

Chemistry 30

Unit 5: Acids & Bases

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Assignment 2 – K_a , K_b , K_w and pH

- 3 1. Given the following balanced ionization reactions for the following weak acids and bases, write the K_a or K_b expressions for each.

a. ascorbic acid: $\text{HC}_6\text{H}_7\text{O}_6(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{C}_6\text{H}_7\text{O}_6^-(\text{aq})$	$K_a = \frac{[\text{H}^+][\text{C}_6\text{H}_7\text{O}_6^-]}{[\text{HC}_6\text{H}_7\text{O}_6]}$
b. boric acid: $\text{H}_3\text{BO}_3(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{H}_2\text{BO}_3^-(\text{aq})$	$K_a = \frac{[\text{H}^+][\text{H}_2\text{BO}_3^-]}{[\text{H}_3\text{BO}_3]}$
c. methyl amine: $\text{CH}_3\text{NH}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{CH}_3\text{NH}_3^+(\text{aq}) + \text{OH}^-(\text{aq})$	$K_b = \frac{[\text{CH}_3\text{NH}_3^+][\text{OH}^-]}{[\text{CH}_3\text{NH}_2]}$

- 2 2. Calculate $[\text{OH}^-]$ in a solution containing 100.0 g of potassium hydroxide in 2.50 L solution. Potassium hydroxide is a strong base.

The molar mass of KOH is $56.1 \text{ g}\cdot\text{mol}^{-1}$

$$[\text{KOH}] = \frac{\text{mol}}{56.1\text{g}} \times \frac{100.0\text{g}}{1} \times \frac{1}{2.50\text{L}} = \frac{0.713\text{mol}}{\text{L}} = 0.713\text{M}$$

$$\text{KOH}(\text{aq}) \rightarrow \text{K}^+(\text{aq}) + \text{OH}^-(\text{aq})$$

Since KOH is a strong base, based on the balanced equation $[\text{OH}^-] = [\text{KOH}] = 0.713 \text{ M}$

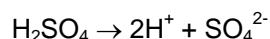
- 2 3. A solution is prepared in which 0.600 mole of hydrogen chloride is dissolved in enough water to make 5.80 L. Calculate the concentration of hydrogen ions in this solution.

$$[\text{HCl}] = \frac{0.600\text{mol}}{5.80\text{L}} = \frac{0.103\text{mol}}{\text{L}} = 0.103\text{M}$$

$$\text{HCl}(\text{aq}) \rightarrow \text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq})$$

Since HCl is a strong acid, based on the balanced equation $[\text{H}^+] = [\text{HCl}] = 0.103 \text{ M}$

- 2 4. A solution is prepared that contains 0.0445 mole of sulfuric acid in a total solution volume of 12.1 L. Sulfuric acid typically undergoes complete ionization according to the equation:

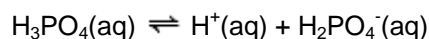


Calculate $[\text{H}^+]$. Sulfuric acid is a strong acid.

$$[\text{H}_2\text{SO}_4] = \frac{0.0445 \text{ mol}}{12.1 \text{ L}} = \frac{3.68 \times 10^{-3} \text{ mol}}{\text{L}} = 3.68 \times 10^{-3} \text{ M}$$

Since H_2SO_4 is a strong acid, based on the balanced equation $[\text{H}^+] = 2 \times [\text{H}_2\text{SO}_4] = 7.36 \times 10^{-3} \text{ M}$

- 4 5. Phosphoric acid is a **weak** acid that undergoes the following ionization reaction:



If there are 1.32×10^{-2} mole of phosphoric acid present in 875 mL of solution, calculate the concentration of hydrogen ions, H^+ , in solution. K_a for phosphoric acid is 7.0×10^{-3} .

Begin by calculating $[\text{H}_3\text{PO}_4]$. Then use K_a to determine $[\text{H}^+]$.

$$[\text{H}_3\text{PO}_4] = \frac{1.32 \times 10^{-2} \text{ mol}}{0.875 \text{ L}} = \frac{1.51 \times 10^{-2} \text{ mol}}{\text{L}} = 1.51 \times 10^{-2} \text{ M}$$

Since H_3PO_4 is a **weak** acid, we must find $[\text{H}^+]$ using K_a for this acid:

$$K_a = \frac{[\text{H}^+][\text{H}_2\text{PO}_4^-]}{[\text{H}_3\text{PO}_4]} \quad \text{point} \quad 7.0 \times 10^{-3} = \frac{(x)(x)}{(1.51 \times 10^{-2})} \quad \text{point} \quad x^2 = (7.0 \times 10^{-3})(1.51 \times 10^{-2})$$

$$x^2 = 1.056 \times 10^{-4}$$

$$x = 1.02 \times 10^{-2}$$

Answer: $[\text{H}^+] = 1.02 \times 10^{-2} \text{ M}$

- 6 Determine the pH of each of the following solutions, and tell whether the solution is acidic or basic.

