 7.14$$

Since $7.14 < 100$, we may not assume that $0.0010 - x \doteq 0.0010$.

ICE Table for the Ionization of $\text{HC}_3\text{H}_5\text{O}_{3(\text{aq})}$			
	$\text{HC}_3\text{H}_5\text{O}_{3(\text{aq})} \rightleftharpoons$	$\text{H}_{(\text{aq})}^+ +$	$\text{C}_3\text{H}_5\text{O}_{3(\text{aq})}^-$
Initial concentration (mol/L)	0.0010	0.0000	0.0000
Change in concentration (mol/L)	$-x$	$+x$	$+x$
Equilibrium concentration (mol/L)	$0.0010 - x$	x	x

$$\frac{x^2}{(0.0010 - x)} = 1.4 \times 10^{-4}$$

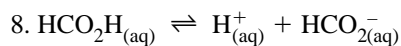
$$x^2 = 1.4 \times 10^{-4}(0.0010 - x)$$

$$x = \frac{-1.4 \times 10^{-4} \pm \sqrt{(1.4 \times 10^{-4})^2 - 4(-1.4 \times 10^{-7})}}{2}$$

$$x = 3.1 \times 10^{-4}$$

$$\begin{aligned} \text{pH} &= -\log [\text{H}_{(\text{aq})}^+] \\ &= -\log [3.1 \times 10^{-4}] \end{aligned}$$

$$\text{pH} = 3.51$$



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

$$K_a = 1.8 \times 10^{-4}$$

Predict whether a simplifying assumption is justified:

$$\frac{[\text{HA}]_{\text{initial}}}{K_a} = \frac{0.150 \text{ mol/L}}{1.8 \times 10^{-4}}$$

$$\frac{[\text{HA}]_{\text{initial}}}{K_a} = 830$$

Since $830 > 100$, we may assume that $0.150 - x = 0.150$.

The equilibrium expression becomes

$$\frac{x^2}{0.150} \doteq 1.8 \times 10^{-4}$$

which yields

$$x^2 \doteq 2.7 \times 10^{-5}$$

$$x \doteq 5.2 \times 10^{-3}$$

Validation of the approximation:

In general, the approximation is valid if

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

$$[\text{HA}]_{\text{initial}} = 0.15 \text{ mol/L}$$

$$\frac{x}{[\text{HA}]_{\text{initial}}} \times 100\% = \frac{5.2 \times 10^{-3}}{0.15} \times 100\%$$

$$\frac{x}{[\text{HA}]_{\text{initial}}} = 3.5\%$$

Since $3.5\% < 5\%$, the assumption is valid.

Therefore,

$$[\text{H}_{(\text{aq})}^+] = 5.2 \times 10^{-3} \text{ mol/L}$$

$$\text{pH} = -\log 5.2 \times 10^{-3}$$

$$\text{pH} = 2.28$$

PRACTICE

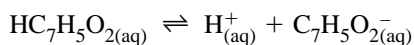
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Understanding Concepts

9. $\text{pH} = 2.40$

$$\begin{aligned} [\text{H}_{(\text{aq})}^+] &= 10^{-\text{pH}} \\ &= 10^{-2.40} \end{aligned}$$

$$[\text{H}_{(\text{aq})}^+] = 3.98 \times 10^{-3} \text{ mol/L}$$



$$K_a = \frac{[\text{H}_{(\text{aq})}^+][\text{C}_7\text{H}_5\text{O}_{2(\text{aq})}^-]}{[\text{HC}_7\text{H}_5\text{O}_{2(\text{aq})}]}$$

$$[\text{H}_{(\text{aq})}^+] = [\text{C}_7\text{H}_5\text{O}_{2(\text{aq})}^-]$$

$$[\text{H}_{(\text{aq})}^+] = 3.98 \times 10^{-3} \text{ mol/L}$$

$$K_a = \frac{(3.98 \times 10^{-3})^2}{0.25}$$

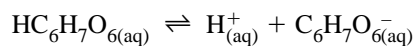
$$K_a = 6.3 \times 10^{-5}$$

The K_a for benzoic acid is 6.3×10^{-5} .

