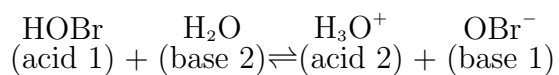


Chemistry 192  
Problem Set 3  
Spring, 2018  
Solutions

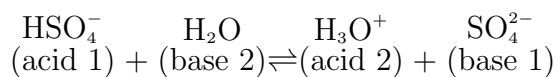
1. Problem 3, page 782, textbook

**Answer**

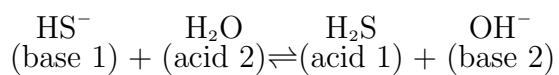
(a)



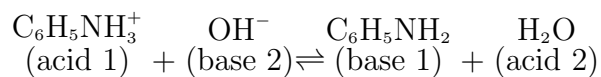
(b)



(c)



(d)



2. Calculate the pH and pOH of the following solutions of strong acids or bases:

(a) 0.0125 M  $\text{HCl}_{(aq)}$

**Answer:**

$$[\text{H}_3\text{O}^+] = 1.25 \times 10^{-2}$$

$$\text{pH} = -\log_{10}(1.25 \times 10^{-2}) = 1.903 \quad \text{pOH} = 14.00 - \text{pH} = 12.10$$

(b) 0.0460 M NaOH<sub>(aq)</sub>

**Answer:**

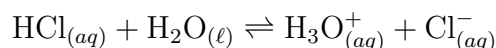
$$[\text{OH}^-] = 4.60 \times 10^{-2}$$

$$\text{pOH} = -\log_{10}(4.60 \times 10^{-2}) = 1.337 \quad \text{pH} = 14.00 - \text{pOH} = 12.66$$

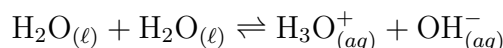
(c)  $1.75 \times 10^{-8}$  M HCl<sub>(aq)</sub>

**Answer:**

There are two equilibrium processes to consider



and



All the OH<sub>(aq)</sub><sup>-</sup> results from the self dissociation of water. We let  $x = [\text{OH}^-]$ . The H<sub>3</sub>O<sup>+</sup> concentration has contributions from both equilibria. Then  $[\text{H}_3\text{O}^+] = x + 1.75 \times 10^{-8}$ . Using the ion product of water

$$[\text{H}_3\text{O}^+][\text{OH}^-] = x(x + 1.75 \times 10^{-8}) = 1.00 \times 10^{-14}$$

or

$$x^2 + 1.75 \times 10^{-8}x - 1.00 \times 10^{-14} = 0$$

Using the quadratic formula

$$x = \frac{-1.75 \times 10^{-8} \pm \sqrt{(1.75 \times 10^{-8})^2 + 4 \times 10^{-14}}}{2} = 9.20 \times 10^{-8}, -1.10 \times 10^{-7}.$$

Only the positive solution for  $x$  is physical and

$$\text{pOH} = -\log_{10}(9.20 \times 10^{-8}) = 7.04 \quad \text{pH} = 14.00 - \text{pOH} = 6.96$$

(d)  $1.75 \times 10^{-8}$  M NaOH<sub>(aq)</sub>

**Answer:**

The method of solution is identical to the solution to problem 2c, except that we set  $x = [\text{H}_3\text{O}^+]$ . Then using similar algebra

$$\text{pH} = -\log_{10}(9.20 \times 10^{-8}) = 7.04 \quad \text{pOH} = 14.00 - \text{pH} = 6.96$$

3. Problem 14, page 782, textbook

**Answer:**

$$n_{\text{OH}^-} = (0.125 \text{ L})(0.606 \text{ mol L}^{-1}) = 0.0758 \text{ mol}$$

$$[\text{OH}^-] = \frac{0.0758 \text{ mol}}{15.0 \text{ L}} = 0.00505 \text{ M}$$

$$\text{pOH} = -\log_{10}(0.00505) = 2.297 \quad \text{pH} = 14.00 - \text{pOH} = 11.70$$

