Key Equations:
\[
[H^+][OH^-] = 1.00 \times 10^{-14} \quad \text{pH} = -\log[H^+] \quad [H^+] = 10^{-\text{pH}} \quad \text{pH} + \text{pOH} = 14
\]
for weak acids in water:
\[
K_a = \frac{[H^+][HA]_{\text{init}}}{[HA]_{\text{init}}} \quad [H^+] = \sqrt{K_a \times [HA]_{\text{init}}}
\]
for weak bases in water:
\[
K_b = \frac{[OH^-][\text{Base}]_{\text{init}}}{[\text{Base}]_{\text{init}}} \quad [OH^-] = \sqrt{K_b \times [\text{Base}]_{\text{init}}}
\]
(above weak acid/base equations assume <5% ionization and assume no alternative source of common ions)

\[K_aK_b = 10^{-14}\] for a conjugate acid/base pair

Quadratic Equation:
\[
x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}
\]

1. Which of the following is false about a system at equilibrium:
   a) The rate of the forward reaction becomes equal to the rate of the reverse reaction
   b) So long as the equilibrium is not disturbed, the relative amounts of products and reactants present will not change no matter how long you wait
   c) In an equilibrium situation, interconversion between reactants and products continues to occur.
   d) The rate constant for the forward reaction becomes equal to the rate constant for the reverse reaction

2. Which of the following statements are true, regarding the equilibrium constant \(K\) for a reaction and the reaction quotient \(Q\):
   1) If \(Q > K\), the reaction is not at equilibrium, and will reach equilibrium by shifting some products over to reactants
   2) If \(K = 3.2 \times 10^{-6}\), the reaction is product favored
   3) If \(K = 5.2 \times 10^4\), the reaction is product favored
   4) If \(Q = K\), the reaction is already at equilibrium.

   a) 1 and 2 only
   b) 1, 2, and 4 only
   c) 1, 3, and 4 only
   d) 3 and 4 only
   e) 2 and 4 only

3. Which of the following statements are true regarding equilibrium constants for the following reaction:
\[
2 \text{NH}_3 (g) \rightleftharpoons \text{N}_2 (g) + 3\text{H}_2 (g) \quad \Delta H^\circ = +92 \text{ kJ}
\]
   a) Increasing the volume of the container will increase the equilibrium constant
   b) Increasing the temperature of the reaction will increase the equilibrium constant
   c) Increasing the volume of the container will increase the concentration of \(\text{NH}_3\) (g)
   d) Increasing the concentration of \(\text{H}_2\) (g) will increase the equilibrium constant
   e) Increasing the concentration of \(\text{NH}_3\) (g) will increase the equilibrium constant
4. Identify the correct equilibrium expression for the following reaction.

\[ A (aq) + 2B (aq) \rightleftharpoons C (aq) + D (s) \]

a) \[ [A] [B]^2 / [C] [D] \]
b) \[ [C] [D] / [A] [B]^2 \]
c) \[ [A] [B]^2 / [C] \]
d) \[ [C] / [A] [B]^2 \]

5. For the following reaction, determine whether the system is at equilibrium when \([CO] = 0.50\) M and \([CO_2] = 0.75\) M. The system _________ at equilibrium, because _______________

\[ C (s) + CO_2 (g) \rightleftharpoons 2 CO (g) \quad K_c = 168 \]

a) Is; the value of Q is 0.33
b) Is not; the value of Q is 0.33
c) Is; the value of Q is 0.67
d) Is not; the value of Q is 0.67
e) More information is needed to answer this question

6. What is the equilibrium constant \(K_c\) for the following reaction, if at equilibrium \([C_4H_{10}] = 0.018\) M, \([C_2H_6] = 0.035\) M, and \([C_2H_4] = 0.035\) M?