10. $\text{pH} = 2.40$

$$\begin{aligned} [\text{H}_{(\text{aq})}^+] &= 10^{-\text{pH}} \\ &= 10^{-2.40} \end{aligned}$$

$$[\text{H}_{(\text{aq})}^+] = 3.98 \times 10^{-3} \text{ mol/L}$$



$$K_a = \frac{[\text{H}_{(\text{aq})}^+][\text{C}_6\text{H}_7\text{O}_{6(\text{aq})}^-]}{[\text{HC}_6\text{H}_7\text{O}_{6(\text{aq})}]}$$

$$[\text{H}_{(\text{aq})}^+] = [\text{C}_6\text{H}_7\text{O}_{6(\text{aq})}^-]$$

$$[\text{H}_{(\text{aq})}^+] = 3.98 \times 10^{-3} \text{ mol/L}$$

$$K_a = \frac{(3.98 \times 10^{-3})^2}{0.20}$$

$$K_a = 7.9 \times 10^{-5}$$

The K_a for ascorbic acid is 7.9×10^{-5} .

Applying Inquiry Skills

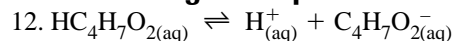
11. Experimental Design

1. Calibrate the pH following the manufacturer's instructions.
2. Add 25 mL of acetic acid to a 100-mL beaker.
3. Measure the pH of the solution using the pH meter.
4. Rinse the pH meter as directed by your teacher.
5. Discard the acetic acid solution as directed by your teacher.

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$$K_a = \frac{[\text{H}_{(\text{aq})}^+][\text{C}_4\text{H}_7\text{O}_{2(\text{aq})}^-]}{[\text{HC}_4\text{H}_7\text{O}_{2(\text{aq})}]}$$

$$K_a = 1.8 \times 10^{-4}$$

Predict whether a simplifying assumption is justified:

$$\frac{[\text{HA}]_{\text{initial}}}{K_a} = \frac{0.150 \text{ mol/L}}{1.8 \times 10^{-4}}$$

$$\frac{[\text{HA}]_{\text{initial}}}{K_a} = 830$$

Since $830 > 100$, we may assume that $0.150 - x \doteq 0.150$.

The equilibrium expression becomes

$$\frac{x^2}{0.150} \doteq 1.8 \times 10^{-4}$$

which yields

$$x^2 \doteq 2.7 \times 10^{-5}$$

$$x \doteq 5.2 \times 10^{-3}$$

Validation of the approximation:

In general, the approximation is valid if

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

$$[\text{HA}]_{\text{initial}} = 0.15 \text{ mol/L}$$

$$\frac{x}{[\text{HA}]_{\text{initial}}} \times 100\% = \frac{5.2 \times 10^{-3}}{0.15} \times 100\%$$

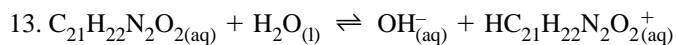
$$\frac{x}{[\text{HA}]_{\text{initial}}} = 3.5\%$$

Since $3.5\% < 5\%$, the assumption is valid.
Therefore,

$$[\text{H}^+_{(\text{aq})}] = 5.2 \times 10^{-3} \text{ mol/L}$$

$$\text{pH} = -\log 5.2 \times 10^{-3}$$

$$\text{pH} = 2.28$$



$$K_b = \frac{[\text{HC}_{21}\text{H}_{22}\text{N}_2\text{O}_{2(\text{aq})}^+][\text{OH}^-_{(\text{aq})}]}{[\text{C}_{21}\text{H}_{22}\text{N}_2\text{O}_{2(\text{aq})}]}$$

$$K_b = 1.0 \times 10^{-6}$$

ICE Table for the Ionization of $\text{C}_{21}\text{H}_{22}\text{N}_2\text{O}_{2(\text{aq})}$			
	$\text{C}_{21}\text{H}_{22}\text{N}_2\text{O}_{2(\text{aq})} +$	$\text{H}_2\text{O}_{(\text{l})} \rightleftharpoons \text{OH}^-_{(\text{aq})} +$	$\text{HC}_{21}\text{H}_{22}\text{N}_2\text{O}_{2(\text{aq})}^+$
Initial concentration (mol/L)	0.001	0.000	0.000
Change in concentration (mol/L)	x	$+x$	$+x$
Equilibrium concentration (mol/L)	$0.001 - x$	x	x