4. Problem 19, page 782, textbook

**Answer:**

Before mixing

$$n_{H^+} = (0.05000 \text{ L})(0.0155 \text{ mol L}^{-1}) = 0.000775 \text{ mol}$$

$$n_{OH^-} = (0.07500 \text{ L})(0.0106 \text{ mol L}^{-1}) = 0.000795 \text{ mol}$$

After mixing

$$n_{OH^-} = 0.000795 \text{ mol} - 0.000775 \text{ mol} = 0.0000200 \text{ mol}$$

$$[OH^-] = \frac{0.0000200 \text{ mol}}{0.125 \text{ L}} = 0.000160 \text{ M}$$

$$\text{pOH} = -\log_{10}(0.000160) = 3.796 \quad \text{pH} = 14.00 - \text{pOH} = 10.20$$

5. Problem 20, Page 782, textbook

**Answer:**

$$[H_3O^+] = 10^{-2.12} = 7.58 \times 10^{-3} \text{ M}$$

$$n_{H_3O^+} = (7.58 \times 10^{-3} \text{ mol L}^{-1})(0.02500 \text{ L}) = 1.94 \times 10^{-4} \text{ mol}$$

$$[OH^-] = 10^{-(14-12.65)} = 4.47 \times 10^{-2} \text{ M}$$

$$n_{OH^-} = (0.02500 \text{ L})(4.47 \times 10^{-2} \text{ mol L}^{-1}) = 1.17 \times 10^{-3} \text{ mol}$$

After mixing

$$n_{OH^-} = 1.17 \times 10^{-3} \text{ mol} - 1.94 \times 10^{-4} \text{ mol} = 9.76 \times 10^{-4} \text{ mol}$$

$$[OH^-] = \frac{9.76 \times 10^{-4} \text{ mol}}{0.0500 \text{ L}} = 1.95 \times 10^{-2} \text{ M}$$

$$\text{pOH} = -\log_{10}(1.95 \times 10^{-2}) = 1.71 \quad \text{pH} = 14.00 - \text{pOH} = 12.29$$

6. Calculate the volume of a 0.100 M aqueous KOH solution that must be added to 100. ml of a 0.100 M aqueous HCl solution to obtain a solution having a pH=2.000.

**Answer:**

Before mixing

$$n_{H_3O^+} = (0.100 \text{ mol L}^{-1})(0.100 \text{ L}) = 0.0100 \text{ mol}$$

Let  $V$  = volume of the KOH solution to be added. Then the final  $H_3O^+$  concentration is

$$\frac{n_{H_3O^+}}{V_{\text{final}}} = \frac{(0.0100 \text{ mol} - (0.100 \text{ mol L}^{-1})V)}{V + 0.100 \text{ L}} = 10^{-2} \text{ mol L}^{-1}$$

$$(10^{-2} \text{ mol L}^{-1})V + 1.00 \times 10^{-3} \text{ mol} = 0.0100 \text{ mol} - (0.100 \text{ mol L}^{-1})V$$

$$(0.110 \text{ mol L}^{-1})V = 0.00900 \text{ mol}$$

$$V = 0.0818 \text{ L} = 81.8 \text{ mL}$$

7. Given that sodium hydroxide is a strong base, calculate the pH of an aqueous  $7.5 \times 10^{-8}$  M sodium hydroxide solution. Approximations do not work for this problem.

**Answer:**

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

$$\text{Let } y = [\text{H}_3\text{O}^+] \quad [\text{OH}^-] = y + 7.5 \times 10^{-8}$$

$$y(y + 7.5 \times 10^{-8}) = 1.0 \times 10^{-14}$$

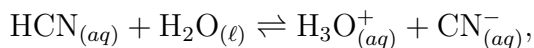
$$y^2 + 7.5 \times 10^{-8}y - 1.0 \times 10^{-14} = 0$$

$$y = \frac{-7.5 \times 10^{-8} \pm [(7.5 \times 10^{-8})^2 + 4.0 \times 10^{-14}]^{1/2}}{2}$$

Keeping the positive solution

$$y = [\text{H}_3\text{O}^+] = 6.9 \times 10^{-8} \text{ M} \quad \text{pH} = -\log_{10}(6.9 \times 10^{-8}) = 7.16$$

