Acid/Base homework booklet #2 Name________________________Period___

Complete all of the questions. Wherever possible, show the work in the question booklet.

1. The pH at which an indicator changes color is known as its
   A. stoichiometric endpoint B. standard point C. equivilence point [✓ D. transition point]

2. What is true about the transition point of all indicators described by the following equilibrium:

$$HIn + H_2O \leftrightarrows H_3O^+ + In^-$$

[✓ A. $[HIn] = [In^-]$]
B. $[H_3O^+] = [OH^-]$
C. pH = pOH
D. pH = K_a

colourless ← 10.0 pH= 0.00 yellow ← 7.3

3. What color would 1.0 M HCl be in an indicator mixture consisting of phenol red and thymophthalein?
[✓ A. yellow]
B. red
C. ~~blue~~
D. ~~colorless~~

pH= 13.00 7.3 → red

4. What color would 0.10 M NaOH be in an indicator mixture consisting of phenol red and bromcresol green? 4.6 → blue
[✓ A. purple]
B. green
C. blue
D. yellow

pH= 13.00 → red 11.05 → red

5. What color would 0.10 M NaOH be in an indicator mixture consisting of phenol red and alizarin yellow ?
[✓ A. red]
B. yellow
C. colorless
D. orange

6. When the indicator thymol blue is added to 0.10 M solution of an unknown acid, the solution is [✓ red]. The acid could be
A. H_2S
B. HF
[✓ C. HNO_3]
D. HCN

thymol blue
red ← 2.0 → yellow ← 8.8 → blue

∴ pH < 2.0

∴ strong acid

7. When the indicator alazarin yellow is added to 0.010 M solution of an unknown compound, the solution is red. ~~The~~ unknown compound could ~~could~~ be
   A. HNO_3
   B. KIO_3
   [✓ C.] NaOH
   D. HCN

0
∴ pH > 11.05 ∴ strong base

red ← 5.4

8. At pH = 4.0 methyl red will be
   A. ~~yellow~~ and [HInd] ~~<~~ $[Ind^-]$
   B. ~~yellow~~ and [HInd] > $[Ind^-]$
   C. red and [HInd] ~~<~~ $[Ind^-]$
   [✓ D.] red and [HInd] > $[Ind^-]$

9. Methyl red is orange in a 0.10 M solution of an acid. The acid could be
   A. ~~NaOH~~
   B. ~~NH_3~~
   C. C_6H_5OH
   D. HI

transition pt

pH = 5.4 ∴ not a strong acid

$C_6H_5OH + H_2O \rightleftharpoons C_6H_5O^- + H_3O^+$
.10 | 0 | 0
−x | +x | +x

pH = 5.4 ∴ $[H_3O^+] =$ 4×10^{-6}

$K_a =$

$K_a = \frac{(4 \times 10^{-6})^2}{(.10 - 4 \times 10^{-6})}$

$= 1.6 \times 10^{-11}$

10. Thymol blue is green in a 0.72 M solution of an unknown solution. The unknown solution could be
   A. ~~NaOH~~
   B. ~~HI~~
   [✓ C.] NaHCOO
   D. ~~HCN~~

∴ pH = 8.8 ∴ weak base

$HCOO^- + H_2O \rightleftharpoons HCOOH + OH^-$
.72 | 0 | 0
−x | +x | +x

pH = 8.8
$[OH^-] = 6.3 \times 10^{-6}$

$K_b = \frac{(6.3 \times 10^{-6})^2}{(.72 - 6.3 \times 10^{-6})}$

$= 5.5 \times 10^{-11}$

$K_a = 1.8 \times 10^{-4}$ ✓

11. Which would produce a yellow solution at pH = 4.0?
   A. methyl violet pH < 0.80
   B. methyl red pH > 5.4
   [✓ C.] chlorophenol red pH < 6.0
   D. indigo carmen pH > 12.2

12. Which would produce an orange solution at pH = 6.0?
   A. thymol blue
   B. phenol red
   [✓ C.] chlorophenol red
   D. methyl red

Do in class

∴ transition point

13. Which would produce a green solution at pH = 6.8?
   [✓ A.] bromthymol blue
   B. bromcresol green
   C. thymol blue
   D. indigo carmine