$$K_b = \frac{[\text{HC}_{21}\text{H}_{22}\text{N}_2\text{O}_{2(\text{aq})}^+][\text{OH}^-_{(\text{aq})}]}{[\text{C}_{21}\text{H}_{22}\text{N}_2\text{O}_{2(\text{aq})}]}$$

$$K_b = 1.0 \times 10^{-6}$$

$$\frac{x^2}{(0.001 - x)} = 1.0 \times 10^{-6}$$

Predicting the validity of the assumption ...

$$\frac{0.001}{1.0 \times 10^{-6}} = 1000$$

Since $1000 > 100$, we may assume that $0.001 - x \doteq 0.001$.

The equilibrium expression becomes

$$\frac{x^2}{0.001} \doteq 1.0 \times 10^{-6}$$

which yields

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

$$x \doteq 3.2 \times 10^{-5}$$

Justifying the simplification assumption ...

$$\frac{3.2 \times 10^{-5}}{0.001} \times 100\% = 3.2\%$$

Since $3.2\% < 5\%$, the assumption is valid.

Therefore,

$$[\text{OH}^-_{(\text{aq})}] = 3.2 \times 10^{-5} \text{ mol/L}$$

$$\text{pOH} = \log 3.2 \times 10^{-5}$$

$$\text{pOH} = 4.5$$

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

$$= 14.0 - 4.5$$

$$\text{pH} = 9.5$$

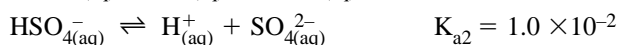
The pH of a 0.001 mol/L solution of strychnine is 9.5.

PRACTICE

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Understanding Concepts

14. (a) Since $\text{H}_2\text{SO}_{4(\text{aq})}$ is a strong acid, the hydrogen ion concentration from K_{a1} is 1.00 mol/L.



ICE Table for the Ionization of $\text{HSO}_{4(\text{aq})}^-$			
	$\text{HSO}_{4(\text{aq})}^- \rightleftharpoons$	$\text{H}^+_{(\text{aq})} +$	$\text{SO}_{4(\text{aq})}^{2-}$
Initial concentration (mol/L)	1.00	0.00	0.00
Change in concentration (mol/L)	$-x$	$+x$	$+x$
Equilibrium concentration (mol/L)	$1.00 - x$	x	x

$$K_a = \frac{[\text{H}^+_{(\text{aq})}][\text{HSO}_{4(\text{aq})}^-]}{[\text{SO}_{4(\text{aq})}^{2-}]}$$

$$\frac{x^2}{1.00 - x} = 1.0 \times 10^{-2}$$

Predicting whether $1.00 - x \doteq 1.00 \dots$

$$\frac{[\text{HA}]_{\text{initial}}}{K_a} = \frac{1.00 \text{ mol/L}}{1.0 \times 10^{-2}}$$

$$\frac{[\text{HA}]_{\text{initial}}}{K_a} = 100$$

Since $100 = 100$, we can assume that $1.00 - x \doteq 1.00$.

$$\frac{x^2}{(1.00 - x)} = 1.0 \times 10^{-2}$$

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

$$x = 1.0 \times 10^{-1}$$

$$x = 0.10$$

Validating the assumption ...

$$\frac{1.0 \times 10^{-2}}{1.00} \times 100\% = 1.0\%$$

Since $1.0\% < 5.0\%$, the assumption is valid.

$$[\text{H}^+_{(\text{aq})}] = 0.10 + 1.00$$

$$[\text{H}^+_{(\text{aq})}] = 1.1 \text{ mol/L}$$

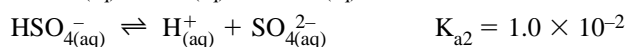
$$\text{pH} = -\log[\text{H}^+_{(\text{aq})}]$$

$$= -\log[1.10]$$

$$\text{pH} = -0.0414$$

The pH of a 0.001 mol/L solution of sulfuric acid is -0.0414 .