\[ C_{4}H_{10} (g) \rightleftharpoons C_{2}H_{6} (g) + C_{2}H_{4} (g) \]

a) 0.068
b) 0.13
c) 14
d) 2.2 \times 10^{-5}

e) More information is needed to answer this question

7. What is the equilibrium concentration of \(N_2O\) (g) (in moles/liter), if at equilibrium \([N_2]= 0.048\) M and \([O_2] = 0.093\) M?

\[ 2N_2 (g) + O_2 (g) \rightleftharpoons 2 N_2O (g) \quad K_c = 1.5 \times 10^{-30} \]

a) 8.2 \times 10^{-17}
b) 1.8 \times 10^{-17}
c) 4.7 \times 10^{-27}
d) 3.4 \times 10^{-28}
e) 3.2 \times 10^{-34}
8. When 1.00 mol NH$_3$ (g) was placed into a 1 L container and allowed to reach equilibrium, the resulting mixture contained 0.60 mol NH$_3$ (g). How many moles of N$_2$ (g) and H$_2$ (g) are present at equilibrium?

\[ 2 \text{NH}_3 (g) \rightleftharpoons \text{N}_2 (g) + 3\text{H}_2 (g) \]

a) 0.40 moles of N$_2$ ; 1.20 moles of H$_2$

b) 0.80 moles of N$_2$ ; 2.40 moles of H$_2$ (g)

c) 0.20 moles of N$_2$ ; 0.60 moles of H$_2$ (g)

d) 0.80 moles of N$_2$ ; 0.27 moles of H$_2$ (g)

e) 0.20 moles of N$_2$ ; 0.40 moles of H$_2$ (g)

9. 0.50 mol of I$_2$ (g) and 0.50 mol of Br$_2$ (g) are placed in a 1.00 L flask and allowed to reach equilibrium. At equilibrium, the flask contains 0.84 mol of IBr. What is the value of K$_c$ for this reaction?

\[ \text{I}_2 (g) + \text{Br}_2 (g) \rightleftharpoons 2\text{IBr} (g) \]

a) 11

b) 4.0

c) 110

d) 6.1

10. When 0.70 mol NO$_2$ was placed in a 1.00 L flask and allowed to reach equilibrium, its concentration was found to be 0.28 M, once equilibrium was established. Calculate K$_c$ for this reaction.

\[ 2\text{NO}_2 (g) \rightleftharpoons 2\text{NO} (g) + \text{O}_2 (g) \]

a) 1.9

b) 0.94

c) 0.47

d) 0.14

11. Calculate the equilibrium concentration of CO (g) and Cl$_2$ (g) if the initial concentration of COCl$_2$ (g) was 0.0627 M.

\[ \text{COCl}_2 (g) \rightleftharpoons \text{CO} (g) + \text{Cl}_2 (g) \quad \text{K}_c = 2.73 \times 10^{-10} \]

a) $2.30 \times 10^8$ M

b) $1.52 \times 10^{-4}$ M

c) $2.03 \times 10^{-3}$ M

d) $4.14 \times 10^{-6}$ M

e) $1.71 \times 10^{-11}$ M
12. Consider the following reaction at equilibrium. Adding N\textsubscript{2} (g) to this reaction will:

\[ 2 \text{NH}_3 (g) \rightleftharpoons \text{N}_2 (g) + 3\text{H}_2 (g) \quad \Delta H^\circ = +92 \text{ kJ} \]

a) Decrease the concentration of NH\textsubscript{3} (g) at equilibrium  
b) Decrease the concentration of H\textsubscript{2} (g) at equilibrium  
c) Increase the value of the equilibrium constant  
d) Cause the reaction to shift to the right

13. Given the following equilibrium, which of the following statements is true?

\[ \text{C (s)} + \text{CO}_2 (g) \rightleftharpoons 2 \text{CO (g)} \quad \Delta H^\circ = +143 \text{ kJ} \]

a) An increase in temperature will cause a shift in the equilibrium position to the left  
b) An increase in the concentration of CO\textsubscript{2} (g) will cause the concentration of CO (g) to decrease  
c) An increase in the amount of carbon will cause the amount of CO (g) to increase  
d) An increase in temperature will make the equilibrium constant get larger  
e) A reduction in volume will cause a shift in the equilibrium position to the right

