Representative Exam Questions
On The Topic of Equilibrium
(Includes Acid / Base Equilibria)

Detailed Answer Key – Answers <u>and</u> Explanations

1. If a chemical equilibrium very much favors the products over the reactants, what would we expect its equilibrium constant to be like?

    * **a) We would expect the equilibrium constant to be a very large number.**

    b) We would expect the equilibrium constant to be a very small number.

    c) We would not be able to make any prediction about the size of the equilibrium constant, because its numerical value conveys no information on what is favored in the equilibrium.

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    The larger the equilibrium constant, the more the
    right (product) side is favored, and the smaller the
    equilibrium constant, the more the left (reactant)
    side is favored.
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2. If a chemical reaction reaches equilibrium very quickly, what would we expect its equilibrium constant to be like?

    a) We would expect the equilibrium constant to be a very large number.

    b) We would expect the equilibrium constant to be a very small number.

    * **c) We would not be able to make any prediction about the size of the equilibrium constant, because its numerical value conveys no information on the time required to reach equilibrium.**

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    The equilibrium constant tells us about the position
    of equilibrium, but not about the time required to get
    there.
    ```

3. What is always true about a chemical reaction when it is at equilibrium?

    a) The concentrations of all the reactants are zero.

    b) All substances shown in the reaction have equal concentrations.

    * **c) The forward reaction rate is equal to the reverse reaction rate.**

    d) Both the forward and reverse reaction have come to a stop.

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Equilibrium is the name given to the condition of
having two opposing processes occurring at the same
rate.  Although all concentrations remain constant at
equilibrium, they are not necessarily equal to each
other.
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4. Chemical equilibrium is said to be

    a) static   *b) **dynamic**

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The forward and reverse reactions continue to occur
when the system is at equilibrium.
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5. A negative equilibrium constant is obtained

    a) when a reaction is elementary

    b) when a reaction is non-elementary

    c) when the reaction is exothermic in the forward direction

    d) when you reverse the direction of a reaction that initially had a positive equilibrium constant

    * **e) under no circumstances**

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Equilibrium constants are calculated by multiplying
and dividing concentrations (for Kc) or pressures (for
Kp) raised to various powers.  Concentrations and
pressures are always positive numbers, and there is no
way to obtain a negative answer when positive numbers
are multiplied, divided, or raised to powers.
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6. For any chemical reaction that reaches an equilibrium, it is possible to write a balanced equilibrium equation and a corresponding equilibrium constant expression. On what factors will the numerical value of the equilibrium constant depend?

    a) None. A given reaction has a characteristic numerical value for its equilibrium constant that applies under all conditions.

    b) Temperature

    c) The choice of coefficients used to balance the equation describing the reaction.

    d) Whether the reaction is elementary or non-elementary

    *   **e) Both b and c**

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Except for reactions that are neither endothermic nor
exothermic (a very rare situation) changing the
temperature shifts the position of equilibrium.
Heating the system drives the reaction in the
endothermic direction, and cooling the system drives
the reaction in the exothermic direction.  If a
chemical equation is multiplied by a constant, the
equilibrium constant is raised to that power.  For
example, multiplying a reaction by 2 results in the
equilibrium constant being squared.  Also, reversing
the direction of a chemical reaction takes the
reciprocal of the equilibrium constant.  For example,
a reaction that has an equilibrium constant of 2 when
written in one direction has an equilibrium constant
of 0.5 when written in the reverse direction.
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7. Given the equilibrium reaction
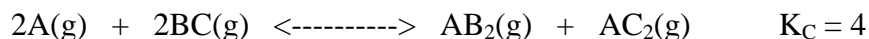
    $2A(g) + 2BC(g) \longleftrightarrow AB_2(g) + AC_2(g) \qquad K_C = 4$

    what is the numerical value of the equilibiurm constant (K_C) for the reaction

    $AB_2(g) + AC_2(g) \longleftrightarrow 2A(g) + 2BC(g) \qquad K_C = ?$
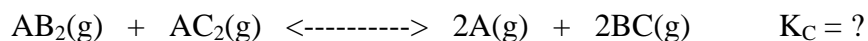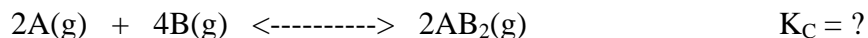
    a) -4   b) -2   *c) 0.25   d) 0.50   e) 4