(b) Since $\text{H}_2\text{SO}_{4(\text{aq})}$ is a strong acid, the hydrogen ion concentration from K_{a1} is 0.001 mol/L.



ICE Table for the Ionization of $\text{HSO}_{4(\text{aq})}^-$			
	$\text{HSO}_{4(\text{aq})}^- \rightleftharpoons$	$\text{H}_{(\text{aq})}^+ +$	$\text{SO}_{4(\text{aq})}^{2-}$
Initial concentration (mol/L)	0.001	0.000	0.000
Change in concentration (mol/L)	$-x$	$+x$	$+x$
Equilibrium concentration (mol/L)	$0.001 - x$	x	x

$$K_a = \frac{[\text{H}_{(\text{aq})}^+][\text{HSO}_{4(\text{aq})}^-]}{[\text{SO}_{4(\text{aq})}^{2-}]}$$

$$\frac{x^2}{0.001 - x} = 1.0 \times 10^{-2}$$

Predicting whether $0.001 - x \doteq 0.001 \dots$

$$\frac{[\text{HA}]_{\text{initial}}}{K_a} = \frac{0.001 \text{ mol/L}}{1.0 \times 10^{-2}}$$

$$\frac{[\text{HA}]_{\text{initial}}}{K_a} = 0.1$$

Since $0.1 < 100$, we cannot assume that $0.001 - x \doteq 0.001$.

$$\frac{x^2}{(0.001 - x)} = 1.0 \times 10^{-2}$$

$$x^2 = 1.0 \times 10^{-2} (0.001 - x)$$

$$x^2 = 1.0 \times 10^{-5} - (1.0 \times 10^{-5}x)$$

$$x^2 + (1.0 \times 10^{-5}x) - (1.0 \times 10^{-5}) = 0$$

$$x = \frac{-1.0 \times 10^{-5} \pm \sqrt{(1.0 \times 10^{-5})^2 - 4(-1.0 \times 10^{-5})}}{2}$$

$$x = 9 \times 10^{-4}$$

$$[\text{H}_{(\text{aq})}^+] = 0.001 + (9 \times 10^{-4})$$

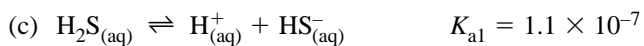
$$[\text{H}_{(\text{aq})}^+] = 0.0019$$

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

$$= -\log[0.0019]$$

$$\text{pH} = 2.7$$

The pH of a 0.001 mol/L solution of sulfuric acid is 2.7.



Since $\text{H}_2\text{S}_{(\text{aq})}$ is a much stronger acid than $\text{HS}_{(\text{aq})}^-$, it will dominate the production of $\text{H}_{(\text{aq})}^+$.

$$K_{a1} = \frac{[\text{H}_{(\text{aq})}^+][\text{HS}_{(\text{aq})}^-]}{[\text{H}_2\text{S}_{(\text{aq})}]}$$

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 $0.010 - x \doteq 0.010 \dots$

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

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

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

1. 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.
2. $\text{CN}^-_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \rightleftharpoons \text{HCN}_{(\text{aq})} + \text{OH}^-_{(\text{aq})}$

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

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 $0.18 - x \doteq 0.18 \dots$

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

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

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

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

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

Validating the assumption ...

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

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

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

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

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

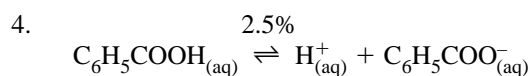
$$\text{pOH} = 2.77$$

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

$$= 14 - 2.77$$

$$\text{pH} = 11.23$$

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



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

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

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

ICE Table for the Ionization of $\text{C}_6\text{H}_5\text{COOH}_{(\text{aq})}$			
	$\text{C}_6\text{H}_5\text{COOH}_{(\text{aq})} \rightleftharpoons$	$\text{H}^+_{(\text{aq})} +$	$\text{C}_6\text{H}_5\text{COO}^-_{(\text{aq})}$
Initial concentration (mol/L)	0.100	0.00	0.00
Change in concentration (mol/L)	-2.5×10^{-3}	$+2.5 \times 10^{-3}$	$+2.5 \times 10^{-3}$
Equilibrium concentration (mol/L)	0.098	2.5×10^{-3}	2.5×10^{-3}

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

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

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

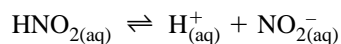
The K_a of benzoic acid is 6.4×10^{-5} .