14. What would be the effect of reducing the volume for the following system, once equilibrium was reestablished:

\[ \text{N}_2 (g) + 3\text{H}_2 (g) \rightleftharpoons 2 \text{NH}_3 (g) \quad \Delta H^\circ = +92 \text{ kJ} \]

a) Decrease the number of moles of NH\textsubscript{3} (g) at equilibrium  
b) Decrease the number of moles of H\textsubscript{2} (g) at equilibrium  
c) Decrease the value of the equilibrium constant  
d) Cause the reaction to shift to the left

15. The [H\textsuperscript{+}] and pH of 0.021 M HNO\textsubscript{3} are:

a) 4.8 x 10\textsuperscript{-13} M and 12.32  
b) 0.021 M and 12.32  
c) 0.021 M and 1.68  
d) 0.021 M and –1.68  
e) 4.8 x 10\textsuperscript{-6} M and 5.32

16. Calculate the hydronium ion concentration in a 0.012 M aqueous solution of NaOH.

a) 7.8 x 10\textsuperscript{-4} M  
b) 5.5 x 10\textsuperscript{-13} M  
c) 5.6 x 10\textsuperscript{-11} M  
d) 8.3 x 10\textsuperscript{-13} M  
e) none of the above
17. What is the [OH\(^-\)] concentration of a solution with pH = 4.50?

a) \(3.2 \times 10^{-5}\) M  
b) \(8.2 \times 10^{-9}\) M  
c) \(8.3 \times 10^{-10}\) M  
d) \(3.2 \times 10^{-10}\) M  
e) none of the above

18. A 0.55 M solution of the weak acid HBrO has a pH of 4.48. What is the value of \(K_a\) for HBrO?

a) \(2.0 \times 10^{-9}\) M  
b) \(1.1 \times 10^{-9}\) M  
c) \(6.0 \times 10^{-5}\) M  
d) \(3.3 \times 10^{-5}\) M  
e) none of the above

19. Calculate the pH of 0.020 M hypochlorous acid, \(K_a = 3.0 \times 10^{-8}\).

a) 2.45  
b) -2.45  
c) 3.60  
d) 9.22  
e) 4.61

20. The basicity constant \(K_b\) for \(C_6H_5NH_2 = 4.3 \times 10^{-10}\). Calculate the pH of a 0.15 M solution of \(C_6H_5NH_3^+\) in water.

a) 11.3  
b) 8.6  
c) 5.2  
d) 2.7  
e) none of the above

21. Calculate the pH of a 0.20 M solution of \(C_4H_5NH_2\) in water. The basicity constant \(K_b\) for \(C_4H_5NH_2 = 3.5 \times 10^{-6}\).

a) 3.1  
b) 4.9  
c) 10.9  
d) 9.6  
e) none of the above
22. The $K_a$ for HF is $7.0 \times 10^{-4}$. What is the value of $K_b$ for NaF?

a) $2.0 \times 10^{-8}$  
b) $1.4 \times 10^{-11}$  
c) $7.0 \times 10^{-18}$  
d) $1.4 \times 10^{-10}$  
e)  

23. Calculate the pH of 0.374 M solution of NaNO$_2$ ($K_a$ for HNO$_2$ = $4.5 \times 10^{-4}$).

a) 8.5  
b) 1.9  
c) 0.013  
d) 12.1  
e) none of the above  

24. Which one of the following is the strongest acid?

a) CH$_3$COOH ($K_a = 1.8 \times 10^{-5}$)  
b) HCOOH ($K_a = 1.0 \times 10^{-4}$)  
c) HClO ($K_a = 3.0 \times 10^{-8}$)  
d) HF ($K_a = 6.8 \times 10^{-4}$)  