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The original reaction has been reversed.  Therefore,
we must take the reciprocal of the original
equilibrium constant.  1 / 4  =  0.25
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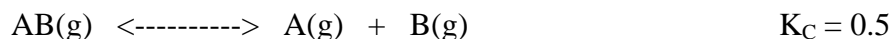
8. Given the equilibrium reaction

$$A(g) + 2B(g) \longleftrightarrow AB_2(g) \qquad\qquad K_C = 9$$

what is the numerical value of the equilibrium constant (K_C) for the reaction

$$2A(g) + 4B(g) \longleftrightarrow 2AB_2(g) \qquad\qquad K_C = ?$$

     a) 3   b) 4.5   c) 9   d) 18   ***e) 81**

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The original reaction was multiplied by 2, so the
equilibrium constant must be raised to the power 2,
that is, squared.  9²  =  81.
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9. Given the equilibrium reactions

$$AB(g) \longleftrightarrow A(g) + B(g) \qquad\qquad K_C = 0.5$$

$$2A(g) + B_2(g) \longleftrightarrow A_2B(g) + B(g) \qquad\qquad K_C = 6$$

what is the equilibrium constant (K_C) for the following reaction?

$$3A(g) + B_2(g) \longleftrightarrow A_2B(g) + AB(g) \qquad\qquad K_C = ?$$

     a) 3   b) 5.5   c) 6.5   ***d) 12**   e) 36

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You can get the target reaction by reversing the first
reaction and then adding the second reaction.
Reversing the first reaction changes its equilibrium
constant from 0.5 to 2.  Recall that reversing a
reaction takes the reciprocal of its equilibrium
constant.  1 / 0.5  =  2.  When chemical reactions are
added, their equilibrium constants are multiplied.
2  x  6  =  12.
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10. The reaction

$$2NH_3(g) \; \text{<----------> } \; N_2(g) \; + \; 3H_2(g)$$
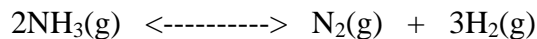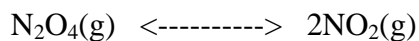
has $K_C = 6.46 \times 10^{-3}$ at 300 °C. What is the value of K_P for the reaction at this same temperature? Note that $R = 0.0820584$ (L atm) / (K mol). Temperatures must be in Kelvin units.

a) 2.62×10^{-4} b) 0.159   c) 0.304

d) 3.91   *e) 14.3

$K_P \;\; = \;\; K_C \; \cdot \; (RT)^{\Delta n} \quad\quad \Delta n \;\; = \;\; (1+3) \; - \; 2 \; = \; 4 \; - \; 2 \; = \; 2$

$K_P \;\; = \;\; 6.46 \times 10^{-3} \; \cdot \; (0.0820584 \times 573.15)^2 \;\; = \;\; 14.3$

11. The reaction

$$N_2O_4(g) \; \text{<----------> } \; 2NO_2(g)$$

has $K_P = 0.145$ at 25 °C. What is K_C for the reaction at this same temperature? Note that $R = 0.0820584$ (L atm) / (K mol). Temperatures must be in Kelvin units.

*a) 5.93×10^{-3} b) 7.07×10^{-2} c) 0.297

d) 3.55   e) 24.6

$K_C \;\; = \;\; K_P \; \cdot \; (RT)^{-\Delta n} \quad\quad\quad \Delta n \; = \; 2 \; - \; 1 \; = \; 1$
$\quad\quad\quad\quad\quad\quad\quad\quad\quad\quad\quad\quad\quad\quad -\Delta n \; = \; -1$

$K_C \;\; = \;\; 0.145 \; \cdot \; (0.0820584 \times 298.15)^{-1} \;\; = \;\; 5.93 \times 10^{-3}$

12. Sulfur trioxide gas, SO_3, was put in a heated reaction vessel and maintained at a temperature of 700 °C. Some of the SO_3 decomposed into SO_2 and O_2 gases, and an equilibrium was established between the SO_3, SO_2, and O_2. The chemical equation describing this equilibrium is

$$2SO_3(g) \; \longleftrightarrow \; 2SO_2(g) \; + \; O_2(g)$$
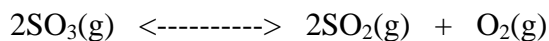
The equilibrium concentrations of these three gases were determined and found to be

$[SO_3]_{eq} \; = \; 1.14 \times 10^{-2} \, M$

$[SO_2]_{eq} \; = \; 8.64 \times 10^{-3} \, M$

$[O_2]_{eq} \; = \; 4.32 \times 10^{-3} \, M$

What is the numerical value of the equilibrium constant for this reaction at 700 °C?