5. $\text{pH} = 2.1$

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

$$= 10^{-2.1}$$

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



$$K_a = \frac{[\text{H}^+_{(\text{aq})}][\text{NO}_{2(\text{aq})}^-]}{[\text{HNO}_{2(\text{aq})}]}$$

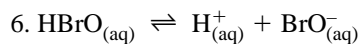
ICE Table for the Ionization of $\text{HNO}_{2(\text{aq})}$			
	$\text{HNO}_{2(\text{aq})} \rightleftharpoons$	$\text{H}^+_{(\text{aq})} +$	$\text{NO}_{2(\text{aq})}^-$
Initial concentration (mol/L)	0.100	0	0
Change in concentration (mol/L)	-8×10^{-3}	$+8 \times 10^{-3}$	$+8 \times 10^{-3}$
Equilibrium concentration (mol/L)	0.09	8×10^{-3}	8×10^{-3}

$$K_a = \frac{[\text{H}^+_{(\text{aq})}][\text{NO}_{2(\text{aq})}^-]}{[\text{HNO}_{2(\text{aq})}]}$$

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

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

The K_a for nitrous acid is 7×10^{-4} .



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

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

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,

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

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

$$= 2.28$$

$$\text{pH} = 4.65$$

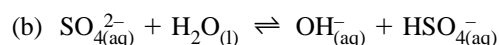
The pH of the hypobromous acid solution is 4.65.

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:

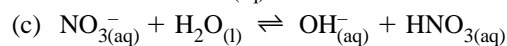
Acid	Approximate pH
$\text{HCN}_{(\text{aq})}$	5
$\text{HCO}_2\text{H}_{(\text{aq})}$	4
$\text{HF}_{(\text{aq})}$	2
$\text{HNO}_{3(\text{aq})}$	0

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) $\text{CN}^-_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \rightleftharpoons \text{OH}^-_{(\text{aq})} + \text{HCN}_{(\text{aq})}$

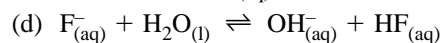
$$K_b = \frac{[\text{HCN}_{(\text{aq})}][\text{OH}^-_{(\text{aq})}]}{[\text{CN}^-_{(\text{aq})}]}$$



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



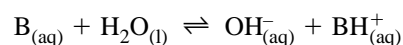
$$K_b = \frac{[\text{HNO}_{3(aq)}][\text{OH}_{(aq)}^-]}{[\text{NO}_{3(aq)}^-]}$$



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

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,

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

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

$$\text{pOH} = 3.6$$

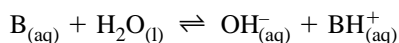
$$\text{pH} = 14 - \text{pOH}$$

$$= 14.0 - 3.6$$

$$\text{pH} = 10.4$$

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,

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

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

$$\text{pOH} = 3.1$$

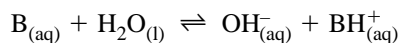
$$\text{pH} = 14 - \text{pOH}$$

$$= 14.0 - 3.1$$

$$\text{pH} = 10.9$$

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,

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

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

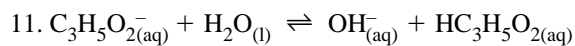
$$\text{pOH} = 2.7$$

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

$$= 14.0 - 2.7$$

$$\text{pH} = 11.3$$

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)	-1.1×10^{-5}	—	$+1.1 \times 10^{-5}$	$+1.1 \times 10^{-5}$
Equilibrium concentration (mol/L)	0.157	—	1.1×10^{-5}	1.1×10^{-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 7.7×10^{-10} .