8. Calculate the pH and percent ionization of a 2.00 M aqueous HCN solution given that for the reaction



$$K_a = 4.1 \times 10^{-10}.$$

**Answer**

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CN}^-]}{[\text{HCN}]}$$

	[HCN]	[H <sub>3</sub> O <sup>+</sup> ]	[CN <sup>-</sup> ]
initial	2.00 M	0 M	0 M
change	-x M	x M	x M
equilibrium	(2.00-x) M	x M	x M

$$\frac{x^2}{2.00 - x} = 4.1 \times 10^{-10}$$

$$x^2 \cong 8.2 \times 10^{-10}$$

$$x = 2.9 \times 10^{-5}$$

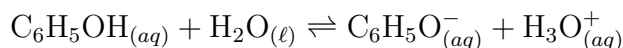
$$\% \text{ ionized} = \frac{2.9 \times 10^{-5}}{2.00} \times 100. = 1.4 \times 10^{-3}\%$$

$$\text{pH} = -\log_{10}(2.9 \times 10^{-5}) = 4.54$$

9. The  $pK_a$  of phenol ( $C_6H_5OH$ ) when acting as a weak acid in aqueous solution is 9.99. Calculate the pH of a 2.0 M phenol solution. Approximations work for this system.

**Answer:**

$$K_a = 10^{-9.99} = 1.0 \times 10^{-10}$$



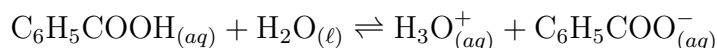
	$[C_6H_5OH]$	$[C_6H_5O^-]$	$[H_3O^+]$
initial	2.0 M	0 M	0 M
change	$-y$	$y$	$y$
equilibrium	$(2.0 - y)$ M	$y$ M	$y$ M

$$\frac{y^2}{2.0 - y} \approx \frac{y^2}{2.0} = 1.0 \times 10^{-10}$$

$$y = \sqrt{2.0 \times 10^{-10}} = 1.4 \times 10^{-5} \text{ M} = [H_3O^+]$$

$$\text{pH} = -\log_{10}(1.4 \times 10^{-5}) = 4.84$$

10. When a 0.072 M aqueous solution of benzoic acid reacts



the pH is found to be 2.68. Calculate the  $pK_a$  for the reaction.

**Answer**

Let  $BA \equiv C_6H_5COOH$  and  $A^- \equiv C_6H_5COO^-$ . Then

$$[H_3O^+] = 10^{-2.68} = 2.1 \times 10^{-3} = [A^-]$$

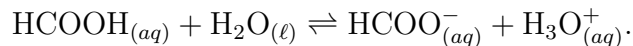
	$[BA]$	$[H_3O^+]$	$[A^-]$
initial	0.072 M	0 M	0 M
change	$-x$ M	$x$ M	$x$ M
equilibrium	$(0.072 - x)$ M	$x$ M	$x$ M

$$x = 2.1 \times 10^{-3}$$

$$K_a = \frac{[H_3O^+][A^-]}{[BA]} = \frac{x^2}{0.072 - x} = \frac{(2.1 \times 10^{-3})^2}{0.072 - 2.1 \times 10^{-3}} = 6.3 \times 10^{-5}$$

$$pK_a = -\log_{10}(6.3 \times 10^{-5}) = 4.20$$

11. Formic acid ( $\text{HCOOH}$ ) is a monoprotic acid with acid dissociation reaction



The pH of a  $1.20 \times 10^{-3}$  M formic acid solution is found to be 3.41. Calculate  $K_a$  and  $\text{p}K_a$  for formic acid.

**Answer:**

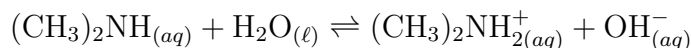
$$[\text{H}_3\text{O}^+] = 10^{-3.41} = 3.9 \times 10^{-4}$$