<div>

a) 1.56×10^{-4} *b) 2.48 x 10^{-3} c) 5.81×10^{-2}

d) 4.72×10^{2} e) 6.38×10^{3}

</div>

$$K_C \; = \; \frac{[SO_2]^2_{eq} \, [O_2]_{eq}}{[SO_3]^2_{eq}}$$

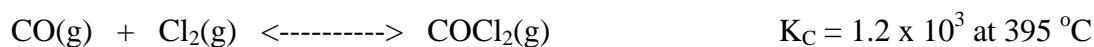$$K_C \; = \; \frac{(8.64 \times 10^{-3})^2 \, (4.32 \times 10^{-3})}{(1.14 \times 10^{-2})^2} \; = \; 2.48 \times 10^{-3}$$

13. At 395 °C, an equilibrium mixture of CO, Cl_2, and $COCl_2$ was found to have the following concentrations:

$$[CO]_{eq} = 6.5 \times 10^{-3} \text{ M}$$

$$[Cl_2]_{eq} = 1.4 \times 10^{-3} \text{ M}$$

$$[COCl_2]_{eq} = \text{unspecified}$$

The equilibrium reaction for this chemical system is as follows:

$CO(g) + Cl_2(g) \longleftrightarrow COCl_2(g)$ $K_C = 1.2 \times 10^3$ at 395 °C

What is the concentration of $COCl_2$ in this mixture at 395 °C?

     a) 7.4×10^{-9} M   b) 8.8×10^{-6} M   c) 3.5×10^{-3} M

     d) 6.9×10^{-3} M   ***e) 1.1×10^{-2} M**

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                         [COCl2]eq
        Kc    =         -------------
                       [CO]eq·[Cl2]eq

Solving this for [COCl2]eq gives

        [COCl2]eq    =    Kc · [CO]eq · [Cl2]eq

                     =    1.2 x 10³ (6.5 x 10⁻³) (1.4 x 10⁻³)

                     =    1.1 x 10⁻²
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14. Into a 10.00 L reaction vessel, a chemist placed 0.381 moles of HI, 0.492 moles of H_2, and 0.773 moles of I_2 at 425 °C. The chemical equation for the equilibrium of HI, H_2, and I_2 is as follows:

$2HI(g)$ <----------> $H_2(g)$ + $I_2(g)$ $\qquad$ $K_C = 1.84$ at 425 °C
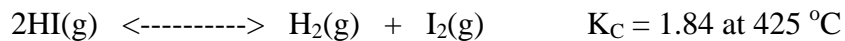
What can be said about this chemical system?

a) This is an equilibrium mixture. No net reaction will occur.

*b) **This is not an equilibrium mixture. Some of the H_2 and I_2 will combine to form HI until equilibrium is reached.**

c) This is not an equilibrium mixture. Some of the HI will decompose into H_2 and I_2 until equilibrium is reached.

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[HI]  = 0.381 mol / 10.00 L  =  0.0381 mol L⁻¹

[H₂] = 0.492 mol / 10.00 L  =  0.0492 mol L⁻¹

[I₂] = 0.773 mol / 10.00 L  =  0.0773 mol L⁻¹

            [H₂] [I₂]         (0.0492) (0.0773)
    Qc  =   ---------    =    -----------------    =   2.62
              [HI]²               (0.0381)²

Qc > Kc so the reaction will proceed to the left.   This
means some of the H₂ and I₂ will react to form HI.
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15. A 5.00 L reaction vessel at 3900 °C initially contained 0.200 moles each of N_2 and O_2 gases. These reacted to form NO as described by the following equation:

$N_2(g) + O_2(g) \leftarrow\text{-----}\rightarrow 2NO(g)$ $K_C = 0.0123$ at 3900 °C
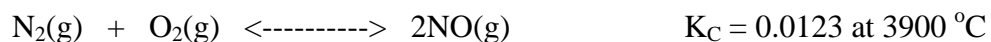
How many moles of NO will be present in the container when the system has reached equilibrium?