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

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

$$K_a K_b = K_w$$

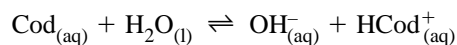
$$K_b = \frac{K_w}{K_a}$$

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

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

The base dissociation constant for the nitrite ion is 1.4×10^{-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 $\text{Cod}_{(\text{aq})}$				
	$\text{Cod}_{(\text{aq})} +$	$\text{H}_2\text{O}_{(\text{l})} \rightleftharpoons$	$\text{OH}^-_{(\text{aq})} +$	$\text{HCod}^+_{(\text{aq})}$
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,

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

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

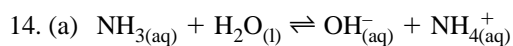
$$\text{pOH} = 3.73$$

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

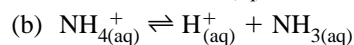
$$= 14.0 - 3.73$$

$$\text{pH} = 10.27$$

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



$$K_b = \frac{[\text{NH}_4^+_{(\text{aq})}][\text{OH}^-_{(\text{aq})}]}{[\text{NH}_{3(\text{aq})}]}$$

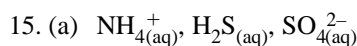


$$K_a = \frac{[\text{H}^+_{(\text{aq})}][\text{NH}_{3(\text{aq})}]}{[\text{NH}_4^+_{(\text{aq})}]}$$

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

Numerically...

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



(b) $NH_3 : K_b = 1.7 \times 10^{-5}$

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

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

16. pH = 10.10

$$pOH = 14 - pH$$

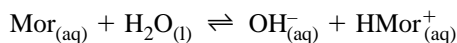
$$= 14.0 - 10.10$$

$$pOH = 3.90$$

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

$$= 10^{-3.90}$$

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



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

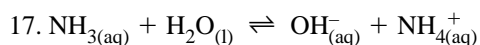
ICE Table for the Ionization of Morphine $Mor_{(aq)}$				
	$Mor_{(aq)}$	$+H_2O_{(l)} \rightleftharpoons$	$OH^-_{(aq)} +$	$HMor^+_{(aq)}$
Initial concentration (mol/L)	0.010	—	0.000	0.000
Change in concentration (mol/L)	-1.26×10^{-4}	—	$+1.26 \times 10^{-4}$	$+1.26 \times 10^{-4}$
Equilibrium concentration (mol/L)	9.87×10^{-3}	—	1.26×10^{-4}	1.26×10^{-4}

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

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

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

The K_b for morphine is 1.6×10^{-6} .



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

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

ICE Table for the Ionization of Ammonia				
	$NH_{3(aq)} +$	$H_2O_{(l)} \rightleftharpoons$	$OH^-_{(aq)} +$	$NH_{4(aq)}^+$
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,

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

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

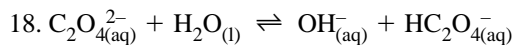
$$\text{pOH} = 2.877$$

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

$$= 14.0 - 2.877$$

$$\text{pH} = 11.123$$

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



$$K_b = \frac{[\text{HC}_2\text{O}_{4(\text{aq})}^-][\text{OH}_{(\text{aq})}^-]}{[\text{C}_2\text{O}_{4(\text{aq})}^{2-}]}$$

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

ICE Table for the Hydrolysis of Oxalate				
	$\text{C}_2\text{O}_{4(\text{aq})}^{2-} +$	$\text{H}_2\text{O}_{(\text{l})} \rightleftharpoons$	$\text{OH}_{(\text{aq})}^- +$	$\text{HC}_2\text{O}_{4(\text{aq})}^-$
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) $\text{pH} = 10.10$

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

$$= 14.0 - 8.81$$

$$\text{pOH} = 5.19$$

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

$$= 10^{-5.19}$$

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

- (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)	-6.5×10^{-6}	—	$+6.5 \times 10^{-6}$	$+6.5 \times 10^{-6}$
Equilibrium concentration (mol/L)	0.010	—	6.5×10^{-6}	6.5×10^{-6}