	$[\text{HCOOH}]$	$[\text{H}_3\text{O}^+]$	$[\text{HCOO}^-]$
initial	$1.20 \times 10^{-3}$ M	0 M	0 M
change	$-3.9 \times 10^{-4}$ M	$3.9 \times 10^{-4}$ M	$3.9 \times 10^{-4}$ M
equilibrium	$8.1 \times 10^{-4}$ M	$3.9 \times 10^{-4}$ M	$3.9 \times 10^{-4}$ M

$$K_a = \frac{[\text{HCOO}^-][\text{H}_3\text{O}^+]}{[\text{HCOOH}]} = \frac{(3.9 \times 10^{-4})^2}{8.1 \times 10^{-4}} = 1.9 \times 10^{-4}$$

$$\text{p}K_a = -\log_{10} K_a = -\log_{10}(1.9 \times 10^{-4}) = 3.73$$

12. Dimethylamine ionizes according to the reaction



A 0.95 M solution of dimethylamine is found to have a pH of 12.32. Calculate the  $\text{p}K_b$  of the reaction.

**Answer:**

$$K_b = \frac{[(\text{CH}_3)_2\text{NH}_2^+][\text{OH}^-]}{[(\text{CH}_3)_2\text{NH}]}$$

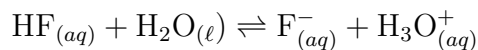
$$\text{pOH} = 14.00 - \text{pH} = 1.68 \quad [\text{OH}^-] = 10^{-1.68} = 2.1 \times 10^{-2}$$

	$[(\text{CH}_3)_2\text{NH}]$	$[(\text{CH}_3)_2\text{NH}_2^+]$	$[\text{OH}^-]$
initial	0.95 M	0 M	0 M
change	$-x$ M	$x$ M	$x$ M
equilibrium	$(0.95-x)$ M	$x$ M	$x$ M

$$\frac{x^2}{0.95 - x} = \frac{(2.1 \times 10^{-2})^2}{0.95 - 2.1 \times 10^{-2}} = 4.7 \times 10^{-4}$$

$$\text{p}K_b = -\log_{10}(4.7 \times 10^{-4}) = 3.32$$

13. Hydrofluoric acid ( $\text{HF}_{(aq)}$ ) is formed when gas-phase hydrogen fluoride ( $\text{HF}_{(g)}$ ) is dissolved in water according to the reaction



The  $\text{p}K_a$  of hydrofluoric acid is 3.18. Calculate the pH of a solution that forms when 0.25 grams of gas-phase HF are dissolved in water to make 1.2 L of solution. Approximations do not work for this problem.

**Answer:**

$$K_a = 10^{-3.18} = 6.6 \times 10^{-4}$$

Initially

$$[\text{HF}] = \frac{(0.25 \text{ g}) \left( \frac{\text{mol}}{(19.0 + 1.0) \text{ g}} \right)}{1.2 \text{ L}} = 1.0 \times 10^{-2} \text{ M}$$

	[HF]	[H <sub>3</sub> O <sup>+</sup> ]	[F <sup>-</sup> ]
initial	$1.0 \times 10^{-2} \text{ M}$	0 M	0 M
change	$-y \text{ M}$	$y \text{ M}$	$y \text{ M}$
equilibrium	$(1.0 \times 10^{-2} - y) \text{ M}$	$y \text{ M}$	$y \text{ M}$

$$6.6 \times 10^{-4} = \frac{[\text{F}^-][\text{H}_3\text{O}^+]}{[\text{HF}]} = \frac{y^2}{1.0 \times 10^{-2} - y}$$

$$y^2 + 6.6 \times 10^{-4}y - 6.6 \times 10^{-6} = 0$$

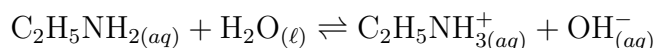
Ignoring the negative solution

$$y = \frac{-6.6 \times 10^{-4} + [(6.6 \times 10^{-4})^2 + 4(6.6 \times 10^{-6})]^{1/2}}{2} = 2.2 \times 10^{-3}$$

$$[\text{H}_3\text{O}^+] = 2.2 \times 10^{-3} \text{ M} \quad \text{pH} = -\log_{10}(2.2 \times 10^{-3}) = 2.66$$

14. Ethylamine,  $\text{C}_2\text{H}_5\text{NH}_2$ , is a weak base in aqueous solution. The pH of a 1.0 M aqueous ethylamine solution is measured to be 12.32. Calculate the  $\text{p}K_b$  of ethylamine.

