

K_a, K_b, K_{sp} Constant Practice

Name: SOLUTIONS

- 1) An acid has a concentration of 0.17 mol/L, it dissociates in water with an ionization percentage of 8.7%. What is the pH of the solution?

$$\textcircled{1} \% = \frac{[\text{H}_3\text{O}^+]}{[\text{HA}]} \times 100$$

$$8.7 = \frac{[\text{H}_3\text{O}^+]}{0.17} \times 100$$

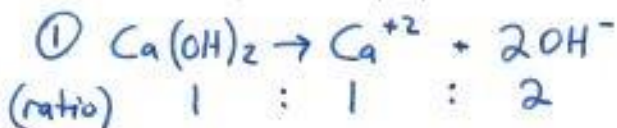
$$[\text{H}_3\text{O}^+] = 0.0148 \text{ mol/L}$$

$$\textcircled{2} \text{pH} = -\log[\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log(0.0148)$$

$$\text{pH} = \underline{\underline{1.83}}$$

- 2) Find the concentration of hydronium ions [H₃O⁺] and concentration of hydroxide ions [OH⁻] for a calcium hydroxide Ca(OH)₂ solution with a concentration of 0.012 mol/L. (Assume 100% dissociation)



if, [Ca(OH)₂] = 0.012 mol/L

then, [Ca⁺²] = 0.012 mol/L

and, [OH⁻] = 2 × 0.012 mol/L = 0.024 mol/L

$$\textcircled{2} K_{\text{water}} = [\text{H}_3\text{O}^+] \cdot [\text{OH}^-]$$

$$10^{-14} = [\text{H}_3\text{O}^+] \cdot (0.024)$$

$$\frac{10^{-14}}{0.024} = [\text{H}_3\text{O}^+]$$

$$[\text{H}_3\text{O}^+] = \underline{\underline{4.2 \times 10^{-13} \text{ mol/L}}}$$

- 3) A solution of cyanic acid (HOCN) with a concentration of 0.01 mol/L has an ionization percentage of 17.04%.

a) What is the pH of this acid solution?

b) What is the value of the acidity constant of the cyanic acid?

$$\textcircled{a} \% = \frac{[\text{H}_3\text{O}^+]}{[\text{HOCN}]} \times 100$$

$$17.04 = \frac{[\text{H}_3\text{O}^+]}{0.01} \times 100$$

$$[\text{H}_3\text{O}^+] = 0.001704 \text{ mol/L}$$

$$\textcircled{b} \text{pH} = -\log[\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log(0.001704)$$

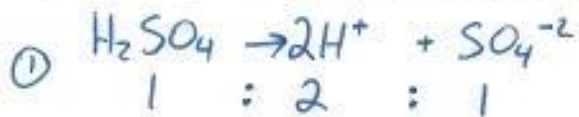
$$\text{pH} = \underline{\underline{2.77}}$$

b)

	HOCN(aq) + H ₂ O(l) ⇌	H ₃ O ⁺ (aq)	+ OCN ⁻ (aq)
T	1	1	1
I	0.01	0	0
C	-0.001704	+0.001704	+0.001704
E	0.008296	0.001704	0.001704

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{OCN}^-]}{[\text{HOCN}]} = \frac{(0.001704)(0.001704)}{0.008296} = K_a = \underline{\underline{0.00035}}$$

4) Sulfuric acid is considered to be a strong acid. It dissociated completely when in solution. If a sample of sulfuric acid initially has a concentration of 0.25 mol/L. What is the pH and the pOH of the solution.



$$\textcircled{3} \quad \text{pH} = -\log[\text{H}^+]$$

$$\text{pH} = -\log(0.5)$$

$$\text{pH} = 0.3$$

$$\textcircled{4} \quad \text{pH} + \text{pOH} = 14$$

$$0.3 + \text{pOH} = 14$$

$$\text{pOH} = 13.7$$

$$\textcircled{2} \quad [\text{H}^+] = 2 \cdot (0.25 \text{ mol/L})$$

$$[\text{H}^+] = 0.5 \text{ mol/L}$$

5) A chlorous acid (HClO_2) solution has a concentration of 0.2 mol/L and an acidity constant value of 1.1×10^{-2} . What is the pH of this acid solution? What is the ionization percentage of this acid at this concentration?

