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 ?

① $\% = \frac{[H_3O^+]}{[HA]} \times 100$

$8.7 = \frac{[H_3O^+]}{0.17} \times 100$

$[H_3O^+] = 0.0148 \text{ mol/L}$

② $pH = -\log[H_3O^+]$

$pH = -\log(0.0148)$

$pH = 1.83$

2) Find the concentration of hydronium ions $[H_3O^+]$ 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)

① $Ca(OH)_2 \rightarrow Ca^{+2} + 2OH^-$

(ratio) 1 : 1 : 2

if, $[Ca(OH)_2] = 0.012 \text{ mol/L}$

then, $[Ca^{+2}] = 0.012 \text{ mol/L}$

and, $[OH^-] = 2 \times 0.012 \text{ mol/L} = 0.024 \text{ mol/L}$

② $K_{water} = [H_3O^+] \cdot [OH^-]$

$10^{-14} = [H_3O^+] \cdot (0.024)$

$\frac{10^{-14}}{0.024} = [H_3O^+]$

$[H_3O^+] = 4.2 \times 10^{-13} \text{ mol/L}$

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?

a) $\% = \frac{[H_3O^+]}{[HOCN]} \times 100$

$17.04 = \frac{[H_3O^+]}{0.01} \times 100$

$[H_3O^+] = 0.001704 \text{ mol/L}$

b) $pH = -\log[H_3O^+]$

$pH - \log(0.001704)$

$pH = 2.77$

b)

	$HOCN_{(aq)} + \cancel{H_2O_{(l)}} \rightleftharpoons$	$H_3O^+_{(aq)}$ +	$OCN^-_{(aq)}$
T	1	1	1
I	0.01	0	0
C	−0.001704	+0.001704	+0.001704
E	0.008296	0.001704	0.001704

$K_a = \frac{[H_3O^+][OCN^-]}{[HOCN]} = \frac{(0.001704)(0.001704)}{0.008296} = K_a = 0.00035$

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.

① $H_2SO_4 \rightarrow 2H^+ + SO_4^{-2}$

$1 : 2 : 1$

② $[H^+] = 2 \cdot (0.25\ mol/L)$

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

③ $pH = -\log[H^+]$

$pH = -\log(0.5)$

$pH = 0.3$

④ $pH + pOH = 14$

$0.3 + pOH = 14$

$pOH = 13.7$

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?

① $HClO_{2\,(aq)} + \cancel{H_2O_{(l)}} \rightleftharpoons H_3O^+_{(aq)} + ClO_2^-{}_{(aq)}$

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

② $$K_a = \frac{[H_3O^+][ClO_2^-]}{[HClO_2]}$$

$$1.1\times10^{-2} = \frac{x \cdot x}{(0.2-x)}$$

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

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

↓

$[H_3O^+] = 0.0417\ mol/L$

③ $pH = -\log[H_3O^+]$

$pH = -\log(0.0417)$

$pH = 1.38$

④ $$ion\,\% = \frac{[H_3O^+]}{[HClO_2]} \times 100$$

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

$= 20.8\%$

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

$$CH_3NH_{2\,(aq)} + H_2O_{(l)} \rightleftharpoons CH_3NH_3{}^+{}_{(aq)} + OH^-{}_{(aq)}$$

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}?

① $pH + pOH = 14$

$9.16 + pOH = 14$

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

② $[OH^-] = 10^{-pOH}$

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

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

③ $CH_3NH_{2\,(aq)} + \cancel{H_2O}_{(l)} \rightleftharpoons CH_3NH_3{}^+{}_{(aq)} + OH^-{}_{(aq)}$

T	|	|	|
I	y	0	0
C	−0.0000145	+0.0000145	+0.0000145
E	y − 0.0000145	0.0000145	0.0000145

④ $K_B = \dfrac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]}$

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

solve for y

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

$\hookrightarrow y = \underset{\text{(initial)}}{[CH_3NH_2]} = 1.49 \times 10^{-5} \text{ mol/L}$

⑤ $\text{ion \%} = \dfrac{[OH^-]}{[CH_3NH_2]} \times 100$

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

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

7) Lactic acid ($CH_3CH(OH)COOH$) is found in dairy and milk products. If a lactic acid solution at 0.2mol/L dissociates as follows:

$$CH_3CH(OH)COOH_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + CH_3CH(OH)COO^-_{(aq)}$$

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 $[H_3O^+] = 10^{-pH}$

$[H_3O^+] = 10^{-2.285}$

$[H_3O^+] = 0.0052$ mol/L

② $CH_3CH(OH)COOH_{(aa)} + \cancel{H_2O_{(l)}} \rightleftharpoons H_3O^+_{(aq)} + CH_3CH(OH)COO^-_{(aq)}$