∴ transition point

14. The chemical indicator bromthymol blue changes from yellow to blue as a result of the addition of

A. 1.0 M NH_4Cl
[✓ B.] 1.0 M K_2CO_3
C. 1.0 M HNO_2
D. 1.0 M HCl

bromthymol | blue
yellow ← | → blue
6.8
∴ adding a base

15. The chemical indicator thymol blue changes from yellow to blue as a result of the addition of

A. 1.0 M HNO_2
[✓ B.] 1.0 M K_2CO_3
C. 1.0 M NH_4Cl
D. 1.0 M HCl

thymol | blue
yellow ← | blue
8.8
∴ adding a base

16. The chemical indicator bromcresol green changes from blue to yellow as a result of the addition of

A. 1.0 M LiCl
B. 1.0 M K_2CO_3
C. 1.0 M $NaNO_2$
[✓ D.] 1.0 M HCl

brom cresol | green
yellow ← | → blue
4.6
∴ add an acid

17. The chemical indicator phenol red changes from red to yellow as a result of the addition of

A. 1.0 M $NaNO_2$
B. 1.0 M K_2CO_3
C. 1.0 M LiCl
[✓ D.] 1.0 M HI

phenol | red
yellow ← | → red
7.3
add acid

18. Consider the following equilibrium for the chemical indicator phenol red, HInd, at a pH = 7.3 (orange).

$$\overset{\longleftarrow}{\overset{\uparrow}{\underset{\text{yellow}}{HInd}} + H_2O \leftrightarrows \overset{\uparrow}{H_3O^+}} + \underset{\downarrow}{\underset{\text{red}}{Ind^-}}$$

When some HCl is added, what stress is imposed on the equilibrium and what colour change occurs?

	Stress	Indicator Colour Change
A.	~~decreased~~ [H_3O^+]	turns ~~red~~
B.	increased [H_3O^+]	turns ~~red~~
C.	~~decreased~~ [H_3O^+]	turns yellow
[✓ D.]	increased [H_3O^+]	turns yellow

19. The indicator phenol red will be red in which of the following solutions?

A. 1.0 M H_2SO_3 Ka
B. 1.0 M NH_4NO_3 Ka
[✓ C.] 1.0 M Na_2HPO_4 Kb > Ka
D. 1.0 M $NaHSO_4$ Ka

phenol | red
yellow ← | → red
7.3
is base

20. The indicator phenolphthalein can be described by the following equilibrium equation:

$$\underset{\text{colorless}}{HIn} + \overset{\longleftarrow}{H_2O \leftrightharpoons} \overset{\uparrow}{H_3O^+} + \underset{\text{pink}}{In^-}$$

Ka is acidic

NH_4Cl is added to a slighly pink sample of this indicator. After equilibrium has been re-established, how do the $[H_3O^+]$ and the colour of the solution compare with the original equilibrium?

	$[H_3O^+]$	Colour of solution
[✓ A.]	increases	turns colourless
B.	de~~cr~~eases	turns ~~more~~ pink
C.	de~~cr~~eases	turns colourless
D.	increases	turns ~~more~~ pink

21. Consider the following indicator equilibrium:

$$\underset{\text{(yellow)}}{HIn} + \overset{\longrightarrow}{H_2O \leftrightharpoons} \overset{}{H_3O^+}\downarrow + \underset{\text{(blue)}}{In^-}$$

What is the result of adding Na_2CO_3 to this indicator? base

	Equilibrium Shift	Colour
A.	~~left~~	~~yellow~~
[✓ B.]	right	blue
C.	right	~~yellow~~
D.	~~left~~	blue

22. Consider the equilibirum for the indicator, thymolphthalein (HThy):

$$HThy + \overset{\longrightarrow}{H_2O \leftrightharpoons} H_3O^+\downarrow + Thy^-$$

colourless ← | blue.
10.0

What happens when NaOH is added to a sample of this indicator in water?