$$\textcircled{1} \quad \text{HClO}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{ClO}_2^-(\text{aq})$$

I	1	1	1
I	0.2	0	0
C	$-x$	$+x$	$+x$
E	$(0.2-x)$	x	x

$$\textcircled{2} \quad K_a = \frac{[\text{H}_3\text{O}^+][\text{ClO}_2^-]}{[\text{HClO}_2]}$$

$$1.1 \times 10^{-2} = \frac{x \cdot x}{(0.2-x)}$$

$$x^2 + 0.011x - 0.0022 = 0$$

$$x = 0.0417 \text{ or } -0.0527$$

↓

$$[\text{H}_3\text{O}^+] = 0.0417 \text{ mol/L}$$

$$\textcircled{3} \quad \text{pH} = -\log[\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log(0.0417)$$

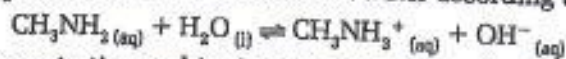
$$\text{pH} = 1.38$$

$$\textcircled{4} \quad \text{ion. \%} = \frac{[\text{H}_3\text{O}^+]}{[\text{HClO}_2]} \times 100$$

$$\text{ion. \%} = \frac{0.0417}{0.2} \times 100$$

$$= 20.8\%$$

- 6) Methanamine (CH_3NH_2) is a base that dissociates in water according to the following equation:



What are the initial concentration and ionization percentage of a methanamine solution with a pH of 9.16, if the value of its basicity constant is 4.60×10^{-4} ?

$$\textcircled{1} \text{ pH} + \text{pOH} = 14$$

$$9.16 + \text{pOH} = 14$$

$$\underline{\underline{\text{pOH} = 4.84}}$$

$$\textcircled{2} [\text{OH}^-] = 10^{-\text{pOH}}$$

$$[\text{OH}^-] = 10^{-4.84}$$

$$\underline{\underline{[\text{OH}^-] = 0.0000145 \text{ mol/L}}}$$



T	1	1	1
I	y	0	0
C	-0.0000145	+0.0000145	+0.0000145
E	y - 0.0000145	0.0000145	0.0000145

$$\textcircled{4} K_B = \frac{[\text{CH}_3\text{NH}_3^+][\text{OH}^-]}{[\text{CH}_3\text{NH}_2]}$$

$$4.6 \times 10^{-4} = \frac{(0.0000145)(0.0000145)}{(y - 0.0000145)}$$

solve for y

$$\underline{\underline{y = 1.49 \times 10^{-5} \text{ mol/L}}}$$

$$\hookrightarrow y = [\text{CH}_3\text{NH}_2] = 1.49 \times 10^{-5} \text{ mol/L}$$

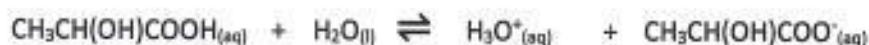
(initial)

$$\textcircled{5} \text{ion \%} = \frac{[\text{OH}^-]}{[\text{CH}_3\text{NH}_2]} \times 100$$

$$\text{ion \%} = \frac{0.0000145}{1.49 \times 10^{-5}} \times 100$$

$$\underline{\underline{= 97\%}}$$

- 7) Lactic acid ($\text{CH}_3\text{CH}(\text{OH})\text{COOH}$) is found in dairy and milk products. If a lactic acid solution at 0.2 mol/L dissociates as follows:



- a) What is the acidity constant (K_a) if the pH is 2.285?
 b) What is the ionization percentage of lactic acid?

a) ① $[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$
 $[\text{H}_3\text{O}^+] = 10^{-2.285}$
 $[\text{H}_3\text{O}^+] = \underline{\underline{0.0052 \text{ mol/L}}}$