   a) 8.91×10^{-3} mol   ***b) 2.10 x 10^{-2} mol**   c) 4.23×10^{-2} mol

   d) 7.66×10^{-2} mol   e) 1.85×10^{-1} mol

$[N_2]_{initial}$ = 0.200 mol / 5.00 L = 0.0400 mol L^{-1}

$[O_2]_{initial}$ = 0.200 mol / 5.00 L = 0.0400 mol L^{-1}

$[NO]_{initial}$ = 0 mol L^{-1}

AN "ICE" TABLE

	$[N_2]$	$[O_2]$	$[NO]$
Initial	0.0400	0.0400	0
Change	-X	-X	+2X
Equilibrium	0.0400 - X	0.0400 - X	2X

$$\frac{[NO]^2_{eq}}{[N_2]_{eq} \; [O_2]_{eq}} = K_C$$

$$\frac{(2X)^2}{(0.0400 - X) \; (0.0400 - X)} = 0.0123$$

$$\frac{(2X)^2}{(0.0400 - X)^2} = 0.0123$$

$$\frac{2X}{(0.0400 - X)} = \pm0.1109 \qquad \text{Note: A square root has}$$
                                      both a positive and
                                      a negative root

$2X = \pm0.1109 (0.0400 - X)$

Case 1: positive root   Case 2: negative root

$2X = 0.1109 (0.0400 - X)$ $2X = -0.1109 (0.0400 - X)$

$2X = 0.004436 - 0.1109X$ $2X = -0.004436 + 0.1109X$

$2X + 0.1109X = 0.004436$ $2X - 0.1109X = -0.004436$

$(2 + 0.1109)X = 0.004436$ $(2 - 0.1109)X = -0.004436$

$2.1109X = 0.004436$ $1.8891X = -0.004436$

$X = 0.00210$ $X = -0.00235$

Since none of the equilibrium concentrations in the "ICE" table can be negative, the acceptable values for X are in the following range:

$$0 < X < 0.0400$$

Therefore, the X value obtained in Case 2 (-0.00235) must be rejected. The value of X obtained in Case 1 (0.00210) falls within the acceptable range. The equilibrium concentrations are now known:

$[N_2]_{eq} = 0.0400 - X = 0.0400 - 0.00210 = 0.0379$

$[O_2]_{eq} = 0.0400 - X = 0.0400 - 0.00210 = 0.0379$

$[NO]_{eq} = 2X = 2 \cdot (0.00210) = 0.00420$

To confirm that these are the equilibrium concentrations, substitute them back into the equilibrium constant expression and confirm that the original K_C is obtained.

$$K_C = \frac{(0.00420)^2}{(0.0379)(0.0379)} = 0.0123$$

The equilibrium concentration of NO is 0.00420 mol L^{-1}. The total number of moles of NO in the 5.00 L container is

0.00420 mol L^{-1} x 5.00 L = 0.0210 mol

$$= 2.10 \times 10^{-2} \text{ mol}$$

16.   What is the pH of a 3.88 M NH3 solution? NH3 is a weak base that establishes
        the following equilibrium when dissolved in water:
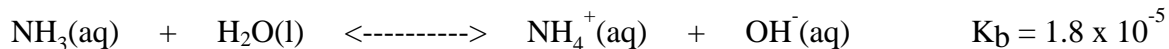
$$NH_3(aq) \quad + \quad H_2O(l) \quad <\text{---------}> \quad NH_4^+(aq) \quad + \quad OH^-(aq) \qquad K_b = 1.8 \times 10^{-5}$$

        The math is no worse than a quadratic, even if you don't approximate, but it
        should be ok to approximate in this problem.

             a) 2.1   b) 8.6   c) 9.2   *d) 11.9   e) 12.8

        $[NH_3]_{initial}$ = 3.88 mol L^{-1}

        $[NH_4^+]_{initial}$ = 0 mol L^{-1}

        $[OH^-]_{initial}$ ≈ 0 mol L^{-1} Note: In an aqueous system,
                                                a small amount of OH$^-$
                                                is always present from
                                                the self-ionization of
                                                H_2O.