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

③ $$K_a = \frac{[H_3O^+][CH_3CH(OH)COO^-]}{[CH_3C(OH)COOH]}$$

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

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

b) $$\%\ ion = \frac{[H_3O^+]}{[CH_3CH(OH)COOH]} \times 100$$

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

$= 2.6\%$

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:

$$C_{18}H_{21}O_3N_{(aq)} + H_2O_{(l)} \rightleftharpoons C_{18}H_{21}O_3NH^+_{(aq)} + OH^-_{(aq)}$$

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)

① $pH + pOH = 14$

$11.2 + pOH = 14$

$pOH = 2.8$

② $[OH^-] = 10^{-pOH}$

$[OH^-] = 10^{-2.8}$

$[OH^-] = 0.00158$ mol/L

③ $C_{18}H_{21}O_3N_{(aq)} + \cancel{H_2O_{(l)}} \rightleftharpoons C_{18}H_{21}O_3NH^+_{(aq)} + OH^-_{(aq)}$

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

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

$K_B = 1.88 \times 10^{-6}$

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 0.00052mol/L.

a) Write out the dissociation equation of calcium oxalate.

b) What is the value of solubility constant (Ksp).

a) $CaC_2O_{4\,(s)} \rightleftharpoons Ca^{+2}_{(aq)} + C_2O_4^{-2}{}_{(aq)}$

1 : 1 : 1

b)

$[Ca^{+2}] = [C_2O_4^{-2}] = CaC_2O_4$ (molar solubility)

$K_{sp} = [Ca^{+2}]^1 \cdot [C_2O_4^{-2}]^1$

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

10) Butanic 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}.

	$C_4H_8O_2$ ⇌	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:

$$Ka = \frac{[H_3O+][C_4H_7O_2]}{[C_4H_8O_2]}$$

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

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

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

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

x = $\underline{\mathbf{0.000381}}$ **or** $\mathbf{-0.000396}$

✓ (valid) ✗ (invalid)

Substitute valid value of x into equilibrium concentrations:

$[C_4H_8O_2]$ = x = 0.000381 mol/L

$[H_3O+]$ = x = 0.000381 mol/L

$[C_4H_7O_2]$ = 0.01 – x = 0.00962 mol/L

Solve for pH:

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

pH = 3.42

11) Determine the mass of barium iodate ($Ba(IO_3)_2$) that must be dissolved in 300 mL of water to obtain a saturated solution of barium iodate with a solubility product constant of 4.01 × 10^{-9} at 25°C, if the volume of the solute is not taken into account.

① $Ba(IO_3)_{2\,(s)} \rightleftharpoons Ba^{+2}{}_{(aq)} + 2\,IO_3^-{}_{(aq)}$

$1 \quad : \quad 1 \quad : \quad 2$

$x \qquad \underbrace{x \quad : \quad 2x}_{\text{concentrations}}$

② $K_{sp} = [Ba^{+2}]^1 \cdot [IO_3^-]^2$

$4.01\times10^{-9} = (x)(2x)^2$

$4.01\times10^{-9} = x\,(4x^2)$

$4.01\times10^{-9} = 4x^3$

$x = 1\times10^{-3}\ mol/L$

③ $Ba(IO_3)_2$ required in 1L is 1×10^{-3} mol

$n = \frac{m}{M}$

$1\times10^{-3} = \frac{m}{487.13}$

$m = 0.4876g$

④ $0.4876g = 1L$

$y = 0.3L$

$y = 0.1463g$ in 300mL

12) Ammonia (a weak base) will dissociate as follows:

Ratio: 1 : 1 : 1

$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$

If at a certain temperature the Kb value of the above is 0.00002, what will the concentration of ammonia at equilibrium if the pH = 10.75

Find $[NH_3^+]$ @ Equilibrium

① $pH + pOH = 14$

$10.75 + pOH = 14$

$pOH = 3.25$

② $[OH^-] = 10^{-pOH}$

$[OH^-] = 10^{-3.25}$

$[OH^-] = 5.62\times10^{-4}\ mol/L$ also equal to $[NH_4^+]$ (equilibrium)

③ $K_B = \frac{[NH_4^+][OH^-]}{[NH_3]}$

$0.00002 = \frac{(5.62\times10^{-4})(5.62\times10^{-4})}{[NH_3]}$

$[NH_3] = 0.0158\ mol/L$

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

$$\underline{H_3PO_{4\,(aq)}} + \cancel{H_2O_{(l)}} \rightleftharpoons \underline{H_3O^+_{(aq)}} + \underline{H_2PO_4^-{}_{(aq)}}$$

