

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

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

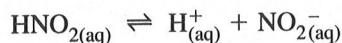
$$p = 6.7\%$$

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Understanding Concepts

3. 5.8%



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

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

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

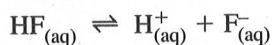
| ICE Table for the Ionization of $\text{HNO}_{2(\text{aq})}$ | | | |
|---|--|------------------------------|------------------------------|
| | $\text{HNO}_{2(\text{aq})} \rightleftharpoons$ | $\text{H}_{(\text{aq})}^+ +$ | $\text{NO}_{2(\text{aq})}^-$ |
| Initial concentration (mol/L) | 0.200 | | |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



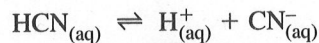
$$K_a = \frac{[\text{H}_{(\text{aq})}^+][\text{F}_{(\text{aq})}^-]}{[\text{HF}_{(\text{aq})}]}$$

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

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

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

(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.

$$5. \quad [\text{H}^+_{(\text{aq})}] = \frac{p}{100} \times [\text{HA}_{(\text{aq})}]$$

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

$$= \frac{1.16 \times 10^{-3} \text{ mol/L}}{[0.100 \text{ mol/L}]} \times 100\%$$

$$p = 1.16\%$$

The percent ionization of propanoic acid is 1.16%.

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

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

$$K_a K_b = K_w$$

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

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

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

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

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

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

$$K_a K_b = K_w$$

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

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

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

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

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

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

$$K_a K_b = K_w$$

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

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

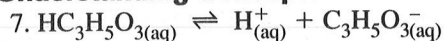
$$K_b = 1.6 \times 10^{-10}$$

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 |

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

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

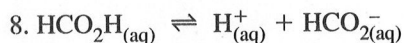
$$x = \frac{-1.4x \times 10^{-4} \pm \sqrt{(1.4x \times 10^{-4})^2 - 4(-1.4 \times 10^{-7})}}{2}$$

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

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

$$= -\log [3.1 \times 10^{-4}]$$

$$\text{pH} = 3.51$$



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

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

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

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

which yields

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

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

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,

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

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

$$\text{pH} = 2.28$$

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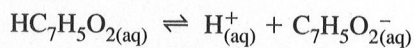
Understanding Concepts

9. $\text{pH} = 2.40$

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

$$= 10^{-2.40}$$

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



$$K_a = \frac{[\text{H}^+_{(\text{aq})}][\text{C}_7\text{H}_5\text{O}_2^-(\text{aq})]}{[\text{HC}_7\text{H}_5\text{O}_2(\text{aq})]}$$

$$[\text{H}^+_{(\text{aq})}] = [\text{C}_7\text{H}_5\text{O}_2^-(\text{aq})]$$

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

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

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

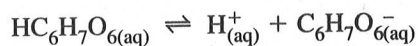
The K_a for benzoic acid is 6.3×10^{-5} .

10. $\text{pH} = 2.40$

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

$$= 10^{-2.40}$$

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



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

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

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

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

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

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

Applying Inquiry Skills

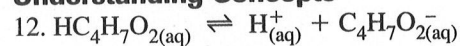
11. Experimental Design

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

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

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

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

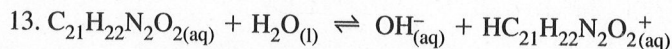
$$\frac{x}{[\text{HA}]_{\text{initial}}} = 3.5\%$$

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,

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

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

$$\text{pOH} = 4.5$$

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

$$= 14.0 - 4.5$$

$$\text{pH} = 9.5$$

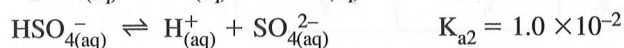
The pH of a 0.001 mol/L solution of strychnine is 9.5.

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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.

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

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

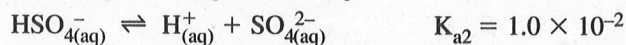
$$\text{pH} = -\log[\text{H}^+_{(\text{aq})}]$$

$$= -\log[1.10]$$

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

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 |
| Change in concentration (mol/L) | -x | | +x |
| Equilibrium concentration (mol/L) | 0.001 - 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 $0.001 - x \approx 0.001$...

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

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

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

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

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

$$x^2 = 1(1.0 \times 10^{-2}x) - (1.0 \times 10^{-5}x)$$

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

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

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

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

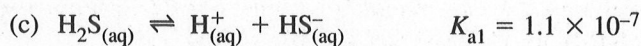
$$[\text{H}^+_{(\text{aq})}] = 0.0019$$

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

$$= -\log[0.0019]$$

$$\text{pH} = 2.7$$

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 $\text{H}^+_{(\text{aq})}$.

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