                        AN "ICE" TABLE

	$[NH_3]$	$[NH_4^+]$	$[OH^-]$
Initial	3.88	0	≈ 0
Change	-X	+X	+X
Equilibrium	3.88 - X	X	X

$$\frac{[NH_4^+]_{eq} \ [OH^-]_{eq}}{[NH_3]_{eq}} = K_b$$

$$\frac{(X)\ (X)}{(3.88 - X)} = 1.8 \times 10^{-5}$$

        Since the initial concentration of NH$_3$ is fairly large,
        and the extent of ionization is small (as indicated by
        the small value of K$_b$), the concentration of NH$_3$ is
        probably not significantly reduced by the ionization.
        That is,

                    3.88 - X ≈ 3.88

        This makes it possible to ignore the subtraction of X
        in the denominator. This greatly simplifies the
        equation. It can now be written as follows:

$$\frac{(X)\ (X)}{3.88} = 1.8 \times 10^{-5}$$

$$\frac{X^2}{3.88} = 1.8 \times 10^{-5}$$

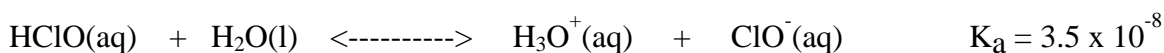$$X^2 = 3.88\ (1.8 \times 10^{-5}) = 7.0 \times 10^{-5}$$

$$X = 8.4 \times 10^{-3} = [OH^-]_{eq}$$

$$[H_3O^+]\ [OH^-] = K_W = 1.0 \times 10^{-14} \quad \text{Numerical value applies}$$
$$\text{at } 25\ ^{\circ}C$$

$$[H_3O^+] = \frac{K_W}{[OH^-]} = \frac{1.0 \times 10^{-14}}{8.4 \times 10^{-3}} = 1.2 \times 10^{-12}$$

$$pH = -\log [H_3O^+] = -\log (1.2 \times 10^{-12}) = 11.9$$

17. What is the $[H_3O^+]$ in a 2.50 M HClO solution? Hypochlorous acid, HClO, is a weak acid with the following equilibrium equation:

$$HClO(aq) \; + \; H_2O(l) \; \text{<---------->} \; H_3O^+(aq) \; + \; ClO^-(aq) \qquad K_a = 3.5 \times 10^{-8}$$

The math is no worse than a quadratic, even if you don't approximate, but it should be ok to approximate in this problem.

a) 3.4×10^{-11} b) 5.5×10^{-8} c) 5.1×10^{-6}

*d) 3.0×10^{-4} e) 2.3×10^{-2}

$[HClO]_{initial} \; = \; 2.50 \; \text{mol L}^{-1}$

$[H_3O^+]_{initial} \; \approx \; 0 \; \text{mol L}^{-1}$ Note: There is always a small amount of H_3O^+ present in an aqueous system due to the self-ionization of water.

$[ClO^-]_{initial} \; = \; 0 \; \text{mol L}^{-1}$

AN "ICE" TABLE

	[HClO]	$[H_3O^+]$	ClO^-]
Initial	2.50	≈ 0	0
Change	- X	+X	+X
Equilibrium	2.50 - X	X	X

$$\frac{[H_3O^+]_{eq} \; [ClO^-]_{eq}}{[HClO]_{eq}} \; = \; K_a$$

$$\frac{(X) \; (X)}{(2.50 - X)} \; = \; 3.5 \times 10^{-8}$$

Since the initial concentration of HClO is fairly large and the extent of ionization is very small (as indicated by the numerical value of K_a), the concentration of HClO is probably not significantly reduced by the ionization. That is,

$$2.50 - X \; \approx \; 2.50$$

With this approximation, we can ignore the subtraction of X in the denominator, which greatly simplifies the equation.
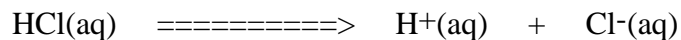
$$\frac{(X)(X)}{2.50} = 3.5 \times 10^{-8}$$

$$\frac{X^2}{2.50} = 3.5 \times 10^{-8}$$

$$X^2 = 2.50 (3.5 \times 10^{-8}) = 8.8 \times 10^{-8}$$

$$X = 3.0 \times 10^{-4} = [H_3O^+]$$

18.   What is the [OH⁻] in a 0.0100 M HCl solution? Note that HCl is a strong acid. It ionizes completely when dissolved in water:

   HCl(aq)   =========>   H⁺(aq)   +   Cl⁻(aq)

Enough HCl was added to provide 0.0100 moles of HCl for each liter of solution, but once added to water, it exists entirely as ions -- no molecular HCl is found in the solution.