T	1	1	1
I	0.2	0	0
C	-0.0052	+0.0052	+0.0052
E	0.1948	0.0052	0.0052

③ $K_a = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{CH}(\text{OH})\text{COO}^-]}{[\text{CH}_3\text{CH}(\text{OH})\text{COOH}]}$

$$K_a = \frac{(0.0052)(0.0052)}{0.1948}$$

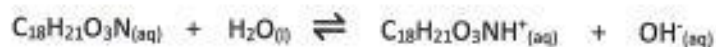
$$\underline{\underline{K_a = 1.39 \times 10^{-4}}}$$

b) $\% \text{ ion} = \frac{[\text{H}_3\text{O}^+]}{[\text{CH}_3\text{CH}(\text{OH})\text{COOH}]} \times 100$

$$\% \text{ ion} = \frac{0.0052}{0.2} \times 100$$

$$\underline{\underline{= 2.6\%}}$$

- 8) Codeine ($C_{18}H_{21}O_3N$) is a weak base that is used to treat pain. Its dissociation is represented by the following:



If a codeine solution has a concentration of 1.33 mol/L at a pH of 11.2, what is the value of the basicity constant (K_b)

$$\begin{aligned} \textcircled{1} \quad pH + pOH &= 14 \\ 11.2 + pOH &= 14 \\ \underline{\underline{pOH &= 2.8}} \end{aligned}$$

$$\begin{aligned} \textcircled{2} \quad [OH^-] &= 10^{-pOH} \\ [OH^-] &= 10^{-2.8} \\ \underline{\underline{[OH^-] &= 0.00158 \text{ mol/L}}} \end{aligned}$$



T	1	1	1
I	1.33	0	0
C	-0.00158	+0.00158	+0.00158
E	1.32842	0.00158	0.00158

$$\textcircled{4} \quad K_b = \frac{[C_{18}H_{21}O_3NH^+][OH^-]}{[C_{18}H_{21}O_3N]} = \frac{(0.00158)(0.00158)}{(1.32842)}$$

$$\underline{\underline{K_b = 1.88 \times 10^{-6}}}$$

- 9) Calcium oxalate (CaC_2O_4) is one of the main solids that make up the crystal structure of kidney stones. If calcium oxalate has a molar solubility of 0.00052 mol/L.

- Write out the dissociation equation of calcium oxalate.
- What is the value of solubility constant (K_{sp}).



b)

$$[Ca^{+2}] = [C_2O_4^{-2}] = \text{CaC}_2O_4 \text{ (molar solubility)}$$

$$K_{sp} = [Ca^{+2}] \cdot [C_2O_4^{-2}]$$

$$K_{sp} = (0.00052)(0.00052) \quad \underline{\underline{K_{sp} = 2.7 \times 10^{-7}}}$$

- 10) Butanoic acid $C_4H_8O_2$ gives parmesan cheese its unique smell. Calculate the pH of a solution of butanoic acid with an initial concentration of 1×10^{-2} mol/L if the value of the acidity constant is 1.51×10^{-5} .

$$C_4H_8O_2 \rightleftharpoons H_3O^+ + C_4H_7O_2^-$$

Theory	1	1	1
Initial	0.01	0	0
Change	-x	+x	+x
Equilibrium	$0.01 - x$	x	x

Substitute:

$$K_a = \frac{[H_3O^+][C_4H_7O_2^-]}{[C_4H_8O_2]}$$

$$1.51 \times 10^{-5} = \frac{x^2}{0.01 - x}$$

$$1.51 \times 10^{-5}(0.01 - x) = x^2$$

$$1.51 \times 10^{-7} - 1.51 \times 10^{-5}x = x^2$$

$$x^2 + 1.51 \times 10^{-5}x - 1.51 \times 10^{-7} = 0$$

$$x = \underline{0.000381} \text{ or } \underline{-0.000396}$$

✓
✗
 (valid) (invalid)

Substitute valid value of x into equilibrium concentrations:

$$[C_4H_8O_2] = x = 0.000381 \text{ mol/L}$$

$$[H_3O^+] = x = 0.000381 \text{ mol/L}$$

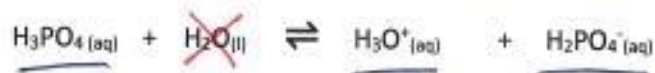
$$[C_4H_7O_2^-] = 0.01 - x = 0.00962 \text{ mol/L}$$

Solve for pH:

$$pH = -\log(3.81 \times 10^{-4})$$

$$pH = 3.42$$

13) Consider the following equation which represents the first dissolution step of phosphoric acid (H_3PO_4):



If a phosphoric acid solution has a concentration of 0.1 mol/L with an acidity constant (K_a) of 7.5×10^{-3} what is the pH of the solution at equilibrium?