If a phosphoric acid solution has a concentration of 0.1 mol/L with an acidity constant (Ka) 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

$$Ka = \frac{[H_3O^+][H_2PO_4^-]}{[H_3PO_4]} \qquad 7.5 \times 10^{-3} = \frac{x \cdot x}{(0.1-x)}$$

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

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

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

solve for x values

$$x = 0.02389 \quad \& \quad x = -0.\cancel{03}139$$

③

$$pH = -\log[H_3O^+]$$

$$pH = -\log[x]$$

$$pH = -\log(0.02389)$$

$$pH = 1.62$$

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 $6.3x10^{-4}$.

$$HF\,(aq) + H_2O\,(l) \rightarrow H_3O^+\,(aq) + F^-\,(aq)$$

	HF	**H_3O^+**	**F^-**
Theory	1	1	1
Initial	0.01	0	0
Change	-x	+x	+x
Equilibrium	0.01 - x	x	x

Substitute:

$$Ka = \frac{[H_3O+][F-]}{[HF]}$$

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

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

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

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

$$\mathbf{x = \underline{0.00221} \quad or \quad -0.00284}$$

✓ (valid) ✗ (invalid)

Substitute valid value of x into equilibrium concentrations:

$[C_4H_8O_2]$ = x = 0.00221 mol/L

$[H_3O+]$ = x = 0.00221 mol/L

$[C_4H_7O_2]$ = 0.01 – x = 0.00779 mol/L

Solve for pH:

$$pH = -\log(0.00221)$$

$$\underline{\underline{\mathbf{pH = 2.66}}}$$

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:

$$HA\,(aq) + \cancel{H_2O}(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$$

① $$\text{Ionization \%} = \frac{[H_3O^+]}{[HA]} \times 100$$

$$5\,\% = \frac{[H_3O^+]}{0.1} \times 100$$

$$\underline{[H_3O^+] = 0.005 \text{ mol/L}}$$

②

T	/	/	/
I	0.1	0	0
C	−0.005	+0.005	+0.005
E	0.095	0.005	0.005

③ $$K_a = \frac{[H_3O^+][A^-]}{[HA]} = \frac{(0.005)(0.005)}{0.095}$$

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

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^-)

$$CN^-_{(aq)} + H_2O_{(l)} \rightleftharpoons HCN_{(aq)} + OH^-_{(aq)}$$

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

b) Aqueous sulphate ($SO4^{-2}$)

$$SO_4^{-2}{}_{(aq)} + H_2O_{(l)} \rightleftharpoons HSO_4^-{}_{(aq)} + OH^-_{(aq)}$$

$$K_B = \frac{[HSO_4^-][OH^-]}{[SO_4^{-2}]}$$

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 (Ksp).

a) Copper(II) chloride — $CuCl_2$

$$CuCl_2(s) \rightleftharpoons Cu^{2+}(aq) + 2Cl^-(aq)$$

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

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b) Barium chloride — $BaCl_2$

$$BaCl_2(s) \rightleftharpoons Ba^{2+}(aq) + 2Cl^-(aq)$$

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

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c) Silver sulfate — Ag_2SO_4

$$Ag_2SO_4(s) \rightleftharpoons 2Ag^+(aq) + SO_4^{2-}(aq)$$

$$K_{sp} = [Ag^+]^2[SO_4^{2-}]$$

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d) Silver carbonate — Ag_2CO_3

$$Ag_2CO_3(s) \rightleftharpoons 2Ag^+(aq) + CO_3^{2-}(aq)$$

$$K_{sp} = [Ag^+]^2[CO_3^{2-}]$$

18) The maximum solubility of silver cyanide (AgCN) is $1.5x10^{-8}$ mol/L at 25°C. Calculate the value of the solubility product constant of silver cyanide.

$$AgCN_{(s)} \rightleftharpoons Ag^{+}_{(aq)} + CN^{-}_{(aq)}$$

1 : 1 : 1

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

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

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

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

19) A saturated solution of calcium fluoride (CaF_2) contains $1.2x10^{20}$ 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

$[CaF_2] =$ 6.02×10^{23} molecules = 1mole
1.2×10^{20} molecules = x } in 1L

$x = 2 \times 10^{-4}$ moles } in 1L

$[CaF_2] = 2 \times 10^{-4}$ mol/L at Equilibrium (when dissolved in (aq))

② find Ksp

1 : 1 : 2

$$CaF_{2}\,(s) \rightleftharpoons Ca^{+2}_{(aq)} + 2F^{-}_{(aq)}$$

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

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

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

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