**Answer:**



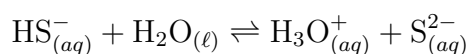
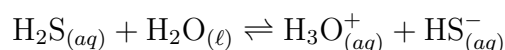
$$\text{pOH} = 14.00 - \text{pH} = 1.68 \quad [\text{OH}^-] = 10^{-1.68} = 2.1 \times 10^{-2}$$

	[C <sub>2</sub> H <sub>5</sub> NH <sub>2</sub> ]	[C <sub>2</sub> H <sub>5</sub> NH <sub>3</sub> <sup>+</sup> ]	[OH <sup>-</sup> ]
initial	1.0 M	0 M	0 M
change	$-2.1 \times 10^{-2} \text{ M}$	$2.1 \times 10^{-2} \text{ M}$	$2.1 \times 10^{-2} \text{ M}$
equilibrium	0.98 M	$2.1 \times 10^{-2} \text{ M}$	$2.1 \times 10^{-2} \text{ M}$

$$K_b = \frac{[\text{C}_2\text{H}_5\text{NH}_3^+][\text{OH}^-]}{[\text{C}_2\text{H}_5\text{NH}_2]} = \frac{(2.1 \times 10^{-2})^2}{0.98} = 4.5 \times 10^{-4}$$

$$\text{p}K_b = -\log_{10}(4.5 \times 10^{-4}) = 3.35$$

15. The two values for the  $\text{p}K_a$  of hydrosulfuric acid according to the reactions



are respectively 7.00 and 19.0. Calculate the concentrations of all species present in a 0.0100 and 0.00100 M aqueous solution of hydrosulfuric acid.

**Answer:**

$$\frac{[\text{H}_3\text{O}^+][\text{HS}^-]}{[\text{H}_2\text{S}]} = 1.00 \times 10^{-7}$$

$$\frac{[\text{H}_3\text{O}^+][\text{S}^{2-}]}{[\text{HS}^-]} = 1.00 \times 10^{-19}$$

(a) 0.0100 M

	$[\text{H}_2\text{S}]$	$[\text{HS}^-]$	$[\text{H}_3\text{O}^+]$
initial	0.0100 M	0 M	0 M
change	$-x$	$x$	$x$
equilibrium	$(0.0100 - x)$ M	$x$ M	$x$ M

$$\frac{x^2}{0.0100 - x} = 1.00 \times 10^{-7} \cong \frac{x^2}{0.0100}$$

$$x^2 = 1.00 \times 10^{-9} \quad x = [\text{H}_3\text{O}^+] = [\text{HS}^-] = \sqrt{1.00 \times 10^{-9}} = 3.16 \times 10^{-5} \text{ M} \quad [\text{H}_2\text{S}] \cong 0.0100 \text{ M}$$

	$[\text{HS}^-]$	$[\text{S}^{2-}]$	$[\text{H}_3\text{O}^+]$
initial	$3.16 \times 10^{-5} \text{ M}$	0 M	$3.16 \times 10^{-5} \text{ M}$
change	$-x$	$x$	$x$
equilibrium	$(3.16 \times 10^{-5} - x)$ M	$x$ M	$(3.16 \times 10^{-5} + x)$ M

$$\frac{x(3.16 \times 10^{-5} + x)}{3.16 \times 10^{-5} - x} \approx x = [\text{S}^{2-}] = 1.00 \times 10^{-19}$$

(b) 0.00100 M. In this case the approximations used in solving the previous case are a bit large, and full solution to the quadratic equations are needed.



	[H <sub>2</sub> S]	[HS <sup>-</sup> ]	[H <sub>3</sub> O <sup>+</sup> ]
initial	0.00100 M	0 M	0 M
change	- $x$	$x$	$x$
equilibrium	(0.00100 - $x$ ) M	$x$ M	$x$ M

$$\frac{x^2}{0.00100 - x} = 1.00 \times 10^{-7}$$

$$x^2 + 1.00 \times 10^{-7}x - 1.00 \times 10^{-10} = 0$$

$$x = \frac{-1.00 \times 10^{-7} \pm \sqrt{(1.00 \times 10^{-7})^2 + 4 \times 1.00 \times 10^{-10}}}{2}$$