	Equilibrium	Colour
A.	shifts right	turns col~~o~~urless
[✓ B.]	shifts right	turns blue
C.	shifts ~~left~~	turns blue
D.	shifts ~~left~~	turns col~~o~~urless

23. A chemical indicator has a $K_a = 4.0 \times 10^{-6}$. What is the pH at the transition point and what is the identity of the indicator?

so pH = 5.39

	pH	Indicator
A.	~~8.6~~	thymol blue
B.	~~8.6~~	phenophthalien
C.	5.4	bromcresol green
[✓ D.]	5.4	methyl red

24. A chemical indicator has a $K_a = 1.6 \times 10^{-7}$. What is the pH at the transition point and what is the identity of the indicator?

pH = 6.80

	pH	Indicator
A.	6.8	phenol red
[✓ B.]	6.8	bromthymol blue
C.	~~7.2~~	thymol blue
D.	6.8	chlorophenol red

unnecessary

25. An indicator changes colour when 4.0 M HCl is added. If the indicator has a $K_a = 1 \times 10^{-10}$, identify the indicator and the pH at its transition point.

pH = 10.0

	Indicator	pH
A.	thymolphthalein	~~4.0~~
B.	phenolphthalein	~~4.0~~
C.	phenolphthalein	10.0
[✓ D.]	thymolphthalein	10.0

26. An indicator is blue at pH of 7.8 and yellow at a pH of 5.6. Identify the indicator and determine its K_a.

yellow ← | → blue

6.8 8.8

	Indicator	K_a
A.	thymol blue both yellow	2×10^{-9}
B.	bromcresol green both blue	3×10^{-5}
[✓ C.]	bromthymol blue	2×10^{-7}
D.	thymol blue	1×10^{-2}

27. What is one of the Ka values for thymol blue?

A. 6×10^{-2}

[✓ B.] 2×10^{-9}

C. 2×10^{-7}

D. 1×10^{-7}

pH = 2.0 Ka = 1.0×10^{-2}

pH = 8.8 Ka = 2×10^{-9}

28. An indicator is often used during acid-base titrations.

a. Define the term ***transition point*** for an indicator.

point where $[HInd] = [Ind^-]$

b. Calculate the K_a value for methyl red.

transition pt pH = 5.4 so $Ka = 4 \times 10^{-6}$

c. A mixture of indicators is made by combining equal amounts of methyl orange and bromthymol blue. Complete the following table showing the colour of each indicator and the mixture at the pH's indicated.

pH	3.8 Colour of methyl orange	6.8 Colour of bromthymol blue	Colour of mixture
pH = 5	yellow	yellow	yellow
pH = 9	yellow	blue	green.

29. An indicator is often used during acid-base titrations.

a. Calculate the K_a value for phenol red.

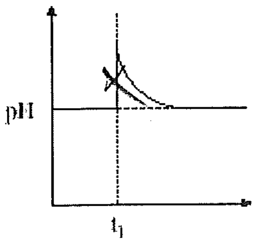
transition pt pH = 7.3 $Ka = 5 \times 10^{-8}$

b. A mixture of indicators is made by combining equal amounts of methyl orange, phenol red and chlorophenol red. Complete the following table showing the colour of each indicator and the mixture at the pH's indicated.

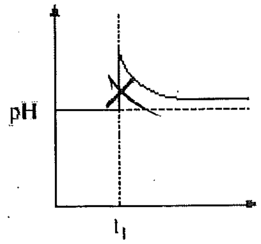
pH	2.0 8.8 Colour of thymol blue	7.3 Colour of phenol red	Colour of mixture
pH = 1.4	red	yellow	orange
pH = 7.8	yellow	red	orange
pH = 10.0	blue	red	purple

30. Which of the following graphs best describes the effect on the pH of a buffer solution with a small amount of acid is added at time t_1?

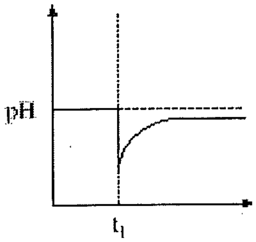
A.



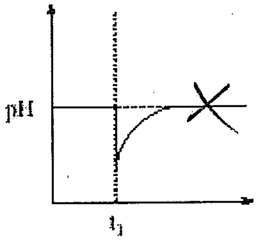
B.



[✓ C.]

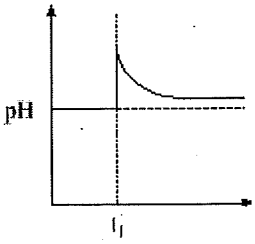


D.