		Acid or Base?
a)	$[\text{H}^+] = 1.0 \times 10^{-3} \text{ M}$ pH = 3	Acid
6 b)	$[\text{H}^+] = 2.5 \times 10^{-5} \text{ M}$ pH = 4.6	Acid
c)	$[\text{OH}^-] = 0.01 \text{ M}$ pH = 12	Base

- 4 7. Calculate both $[H^+]$ and $[OH^-]$ for the following solutions. All are either strong acids or strong bases. Be sure to clearly identify all answers.

a) 2.5 M NaOH

$$[OH^-] = [NaOH] = 2.5 \text{ M}$$

$$[H^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{2.5} = 4.0 \times 10^{-15} \text{ M}$$

b) 0.045 M HCl

$$[H^+] = [HCl] = 0.045 \text{ M}$$

$$[OH^-] = \frac{K_w}{[H^+]} = \frac{1.0 \times 10^{-14}}{0.045} = 2.2 \times 10^{-13} \text{ M}$$

- 3 8. Calculate the pH of a 0.1 M solution of sodium hydroxide, NaOH, a strong base.

Since NaOH is a strong base, and based on the balanced equation $NaOH \rightarrow Na^+ + OH^-$

$$[OH^-] = [NaOH] = 0.1 \text{ M}$$

Use K_w to find $[H^+]$:
$$[H^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{0.1} = 1.0 \times 10^{-13} \text{ M}$$

$$pH = -\log[H^+] = -\log(1.0 \times 10^{-13}) = 13$$

- 9 a) Determine the concentration of hydrogen ions, $[H^+]$ in a solution whose pH is 5.17.

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$$[H^+] = \text{antilog}(-pH) = \text{antilog}(-5.17) = 6.8 \times 10^{-6} \text{ M}$$

- b) Calculate the hydroxide ion concentration, $[OH^-]$, for this solution.

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$$[OH^-] = \frac{K_w}{[H^+]} = \frac{1.0 \times 10^{-14}}{6.8 \times 10^{-6}} = 1.5 \times 10^{-9} \text{ M}$$

10. Determine $[H_3O^+]$ in a solution whose pH = 9.22. (Hint: $[H_3O^+] = [H^+]$)

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$$[H_3O^+] = \text{antilog}(-pH) = \text{antilog}(-9.22) = 6.03 \times 10^{-10} \text{ M}$$

- 5 11. A 2.67 g sample of hydrogen fluoride gas (HF) is dissolved in sufficient water to make 1.05 L of solution at 25°C to form an acidic solution. Hydrogen fluoride is a weak acid with $K_a = 6.6 \times 10^{-4}$.

Calculate the pH of this solution.

Begin by calculating [HF]. Then use K_a to determine $[H^+]$. Finally convert $[H^+]$ to pH.

The molar mass of HF is $20.0 \text{ g} \cdot \text{mol}^{-1}$

$$[\text{HF}] = \frac{\text{mol}}{20.0\text{g}} \times \frac{2.67\text{g}}{1} \times \frac{1}{1.05\text{L}} = \frac{0.127\text{mol}}{\text{L}} = 0.127\text{M}$$

Since HF is a **weak** acid, we must find $[H^+]$ using K_a for this acid:

$$K_a = \frac{[H^+][F^-]}{[\text{HF}]} \quad \text{☞} \quad 6.6 \times 10^{-4} = \frac{(x)(x)}{(0.127)} \quad \text{☞} \quad x^2 = (6.6 \times 10^{-4})(0.127)$$

$$x^2 = 8.39 \times 10^{-5}$$

$$x = 9.16 \times 10^{-3}$$

$$[H^+] = 9.16 \times 10^{-3} \text{ M}$$

$$\text{pH} = -\log[H^+] = -\log(9.16 \times 10^{-3}) = \mathbf{2.04} \quad \mathbf{ANSWER}$$

- 5 12. The formula for ascorbic acid, better known as Vitamin C, is $\text{HC}_6\text{H}_7\text{O}_6$. K_a for ascorbic acid is 8.00×10^{-5} . Determine the pH of a solution prepared by dissolving a 500.0 mg vitamin C tablet in enough water to make 200.0 mL of solution.

The molar mass of $\text{HC}_6\text{H}_7\text{O}_6$ is $176.0 \text{ g} \cdot \text{mol}^{-1}$

$$[\text{HF}] = \frac{\text{mol}}{176.0\text{g}} \times \frac{0.500\text{g}}{1} \times \frac{1}{0.200\text{L}} = \frac{0.142\text{mol}}{\text{L}} = 0.142\text{M}$$

Since $\text{HC}_6\text{H}_7\text{O}_6$ is a **weak** acid, we must find $[H^+]$ using K_a for this acid:

$$K_a = \frac{[H^+][\text{C}_6\text{H}_7\text{O}_6^-]}{[\text{HC}_6\text{H}_7\text{O}_6]} \quad \text{☞} \quad 8.0 \times 10^{-5} = \frac{(x)(x)}{(0.142)} \quad \text{☞} \quad x^2 = (8.0 \times 10^{-5})(0.142)$$

$$x^2 = 1.14 \times 10^{-6}$$

$$x = 1.07 \times 10^{-3}$$

$$[H^+] = 1.07 \times 10^{-3} \text{ M}$$

$$\text{pH} = -\log[H^+] = -\log(1.07 \times 10^{-3}) = \mathbf{3.0} \quad \mathbf{ANSWER}$$