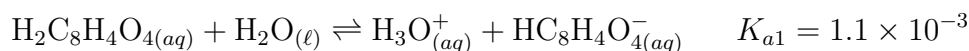
Keeping only the positive solution

$$x = [\text{H}_3\text{O}^+] = [\text{HS}^-] = 9.95 \times 10^{-6} \quad [\text{H}_2\text{S}] \cong 0.00100 \text{ M}$$

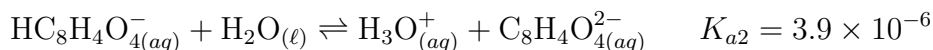
	[HS <sup>-</sup> ]	[S <sup>2-</sup> ]	[H <sub>3</sub> O <sup>+</sup> ]
initial	$9.95 \times 10^{-6} \text{ M}$	0 M	$9.95 \times 10^{-6} \text{ M}$
change	- $x$	$x$	$x$
equilibrium	$(9.95 \times 10^{-6} - x) \text{ M}$	$x \text{ M}$	$(9.95 \times 10^{-6} + x) \text{ M}$

$$\frac{x(9.95 \times 10^{-6} + x)}{9.95 \times 10^{-6} - x} \approx x = [\text{S}^{2-}] = 1.00 \times 10^{-19}$$

16. When phthalic acid combines with water, two equilibrium reactions occur



and



Calculate the pH and concentrations of all species present in a 0.010 M phthalic acid solution.

**Answer:**

$$K_{a1} = \frac{[\text{H}_3\text{O}^+][\text{HC}_8\text{H}_4\text{O}_4^-]}{[\text{H}_2\text{C}_8\text{H}_4\text{O}_4]} = 1.1 \times 10^{-3}$$

	[H <sub>2</sub> C <sub>8</sub> H <sub>4</sub> O <sub>4</sub> ]	[HC <sub>8</sub> H <sub>4</sub> O <sub>4</sub> <sup>-</sup> ]	[H <sub>3</sub> O <sup>+</sup> ]
initial	0.010 M	0 M	0 M
change	- $x$	$x$	$x$
equilibrium	(0.010 - $x$ ) M	$x \text{ M}$	$x \text{ M}$

$$\frac{x^2}{0.010 - x} = 1.1 \times 10^{-3}$$

$$x^2 + 1.1 \times 10^{-3}x - 1.1 \times 10^{-5} = 0$$

$$x = \frac{-1.1 \times 10^{-3} \pm \sqrt{(1.1 \times 10^{-3})^2 + 4 \times 1.1 \times 10^{-5}}}{2}$$

$$x = [\text{H}_3\text{O}^+] = [\text{HC}_8\text{H}_4\text{O}_4^-] = 2.8 \times 10^{-3} \text{ M}$$

$$K_{a2} = \frac{[\text{H}_3\text{O}^+][\text{C}_8\text{H}_4\text{O}_4^{2-}]}{[\text{HC}_8\text{H}_4\text{O}_4^-]} = 3.9 \times 10^{-6}$$

	$[\text{HC}_8\text{H}_4\text{O}_4^-]$	$[\text{C}_8\text{H}_4\text{O}_4^{2-}]$	$[\text{H}_3\text{O}^+]$
initial	$2.8 \times 10^{-3} \text{ M}$	0 M	$2.8 \times 10^{-3} \text{ M}$
change	$-x$	$x$	$x$
equilibrium	$(2.8 \times 10^{-3} - x) \text{ M}$	$x \text{ M}$	$(2.8 \times 10^{-3} + x) \text{ M}$

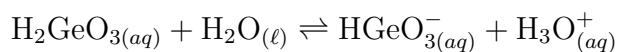
$$\frac{x(2.8 \times 10^{-3} + x)}{2.8 \times 10^{-3} - x} \approx x = 3.9 \times 10^{-6} = [\text{C}_8\text{H}_4\text{O}_4^{2-}]$$

$$\text{pH} = -\log_{10} 2.8 \times 10^{-3} = 2.55$$

17. Germanic acid ( $\text{H}_2\text{GeO}_3$ ) is a diprotic acid with  $\text{p}K_a$ 's for the two protons in aqueous solution given by  $\text{p}K_{a,1} = 9.01$  and  $\text{p}K_{a,2} = 12.30$ . Calculate the equilibrium concentrations of  $\text{HGeO}_3^-$ ,  $\text{GeO}_3^{2-}$  and  $\text{H}_3\text{O}^+$  and the pH for a 1.50 M aqueous solution of germanic acid. Approximations work for this problem.