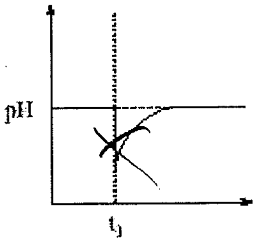


31. Which of the following graphs best describes the effect on the pH of a buffer solution with a small amount of base is added at time t_1?

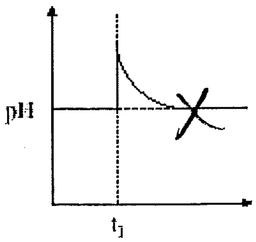
[✓ A.]



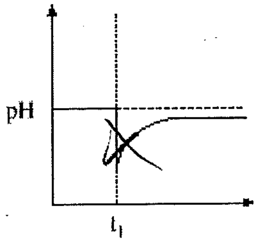
B.



C.



D.



32. a. What is the main function of a buffer solution?

keep pH relatively constant

b. Describe how you would prepare a buffer solution.

significant amounds of weak acid and its conjugated base.

33. Which of the following would form a buffer solution when equal moles are mixed together?

A. $\cancel{Na_2S}O_4$ and $N\cancel{aO}H$
B. $\cancel{HCl}$ and NaCl
C. $K\cancel{NO}_3$ and KOH
[✓ D.] HCN and NaCN

34. Equal moles of which of the following chemicals could be used to make a buffer solution that has a pH >7.0?

$Ka = [H_3O^+]$

A. ~~HCl and NaCl~~
[✓ B.] HCN and NaCN pH= 9. —
C. $K\cancel{Br}$ and $N\cancel{aN}O_3$
D. HF and NaF pH= 3. —

35. Equal moles of which of the following chemicals could be used to make a buffer solution with a pH <7.0?

[✓ A.] HNO_2 and $NaNO_2$ pH= 3. —
B. $NaHCO_3$ and Na_2CO_3 pH= 10. —
C. $HC\cancel{lO_4}$ and $Na\cancel{Cl}O_4$
D. H_2CO_3 and $N\cancel{aN}O_3$

36. Which of the following pairs of chemicals could be used to make a buffer solution?

[✓ A.] NH_3 and NH_4Cl
B. NH_3 and $\cancel{H_2}O$
C. $\cancel{HCl}$ and $N\cancel{a}Cl$
D. CH_3COOH and $\cancel{HCl}$

37. Which of the following pairs of chemicals could be used to make a buffer solution?

[✓ A.] HCOOH and NaHCOO
B. NH_3 and $N\cancel{aO}H$
C. HCN and $N\cancel{aO}H$
D. NH_3 and $H\cancel{Cl}$

38. Which of the following could tpically be used to prepare a buffer solution?
   A. HNO_2 and Na~~NO_3~~
   [✓ B.] H_2S and NaHS
   C. H_2S and ZnS
   D. HNO_3 and $NaNO_3$

39. Consider the following buffer equilibrium:

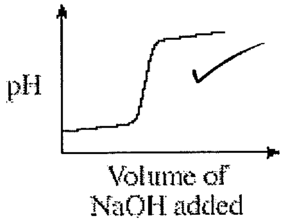
$$\underset{\downarrow}{HCN} + H_2O \overset{\longrightarrow}{\leftrightarrows} \overset{\uparrow}{CN^-} + \underset{\downarrow}{H_3O^+}$$

When a few drops of KOH are added the buffer, the equilibrium
   [✓ A.] shifts right and the [CN^-] increases
   B. shifts ~~left~~ and the [CN^-] ~~decreases~~
   C. shifts right and the [CN^-] ~~decreases~~
   D. shifts ~~left~~ and the [CN^-] increases

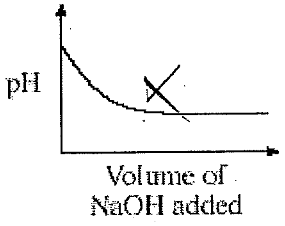
40. Which of the following graphs describes the relationship between the pH of a buffer and the volume of NaOH added to the buffer?

adding NaOH makes pH ↑

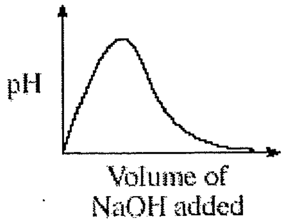
[✓ A.]



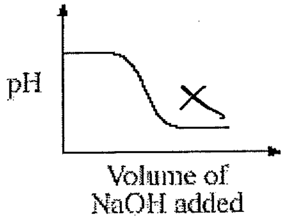
B.