   *a) 1.0 x 10⁻¹²   b) 1.0 x 10⁻¹⁰   c) 1.0 x 10⁻⁸

   d) 1.0 x 10⁻⁴   e) 1.0 x 10⁻²

Note that the equation is written to show H⁺ rather than H₃O⁺. You will find that H⁺ and H₃O⁺ are often used interchangeably. If you prefer, you can add H₂O to both sides of the equation and write it as

   HCl(aq)   +   H₂O(l)   =========>   H₃O⁺(aq)   +   Cl⁻(aq)
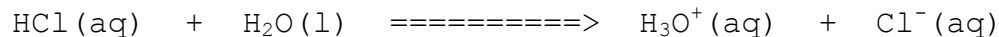
HCl is a strong acid. We regard the ionization of a strong acid to be 100% complete, rather than having the ions in equilibrium with the unionized acid. We can still set up an "ICE" table, but there is no unknown X in it. Rather, numerical values can be obtained for every cell in the table.

[HCl]$_{initial}$ =   0.0100 mol L⁻¹

[H₃O⁺]$_{initial}$ ≈   0 mol L⁻¹   Note: There is always a small amount of H₃O⁺ present in an aqueous system due to the self-ionization of water.

[Cl⁻]$_{initial}$ =   0 mol L⁻¹

                    AN "ICE" TABLE

	[HCl]	[H₃O⁺]	[Cl⁻]
Initial	0.0100	≈ 0	0
Change	-0.0100	+0.0100	+0.0100
Equilibrium	0	0.0100	0.0100

$[H_3O^+]_{eq}$ = 0.0100 mol L^{-1}

$[H_3O^+]_{eq} [OH^-]_{eq}$ = K_W = 1.0 X 10^{-14} Numerical value
                                                             applies at 25 °C

$[OH^-]_{eq}$ = $\dfrac{K_W}{[H_3O^+]_{eq}}$ = $\dfrac{1.0 \times 10^{-14}}{0.0100}$ = 1.0 x 10^{-12}

19.   What is the pOH of a 4.20 M hydrocyanic acid (HCN) solution? Note that
       hydrocyanic acid is a weak acid that establishes the following equilibrium in
       water:
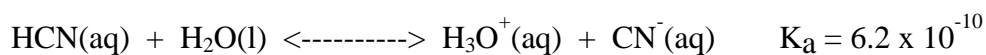
       $HCN(aq) + H_2O(l) \longleftrightarrow H_3O^+(aq) + CN^-(aq)$ $K_a = 6.2 \times 10^{-10}$

       The math is no worse than a quadratic, even if you don't approximate, but it
       should be ok to approximate in this problem.

              a) 1.6   b) 4.3   c) 8.1   *d) 9.7   e) 12.4

$[HCN]_{initial}$ = 4.20 mol L^{-1}

$[H_3O^+]_{initial}$ ≈ 0   Note: There is always a small
                                                 amount of H_3O^+ present
                                                 in an aqueous system
                                                 due to the self-
                                                 ionization of water.

$[CN^-]_{initial}$ = 0

                    AN "ICE" TABLE

	[HCN]	$[H_3O^+]$	$[CN^-]$
Initial	4.20	≈ 0	0
Change	-X	+X	+X
Equilibrium	4.20 - X	X	X

$\dfrac{[H_3O^+]_{eq} [CN^-]_{eq}}{[HCN]_{eq}}$ = K_a

$\dfrac{(X) (X)}{4.20 - X}$ = 6.2 x 10^{-10}

Since the initial concentration of HCN is fairly large and the extent of ionization if very small (as indicated by the small value of K_a), the concentration of HCN is probably not significantly reduced by the ionization. That is,

$$4.20 - X \approx 4.20$$

With this approximation, the subtraction of X in the denominator can be ignored. This greatly simplifies the equation.