**Answer:**

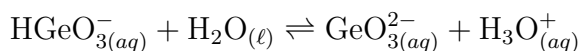
$$K_{a1} = 10^{-9.01} = 9.8 \times 10^{-10} \quad K_{a2} = 10^{-12.30} = 5.0 \times 10^{-13}$$



	$[\text{H}_2\text{GeO}_3]$	$[\text{HGeO}_3^-]$	$[\text{H}_3\text{O}^+]$
initial	1.50 M	0 M	0 M
change	$-y$	$y$	$y$
equilibrium	$(1.50 - y) \text{ M}$	$y \text{ M}$	$y \text{ M}$

$$9.8 \times 10^{-10} = \frac{[\text{HGeO}_3^-][\text{H}_3\text{O}^+]}{[\text{H}_2\text{GeO}_3]} = \frac{y^2}{1.50 - y} \approx \frac{y^2}{1.50}$$

$$y = [\text{H}_3\text{O}^+] = [\text{HGeO}_3^-] = 3.8 \times 10^{-5} \text{ M} \quad \text{pH} = -\log_{10}(3.8 \times 10^{-5}) = 4.42$$

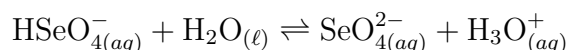


	$[\text{HGeO}_3^-]$	$[\text{GeO}_3^{2-}]$	$[\text{H}_3\text{O}^+]$
initial	$3.8 \times 10^{-5} \text{ M}$	0 M	$3.8 \times 10^{-5} \text{ M}$
change	$-y$	$y$	$y$
equilibrium	$(3.8 \times 10^{-5} - y) \text{ M}$	$y \text{ M}$	$(3.8 \times 10^{-5} + y) \text{ M}$

$$5.0 \times 10^{-13} = \frac{[\text{GeO}_3^{2-}][\text{H}_3\text{O}^+]}{[\text{HGeO}_3^-]} = \frac{y(3.8 \times 10^{-5} + y)}{3.8 \times 10^{-5} - y} \approx y$$

$$[\text{GeO}_3^{2-}] = 5.0 \times 10^{-13} \text{ M}$$

18. Selenic acid ( $\text{H}_2\text{SeO}_4$ ) is a diprotic acid. Like sulfuric acid, the first proton is essentially 100% ionized in aqueous solution. The second proton dissociates according to the reaction



with a  $\text{p}K_a = 1.92$ . Consider a 0.10 M solution of selenic acid. Calculate the equilibrium concentrations of  $\text{HSeO}_4^-$  and  $\text{SeO}_4^{2-}$ , and the pH of the solution at equilibrium. Approximations work for the dissociation equilibrium.

**Answer:**

	$[\text{HSeO}_4^-]$	$[\text{SeO}_4^{2-}]$	$[\text{H}_3\text{O}^+]$
initial	0.10 M	0 M	0.10 M
change	$-y$	$y$	$y$
equilibrium	$(0.10 - y) \text{ M}$	$y \text{ M}$	$(0.10 + y) \text{ M}$

$$K_a = 10^{-1.92} = 1.2 \times 10^{-2} = \frac{[\text{SeO}_4^{2-}][\text{H}_3\text{O}^+]}{[\text{HSeO}_4^-]}$$

$$= \frac{y(0.10 + y)}{0.10 - y} \approx y$$

$$[\text{SeO}_4^{2-}] = 1.2 \times 10^{-2}$$

$$[\text{HSeO}_4^-] = 0.10 - 1.2 \times 10^{-2} = 0.10 \text{ (to two significant figures)}$$

$$[\text{H}_3\text{O}^+] = 0.10 + 1.2 \times 10^{-2} = 0.11 \text{ M}$$

$$\text{pH} = -\log_{10}(0.11) = 0.96$$