C.



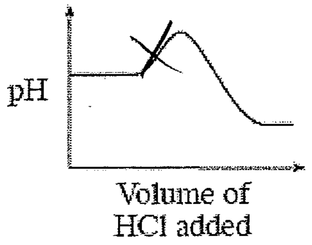
D.



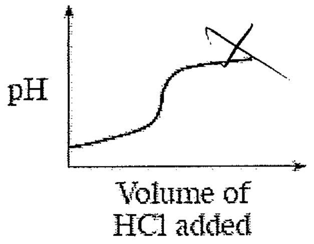
41. Which of the following graphs best describes the changes in pH when HCl is added to a buffer solution?

adding acid makes pH ↓

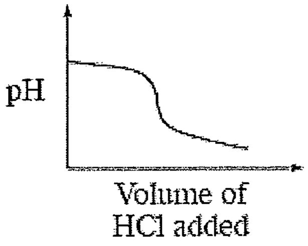
A.



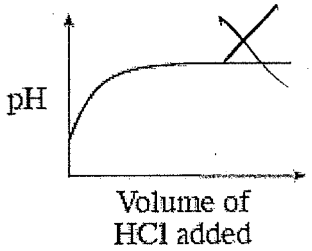
B.



[✓ C.]



D.



do in class

42. Acid is added to a buffer solution. When equilibrium is reestablished the buffering effect has resulted in $[H_3O^+]$

A. decreasing considerably
[✓ B.] increasing slightly
C. decreasing slightly
D. increasing considerably

do in class

43. A few drops of KOH are added to a buffer solution. When equilibrium is reestablished the buffering effect has resulted in pH

A. decreasing considerably
B. decreasing slightly
[✓ C.] increasing slightly
D. increasing considerably

44. Consider the buffer equilibrium:

$$\overset{\longleftarrow}{\underset{\text{high concentration}}{HNO_{2\,(aq)}} + H_2O_{(l)}} \leftrightarrows \underset{\text{low concentration}}{H_3O^{+}{}_{(aq)}}\uparrow + \underset{\text{high concentration}}{NO_2^{-}{}_{(aq)}}\downarrow$$

What happens when a small amount of $HCl_{(aq)}$ is added to the equilibrium system?

[✓ A.] the pH decreases slightly
B. the pH increases slightly
C. the equilibrium does not shift due to the levelling effect
D. the equilibrium shifts to the right

45. Consider the buffer equilibrium:

$HNO_{2\,(aq)}$	$+\ H_2O_{(l)}$	$\overset{\longrightarrow}{\leftrightarrows}$	$H_3O^+_{(aq)}\downarrow$	$+$	$NO_2^-{}_{(aq)}$
high concentration			low concentration		high concentration

What happens when a small amount of KOH $_{(aq)}$ is added to the equilibrium system?

[✓ A.] the pH increases slightly
B. the equilibrium shifts to the right
C. the equilibrium does not shift due to the levelling effect
D. the pH decreases slightly

46. Consider the buffer equilibrium:

$HNO_{2\,(aq)}$	$+\ H_2O_{(l)}$	$\overset{\longrightarrow}{\leftrightarrows}$	$H_3O^+_{(aq)}\downarrow$	$+$	$NO_2^-{}_{(aq)}$
high concentration			low concentration		high concentration

What happens when a small amount of $Na_2CO_{3\,(aq)}$ is added to the equilibrium system?

[✓ A.] the pH increases slightly
B. the equilibrium does not shift due to the levelling effect
C. the pH decreases slightly
D. the equilibrium shifts to the left

47. Consider the following buffer equilibrium:

$$H_2CO_{3\,(aq)} + H_2O_{(l)} \overset{\longrightarrow}{\leftrightarrows} H_3O^+_{(aq)}\downarrow + HCO_3^-{}_{(aq)}$$

What happens when a small amount of NaOH (aq) is added?

A. $[H_3O^+]$ decreases, then the equilibrium shifts to the ~~left~~.
B. $[H_3O^+]$ in~~cr~~eases, then the equilibrium shifts to the ~~left~~.
[✓ C.] $[H_3O^+]$ decreases, then the equilibrium shifts to the right.
D. $[H_3O^+]$ in~~cr~~eases, then the equilibrium shifts to the right.