$$\frac{(X)\,(X)}{4.20} = 6.2 \times 10^{-10}$$

$$\frac{X^2}{4.20} = 6.2 \times 10^{-10}$$

$$X^2 = 4.20\,(6.2 \times 10^{-10}) = 2.6 \times 10^{-9}$$

$$X = 5.1 \times 10^{-5} = [H_3O^+]_{eq}$$

$$[H_3O^+]_{eq}\,[OH^-]_{eq} = K_W = 1.0 \times 10^{-14} \qquad \text{Numerical value applies at 25 }^{\circ}\text{C}$$

$$[OH^-]_{eq} = \frac{K_W}{[H_3O^+]_{eq}} = \frac{1.0 \times 10^{-14}}{5.1 \times 10^{-5}} = 2.0 \times 10^{-10}$$

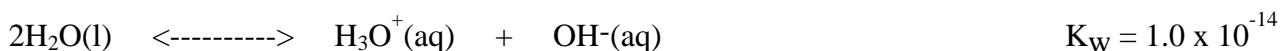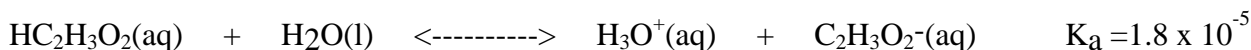$$pOH = -\log [OH^-]_{eq} = -\log (2.0 \times 10^{-10}) = 9.7$$

20. What is the $[H_3O^+]$ in a 3.00 M sodium acetate ($NaHC_2H_3O_2$) solution? Note that sodium acetate is a salt that undergoes hydrolysis after being dissolved in water. The salt is freely soluble, and dissolves without establishing an equilibrium:

$$NaHC_2H_3O_2(aq) \quad =========> \quad Na^+(aq) \quad + \quad C_2H_3O_2^-(aq)$$
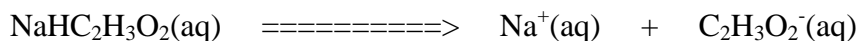
The Na^+, being the conjugate acid of a strong base ($NaOH$) is too weak an acid to react with water. The $C_2H_3O_2^-$, being the conjugate base of a weak acid ($HC_2H_3O_2$) has enough base strength to react with water:

$$C_2H_3O_2^-(aq) + H_2O(l) \quad <----------> \quad HC_2H_3O_2(aq) \quad + \quad OH^-(aq)$$
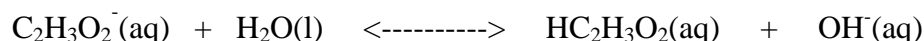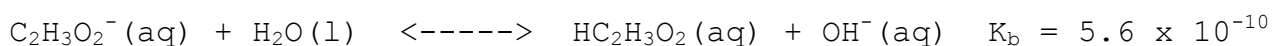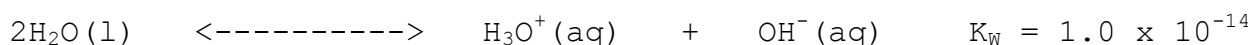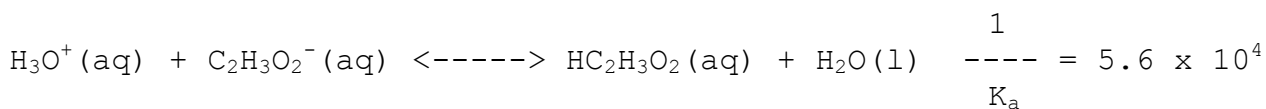
The numerical value of the equilibrium constant for the above reaction is normally not published in tables, because it can be obtained from the equilibrium constants of two other reactions that ARE published:

$$HC_2H_3O_2(aq) \quad + \quad H_2O(l) \quad <----------> \quad H_3O^+(aq) \quad + \quad C_2H_3O_2^-(aq) \qquad K_a = 1.8 \times 10^{-5}$$

$$2H_2O(l) \quad <----------> \quad H_3O^+(aq) \quad + \quad OH^-(aq) \qquad K_w = 1.0 \times 10^{-14}$$

By combining these two reactions in the appropriate way, you can produce the reaction for which the numerical value of the equilibrium constant is desired. Once you have that, you can do an ICE table calculation in the usual way. It should be possible to make an approximation in your solution to the ICE table calculation to avoid having to solve the quadratic formula.