25. What is the conjugate acid of C$_4$H$_7$NH$_2$?

a) C$_4$H$_7$NH$^+$  
b) C$_4$H$_7$NH$^-$  
c) C$_4$H$_7$NH$_3^+$  
d) C$_4$H$_7$NH$_3^-$  

26. Which one of the following 0.1 M solutions would have a pH of 7.0?

a) Na$_2$S  
b) CoCl$_3$  
c) NaNO$_3$  
d) NH$_4$Cl  
e) None of these  

27. Given the $K_a$ values shown, which one of the anions shown is the strongest base?

- CH$_3$COOH ($K_a = 1.8 \times 10^{-5}$)  
- HCOOH ($K_a = 1.0 \times 10^{-4}$)  
- HClO ($K_a = 3.0 \times 10^{-8}$)  
- HF ($K_a = 6.8 \times 10^{-4}$)  

a) CH$_3$COO$^-$  
b) HCOO$^-$  
c) ClO$^-$  
d) F$^-$  

28. For the reaction shown, which of the following statements would be false?

\[ \text{H}_2\text{CO}_3 \text{(aq)} + \text{CH}_3\text{COO}^- \text{(aq)} \rightleftharpoons \text{CH}_3\text{COOH (aq)} + \text{HCO}_3^- \text{(aq)} \quad K = 2.3 \times 10^{-2} \]

a) CH$_3$COOH is the strongest acid
b) HCO$_3^-$ anion is the strongest base
c) H$_2$CO$_3$ is the strongest acid
d) The solution will contain more H$_2$CO$_3$ than CH$_3$COOH at equilibrium

29. Which of the following would give an acidic solution in water?

a) NaCN
b) KF
c) NH$_3$
d) CH$_3$COOH

30. Rank the relative basicity of NH$_3$, OH$^-$, F$^-$, HSO$_4^-$, given the following acidity data:

\[ \text{NH}_4^+ \text{ (K}_a \text{ = 1.8 x 10}^{-5}) \quad \text{HF (K}_a \text{ = 7.2 x 10}^{-4}) \]

a) OH$^-$ > NH$_3$ > HSO$_4^-$ > F$^-$
b) OH$^-$ > F$^-$ > NH$_3$ > HSO$_4^-$
c) HSO$_4^-$ > F$^-$ > NH$_3$ > OH$^-$
d) OH$^-$ > NH$_3$ > F$^-$ > HSO$_4^-$
e) None of the above

31. Which of the following would not give an acidic solution?

a) H$_2$S
b) NH$_4$Cl
c) NaNO$_2$
d) FeCl$_3$
e) None of these

32. Which of the following acidity relationships is true?

a) H$_2$SO$_3$ > H$_2$SO$_4$
b) H$_3$PO$_4^-$ > HPO$_4^{2-}$
c) HF > HClO$_4$
d) H$_2$CO$_3$ > HNO$_3$
e) None of these

33. For the reaction shown, which of the following statements would be false?

\[ \text{H}_2\text{SO}_3 \text{(aq)} + \text{HS}^- \text{(aq)} \rightleftharpoons \text{HSO}_3^- \text{(aq)} + \text{H}_2\text{S} \text{(aq)} \]

a) H$_2$SO$_3$ and H$_2$S are acids
b) HS$^-$ and HSO$_3^-$ are bases
c) The equilibrium will favor the side with the weaker acid and the weaker base
d) H$_2$SO$_3$ and HS$^-$ are a conjugate acid/base pair
1. D
2. C
3. B
4. D
5. B
6. A
7. B
8. C
9. C
10. C
11. D
12. B
13. D
14. B
15. C
16. D
17. D
18. A
19. E
20. D
21. C
22. B
23. A
24. D
25. D
26. C
27. C
28. C
29. D
30. D
31. C
32. B
33. D