48. In the human bloodstream, a buffer exists that is made of H_2CO_3 and $NaHCO_3$.

do in class

a. Explain what the purpose for this buffer is:

Keep the pH relatively constant

b. Approximately what pH level would this buffer operate at? Assume that there are equal moles of H_2CO_3 and $NaHCO_3$.

$K_a = \frac{[H\cancel{CO_3^-}][H_3O^+]}{[H_2\cancel{CO_3}]}$ $\quad 4.3 \times 10^{-7} = [H_3O^+]$

pH = 6.37

c. When a person exercises strenuously, the muscles produce lactic acid as a waste product. After strenuous exercise, that acid would make its way into the blood stream. What would happen to the pH of the blood?

$H_2CO_3 + H_2O \rightleftharpoons HCO_3^- + \boxed{H_3O^+}$

decrease, slightly

49. A scientist wants a buffer solution that will work at a pH level of 3.75.

a. Describe what would be required to make a suitable buffer solution.

if pH = 3.75 then $[H_3O^+] = K_a = 1.8 \times 10^{-4}$

$\therefore$ acid with $K_a = 1.8 \times 10^{-4}$

b. Which weak acid and conjugate base would work? HCOOH and its conjugate base $HCOO^-$

c. Explain what would happen if a few drops of NaOH would be added to this buffer. Would the pH change? If so, how much and would it increase or decrease?

pH would increase slightly.

Explanation: stress $[H_3O^+] \downarrow$ $\therefore$ shift to the right to minimize stress

or stress $[OH^-] \uparrow$ $\therefore$ shift to the left to minimize stress

50. Consider the following buffer equilibrium:

$$\underset{\text{high concentration}}{HF_{(aq)}} + H_2O_{(l)} \leftrightarrows \underset{\text{low concentration}}{H_3O^+{}_{(aq)}} + \underset{\text{high concentration}}{F^-{}_{(aq)}}$$

Using LeChatelier's Principle, explain what happens to the pH of the buffer solution when a small amount of NaOH is added.

pH ↑ slightly   stress $[H_3O^+]$ ↓

so shift to the right to minimize the stress.

If equal moles of HF and F⁻ are used, what will be the approximate pH level that this buffer will work at?

$$K_a = \frac{[\cancel{F^-}][H_3O^+]}{[\cancel{HF}]} \qquad 3.5 \times 10^{-4} = [H_3O^+]$$

pH = 3.46

51. State the sequence of events that occur when a small amount of HCl (aq) is added to a buffer such as:

$$NH_{3\,(aq)} + H_2O_{(l)} \overset{\longrightarrow}{\leftrightarrows} NH_4^+{}_{(aq)} + \underset{\downarrow}{OH^-}{}_{(aq)}$$

Be sure to describe the stress, the shift and the effect on the pH what occur.

Stress: $[OH^-]$

Shift: right

Effect on pH decrease slightly

52. If 1.00 moles of HCN and 1.00 moles of NaCN are added to 1.00 L of water, what pH will the buffer remain relatively constant at?

A. 0.00

B. 7.00

C. 4.69

[✓ D. 9.31]

53. If 1.00 moles of HF and 1.00 moles of NaF are added to 1.00 L of water, what pH will the buffer remain relatively constant at?
A. 7.00
B. 10.54
C. 0.00
[✓ D. 3.46]

54. If 1.00 moles of CH_3COOH and 1.00 moles of $NaCH_3COO$ are added to 1.00 L of water, what pH will the buffer remain relatively constant at?
A. 0.00
B. 7.00
C. 9.26
[✓ D. 4.74]

55. If 1.00 moles of NH_4NO_3 and 1.00 moles of NH_3 are added to 1.00 L of water, what pH will the buffer remain relatively constant at?
A. 7.00
B. 4.75
C. 0.00
[✓ D. 9.25]

56. Consider the following buffer equilibrium:

$$HF_{(aq)} + H_2O_{(l)} \overset{\longleftarrow}{\leftrightarrows} \overset{\uparrow}{H_3O^+_{(aq)}} + \boxed{F^-_{(aq)}}$$

What would limit the buffering action if HCl were added?
A. [HF]
[✓ B. [F$^-$]]
C. [H_3O^+]
D. [H_2O]