①

T			
I	0.1	0	0
C	-x	+x	+x
E	(0.1-x)	x	x

② $K_a = \frac{[\text{H}_3\text{O}^+][\text{H}_2\text{PO}_4^-]}{[\text{H}_3\text{PO}_4]} \quad 7.5 \times 10^{-3} = \frac{x \cdot x}{(0.1-x)}$

$$7.5 \times 10^{-3}(0.1-x) = x^2$$

$$7.5 \times 10^{-4} - 7.5 \times 10^{-3}x = x^2$$

$$x^2 + 7.5 \times 10^{-3}x - 7.5 \times 10^{-4} = 0$$

solve for x values

$$x = 0.02389 \quad \& \quad x = -0.03139$$

③ $\text{pH} = -\log[\text{H}_3\text{O}^+]$

$$\text{pH} = -\log[x]$$

$$\text{pH} = -\log(0.02389)$$

$$\text{pH} = 1.62$$

- 14) A solution of hydrofluoric acid (HF) has an initial concentration of 0.01 mol/L. What is the pH of the is solution if the value of the acidity constant is 6.3×10^{-4} .



	HF	H ₃ O ⁺	F ⁻
Theory	1	1	1
Initial	0.01	0	0
Change	-x	+x	+x
Equilibrium	0.01 - x	x	x

Substitute:

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

$$K_a = \frac{(x)(x)}{0.01 - x}$$

$$6.3 \times 10^{-4}(0.01 - x) = x^2$$

$$6.3 \times 10^{-6} - 6.3 \times 10^{-4}x = x^2$$

$$x^2 + 6.3 \times 10^{-4}x - 6.3 \times 10^{-6} = 0$$

$$x = \frac{0.00221}{\checkmark} \text{ or } \frac{-0.00284}{\times}$$

(valid) (invalid)

Substitute valid value of x into equilibrium concentrations:

$$[\text{C}_2\text{H}_3\text{O}_2] = x = 0.00221 \text{ mol/L}$$

$$[\text{H}_3\text{O}^+] = x = 0.00221 \text{ mol/L}$$

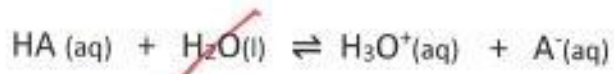
$$[\text{C}_4\text{H}_7\text{O}_2] = 0.01 - x = 0.00779 \text{ mol/L}$$

Solve for pH:

$$\text{pH} = -\log(0.00221)$$

$$\text{pH} = 2.66$$

15) 5% of a 0.1 mol/L solution of a weak acid dissociates. Calculate the value of the acidity constant if the reaction is represented by:



$$\textcircled{1} \quad \text{Ionization \%} = \frac{[\text{H}_3\text{O}^+]}{[\text{HA}]} \times 100$$

$$5\% = \frac{[\text{H}_3\text{O}^+]}{0.1} \times 100$$

$$\underline{[\text{H}_3\text{O}^+] = 0.005 \text{ mol/L}}$$

$\textcircled{2}$

T	1	1	1
I	0.1	0	0
C	-0.005	+0.005	+0.005
E	0.095	0.005	0.005

$$\textcircled{3} \quad K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} = \frac{(0.005)(0.005)}{0.095}$$

$$\underline{\underline{K_a = 0.00026}}$$

16) For each of the following weak bases, write the equation of the chemical equilibrium and the equation of the basicity equilibrium constant.