a) 1.6×10^{-13} *b) 2.4×10^{-10} c) 2.5×10^{-7}

d) 4.1×10^{-5} e) 1.8×10^{-3}

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If we reverse the acetic acid ionization reaction and
add the water ionization reaction, we obtain the
desired reaction:
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$$H_3O^+(aq) + C_2H_3O_2^-(aq) <-----> HC_2H_3O_2(aq) + H_2O(l) \quad \frac{1}{K_a} = 5.6 \times 10^4$$

$$2H_2O(l) \quad <----------> \quad H_3O^+(aq) \quad + \quad OH^-(aq) \qquad K_w = 1.0 \times 10^{-14}$$

$$C_2H_3O_2^-(aq) + H_2O(l) \quad <-----> \quad HC_2H_3O_2(aq) + OH^-(aq) \quad K_b = 5.6 \times 10^{-10}$$

Recall that when we reverse the direction of a chemical reaction, we take the reciprocal of its equilibrium constant, and when we add chemical reactions, we multiply their equilibrium constants. Using these principles, we have determined that the desired reaction has an equilibrium constant of 5.6×10^{-10}. This equilibrium constant has been labeled K_b since the $C_2H_3O_2^-$ is acting as a base. We know that it acts as a base because it generates OH^- ions in its reaction with water. We refer to $C_2H_3O_2^-$ as the conjugate base of $HC_2H_3O_2$ because $C_2H_3O_2^-$ is what is left over after $HC_2H_3O_2$, acting as an acid, donates an H^+ to H_2O to form H_3O^+.

Notice that what has been shown above is that

$$\frac{1}{K_a} \cdot K_W = K_b$$

This can be rewritten in the form

$$K_a \cdot K_b = K_W$$

This is a general relationship between the equilibrium constants of any acid and its conjugate base. Since K_W is a known constant, we can calculate K_b if we know K_a and we can calculate K_a if we know K_b. Keeping this formula in mind, it will not be necessary to actually set up the chemical arithmetic each time, as I have done above. Once you know the pattern, you can quickly and easily find the needed equilibrium constant.

Now that we have the equilibrium constant for the desired reaction, we can calculate the equilibrium concentrations in the usual manner.

Since the $NaC_2H_3O_2$ is 100% ionized in water, a 3.00 M solution of $NaC_2H_3O_2$ is really a solution that is 3.00 M in both Na^+ and $C_2H_3O_2^-$. Therefore, the initial concentration of $C_2H_3O_2^-$ is 3.00 M.

$[C_2H_3O_2^-]_{initial}$ = 3.00 mol L^{-1}

$[HC_2H_3O_2^-]_{initial}$ = 0 mol L^{-1}

$[OH^-]_{initial}$ ≈ 0   Note: In an aqueous solution, there will
                          always be a small amount of OH^-
                          present from the self-ionization of
                          water.

AN "ICE" TABLE

	$[C_2H_3O_2^-]$	$[HC_2H_3O_2]$	$[OH^-]$
Initial	3.00	0	≈ 0
Change	-X	+X	+X
Equilibrium	3.00 - X	X	X

$$\frac{[HC_2H_3O_2]_{eq}\ [OH^-]_{eq}}{[C_2H_3O_2^-]_{eq}} = K_b$$

$$\frac{(X)\ (X)}{3.00 - X} = 5.6 \times 10^{-10}$$

Since the initial concentration of $C_2H_3O_2^-$ is fairly large and
the extent of reaction is small (as indicated by the small
value of K_b), the concentration of $C_2H_3O_2^-$ is probably not
significantly reduced by the reaction. That is,

$$3.00 - X ≈ 3.00$$

With this approximation, we can ignore the subtraction of X
in the denominator. This greatly simplifies the equation.

$$\frac{(X)\ (X)}{3.00} = 5.6 \times 10^{-10}$$

$$\frac{X^2}{3.00} = 5.6 \times 10^{-10}$$

$$X^2 = 3.00\ (5.6 \times 10^{-10}) = 1.7 \times 10^{-9}$$

$$X = 4.1 \times 10^{-5} = [OH^-]_{eq}$$

$$[H_3O^+]_{eq}\ [OH^-]_{eq} = K_W = 1.0 \times 10^{-14}$$ Numerical value applies at 25 °C

$$[H_3O^+]_{eq} = \frac{K_W}{[OH^-]_{eq}} = \frac{1.0 \times 10^{-14}}{4.1 \times 10^{-5}} = 2.4 \times 10^{-10}$$