57. Consider the following buffer equilibrium:

$$\boxed{HF_{(aq)}} + H_2O_{(l)} \overset{\longrightarrow}{\leftrightarrows} \underset{\downarrow}{H_3O^+_{(aq)}} + F^-_{(aq)}$$

What would limit the buffering action if KOH were added?
A. [F$^-$]
B. [H_3O^+]
[✓ C. [HF]]
D. [H_2O]

58. Consider the following buffer equilibrium:

$$HCN + H_2O \leftrightarrows CN^- + H_3O^+$$

When 25 mL of 0.200 M KOH are added, the pH rises dramatically. Why?
A. The KOH is a strong base and forces the CN^- to act as an acid.
B. The KOH becomes part of the buffer solution.
C. The KOH reacts with the HCN instead of the H_3O^+, causing a shift left instead of a shift right.
[✓ D.] The KOH exceeds the buffer capacity.

59. In acid-base titrations, the solution of known concentration is called a (an)
   - A. indicating solution
   - B. basic solution
   - [✓ C. standard solution]
   - D. acidic solution

60. *At a certain point in a strong acid-strong base titration, the moles of H^+ are equal to the moles of OH^-.* This is a definition of which of the following?
   A. titration point B. end point [✓ C. equivalence point] D. transition point

61. When performing a titration experiment, the indicator must always have
   - [✓ A. a transition point that is close to the equivilence point]
   - B. a distinct colour change a pH = 7.0
   - C. the ability to change from colourless to pink
   - D. an equivalence point that is close to the stoichiometric point

62. A weak acid is titrated with a strong base using the indicator phenophthalein to detect the end point. What is the approximate pH at the transition point?
   - [✓ A. 9.0]
   - B. 7.0
   - C. 8.0
   - D. 10.0

change to equivalence

63. Which of the following would be the net ionic equation for the reaction between HCl and KOH?
   - A. $H^+ + Cl^- + K^+ + OH^- \rightarrow H_2O + Cl^- + K^+$
   - [✓ B. $H^+ + OH^- \rightarrow H_2O$]
   - C. $HCl + KOH \rightarrow H_2O + KCl$
   - D. $H^+ + Cl^- + K^+ + OH^- \rightarrow H_2O + KCl$

64. Which of the following is the net ionic equation for the neutralization of CH_3COOH with $NaOH?_{(aq)}$
   - A. $CH_3COOH_{(aq)} + H^+_{(aq)} + Na^+_{(aq)} + OH^-_{(aq)} \rightarrow Na^+ + CH_3COO^-_{(aq)} + H_2O_{(l)}$
   - [✓ B. $CH_3COOH_{(aq)} + OH^-_{(aq)} \rightarrow CH_3COO^-_{(aq)} + H_2O_{(l)}$]
   - C. $CH_3COO^-_{(aq)} + OH^-_{(aq)} \rightarrow CH_3COOH_{(aq)} + O^{-2}_{(aq)}$
   - D. $CH_3COOH_{(aq)} + NaOH_{(aq)} \rightarrow NaCH_3COO^-_{(aq)} + H_2O_{(l)}$

WA

65. Write the net ionic equation for the neutralization of $HF_{(aq)}$ with $Sr(OH)_{2\,(aq)}$.
   - A. $HF_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)} + F^-_{(aq)}$
   - [✓ B. $2HF_{(aq)} + Sr^{+2}_{(aq)} + 2\,OH^-_{(aq)} \rightarrow SrF_{2\,(s)} + 2H_2O_{(l)}$]
   - C. $HF_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + F^-_{(aq)}$
   - D. $2HF_{(aq)} + Sr(OH)_{2\,(aq)} \rightarrow SrF_{2\,(s)} + 2\,H_2O_{(l)}$