a) Aqueous cyanide (CN⁻)



$$K_b = \frac{[\text{HCN}][\text{OH}^{-}]}{[\text{CN}^{-}]}$$

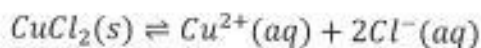
b) Aqueous sulphate (SO₄²⁻)



$$K_b = \frac{[\text{HSO}_4^{-}][\text{OH}^{-}]}{[\text{SO}_4^{2-}]}$$

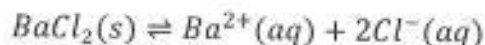
17) Write the balanced equation representing the dissociation of each of the following solids when dissolved in water. Also write the expressions for the solubility product constant (K_{sp}).

a) Copper(II) chloride — CuCl₂



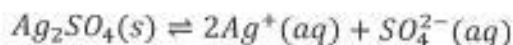
$$K_{sp} = [\text{Cu}^{2+}][\text{Cl}^{-}]^2$$

b) Barium chloride — BaCl₂



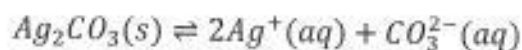
$$K_{sp} = [\text{Ba}^{2+}][\text{Cl}^{-}]^2$$

c) Silver sulfate — Ag₂SO₄



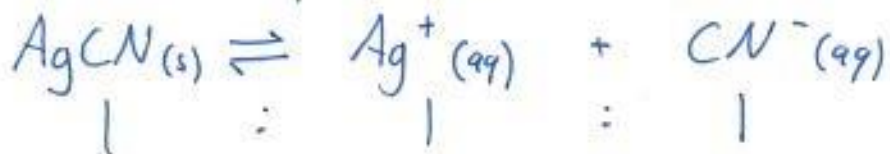
$$K_{sp} = [\text{Ag}^{+}]^2[\text{SO}_4^{2-}]$$

d) Silver carbonate — Ag₂CO₃



$$K_{sp} = [\text{Ag}^{+}]^2[\text{CO}_3^{2-}]$$

- 18) The maximum solubility of silver cyanide (AgCN) is 1.5×10^{-8} mol/L at 25°C . Calculate the value of the solubility product constant of silver cyanide.



$$[\text{Ag}^+] = [\text{CN}^-] = [\text{AgCN}] = 1.5 \times 10^{-8} \text{ mol/L}$$

$$K_{sp} = [\text{Ag}^+][\text{CN}^-]$$

$$K_{sp} = (1.5 \times 10^{-8})(1.5 \times 10^{-8})$$

$$\underline{\underline{K_{sp} = 2.25 \times 10^{-16}}}$$

- 19) A saturated solution of calcium fluoride (CaF_2) contains 1.2×10^{20} molecules of calcium fluoride per liter of solution. Calculate the value of the solubility product constant of the calcium fluoride.

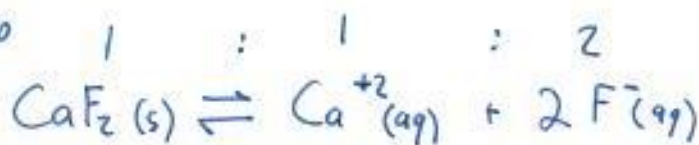
① Convert 1.2×10^{20} molecules/L in moles/L

$$[\text{CaF}_2] = \frac{6.02 \times 10^{23} \text{ molecules}}{1.2 \times 10^{20} \text{ molecules}} = \frac{1 \text{ mole}}{x} \quad \left. \vphantom{\frac{6.02 \times 10^{23} \text{ molecules}}{1.2 \times 10^{20} \text{ molecules}}} \right\} \text{ in 1L}$$

$$x = 2 \times 10^{-4} \text{ moles} \quad \left. \vphantom{x} \right\} \text{ in 1L}$$

$$[\text{CaF}_2] = 2 \times 10^{-4} \text{ mol/L} \quad \text{at Equilibrium} \\ \text{(when dissolved in (aq))}$$

② find K_{sp}



$$[\text{Ca}^{+2}] = 2 \times 10^{-4} \text{ mol/L}$$

$$[\text{F}^-] = (2) \cdot 2 \times 10^{-4} = \underline{\underline{4 \times 10^{-4} \text{ mol/L}}}$$

$$K_{sp} = [\text{Ca}^{+2}][\text{F}^-]^2$$

$$K_{sp} = (2 \times 10^{-4})(4 \times 10^{-4})^2 = \underline{\underline{3.2 \times 10^{-11}}}$$