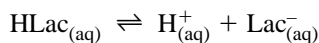
$$K_b = \frac{[\text{HC}_6\text{H}_5\text{NH}_{2(\text{aq})}^+][\text{OH}^-_{(\text{aq})}]}{[\text{C}_6\text{H}_5\text{NH}_{2(\text{aq})}]}$$

$$= \frac{(6.5 \times 10^{-6})^2}{0.010}$$

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

The K_b for aniline is 4.2×10^{-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)	5.6×10^{-3}	0	0
Change in concentration (mol/L)	$-x$	$+x$	$+x$
Equilibrium concentration (mol/L)	$5.6 \times 10^{-3} - x$	x	x

$$\frac{x^2}{(5.6 \times 10^{-3} - x)} = 7.94 \times 10^{-5}$$

Predict whether a simplifying assumption is justified ...

$$\frac{[\text{HA}]_{\text{initial}}}{K_a} = \frac{5.6 \times 10^{-3}}{7.94 \times 10^{-5}}$$

$$\frac{[\text{HA}]_{\text{initial}}}{K_a} = 70$$

Since $70 < 100$, we may not assume that $5.6 \times 10^{-3} - x \doteq 5.6 \times 10^{-3}$.

$$\frac{x^2}{(5.6 \times 10^{-3} - x)} = 7.94 \times 10^{-5}$$

$$x^2 = 7.94 \times 10^{-5} (5.6 \times 10^{-3} - x)$$

$$x^2 + (7.94 \times 10^{-5} x) - (4.45 \times 10^{-7}) = 0$$

$$x = \frac{-7.94 \times 10^{-5} \pm \sqrt{(7.94 \times 10^{-5})^2 - 4(-4.45 \times 10^{-7})}}{2}$$

$$x = 6.29 \times 10^{-4} \text{ mol/L}$$

$$\begin{aligned} \text{pH} &= -\log[\text{H}_{(\text{aq})}^+] \\ &= -\log[6.29 \times 10^{-4}] \end{aligned}$$

$$\text{pH} = 3.20$$

The pH of the lactic acid in the runner's muscles is 3.20.

- (b) Lactic acid buildup in muscles causes fatigue, pain, and muscle stiffness.
- (c) Muscles can oxidize glucose aerobically or anaerobically to release energy. The energy released is stored temporarily in the molecule ATP that can then be used by muscles to do mechanical work. Anaerobic oxidation of glucose, however, is not very efficient. In a sprint, for example, a great deal of energy is required in a short period of time. Anaerobic oxidation supplies most of this energy but is very inefficient. Short-term energy reserves are depleted quickly and lactic acid accumulates.

Muscles used in long-distance running rely more on aerobic oxidation of glucose for their energy. When running at a comfortable pace, both systems of oxidation are used but the ratio of anaerobic: aerobic is low enough to prevent lactic acid from accumulating. As the pace increases, the anaerobic: aerobic ratio increases to the point where lactic acid begins to accumulate in the blood. This is known as the lactic acid threshold. In order to improve performance, long-distance runners try to train at the speed at which the lactic acid threshold occurs. This serves to increase the threshold and overall performance.

8.3 ACID-BASE PROPERTIES OF SALT SOLUTIONS

PRACTICE

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Understanding Concepts

1. (a) The ammonium ion is a weak acid with $K_a = 5.8 \times 10^{-10}$. The phosphate ion is a base with $K_b = 2.4 \times 10^{-2}$. Since K_b is larger than K_a , an ammonium phosphate solution is basic.
- (b) The ammonium ion is a weak acid with $K_a = 5.8 \times 10^{-10}$. The sulfate ion is a base with $K_b = 1.0 \times 10^{-12}$. Since K_a is larger than K_b , an ammonium sulfate solution is slightly acidic.
- (c) Magnesium oxide reacts with water to form magnesium hydroxide (a base).
- (d) $\text{MgO}_{(\text{s})} + \text{H}_2\text{O}_{(\text{l})} \rightarrow \text{Mg}_{(\text{aq})}^{2+} + 2 \text{OH}_{(\text{aq})}^-$. This makes a solution of magnesium oxide basic.
2. A solution of sodium sulfite will be basic.
3. $\text{NH}_4\text{NO}_{3(\text{aq})} \rightarrow \text{NH}_{4(\text{aq})}^+ + \text{NO}_{3(\text{aq})}^-$