$SrF_{2\,(s)}$ low solubility

66. Write the net ionic equation for the neutralization of $H_2S_{(aq)}$ with $Sr(OH)_{2\,(aq)}$.

A. $H_2S_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + HS^-_{(aq)}$

B. $H_2S_{(aq)} + Sr(OH)_{2\,(aq)} \rightarrow SrS_{(aq)} + 2\,H_2O_{(l)}$

C. $2H^+_{(aq)} + S^{-2}_{(aq)} + Sr^{+2}_{(aq)} + 2\,OH^-_{(aq)} \rightarrow SrS_{(s)} + 2H_2O_{(l)}$

[✓ D.] $H_2S_{(aq)} + 2OH^-_{(aq)} \rightarrow 2H_2O_{(l)} + S^{-2}_{(aq)}$

67. Write the net ionic equation for the neutralization of H~~Br~~ $_{(aq)}$ with ~~Sr~~$(OH)_{2\,(aq)}$.

A. $2H^+_{(aq)} + 2\,Br^-_{(aq)} + Sr^{+2}_{(aq)} + 2\,OH^-_{(aq)} \rightarrow SrBr_{2\,(aq)} + 2H_2O_{(l)}$

[✓ B.] $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

C. $HBr_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)} + Br^-_{(aq)}$

D. $2HBr_{(aq)} + Sr(OH)_{2\,(aq)} \rightarrow SrBr_{2\,(aq)} + 2\,H_2O_{(l)}$

68. What is the formula equation for the neutralization of 0.1 M $Sr(OH)_2$ (aq) with 0.1 M H_2SO_4 (aq)?

[✓ A.] $Sr(OH)_{2\,(aq)} + H_2SO_{4\,(aq)} \rightarrow SrSO_{4\,(s)} + 2H_2O_{(l)}$

B. $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

C. $Sr^{+2}_{(aq)} + SO_4{}^{-2}_{(aq)} \rightarrow SrSO_{4\,(s)}$

D. $Sr^{+2}_{(aq)} + 2OH^-_{(aq)} + 2H^+_{(aq)} + SO_4{}^{-2}_{(aq)} \rightarrow SrSO_{4\,(s)} + 2\,H_2O_{(l)}$

69. The strong acid, $HNO_{3\,(aq)}$ is titrated with the weak base, $NH_{3\,(aq)}$. What is the net ionic equation for this reaction?

A. $H^+_{(aq)} + NO_3^-{}_{(aq)} + NH_3\ (aq) \rightarrow NH_4{}^+_{(aq)} + NO_3^-{}_{(aq)}$

[✓ B.] $H^+_{(aq)} + NH_{3\,(aq)} \rightarrow NH_4^+{}_{(aq)}$

C. $H^+_{(aq)} + OH^- \rightarrow H_2O_{(l)}$

D. $HNO_{3\,(aq)} + NH_{3\,(aq)} \rightarrow NH_4NO_{3\,(aq)}$

70. Which of the following is the complete ionic equation for the titration of $HCl_{(aq)}$ with KOH (aq)?

A. $HCl_{(aq)} + KOH_{(aq)} \rightarrow KCl_{(aq)} + H_2O_{(l)}$

B. $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

[✓ C.] $H^+_{(aq)} + Cl^-_{(aq)} + K^+_{(aq)} + OH^-_{(aq)} \rightarrow K^+_{(aq)} + Cl^-_{(aq)} + H_2O_{(l)}$

D. $H^+_{(aq)} + Cl^-_{(aq)} + K^+_{(aq)} + OH^-_{(aq)} \rightarrow KCl_{(aq)} + H_2O_{(l)}$

71. What is the net ionic equation for the titration of $\overset{WA}{H_3PO_{4}}{}_{(aq)}$ with $\overset{SB}{Sr(OH)_{2}}{}_{(aq)}$?

A. $6H^+_{(aq)} + 2PO_4{}^{-3}{}_{(aq)} + 3Sr^{+2}_{(aq)} + 6OH^-_{(aq)} \rightarrow 3Sr^{+2}_{(aq)}\ \ 2PO_4{}^{-3}{}_{(aq)} + 6H_2O_{(l)}$

B. $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

C. $6H^+_{(aq)} + 6OH^-_{(aq)} \rightarrow 6H_2O_{(l)}$

[✓ D.] $2H_3PO_{4\,(aq)} + 3Sr^{+2}_{(aq)} + 6OH^-_{(aq)} \rightarrow Sr_3(PO_4)_{2\,(s)} + 6H_2O_{(l)}$

72. Write the net ionic equation for the acid-base reaction that occurs between $NaNO_2$ (aq) and NH_4Cl (aq).

$NO_2^- + NH_4^+ \rightarrow NH_3 + HNO_2$

73. For the titration between 0.20 M $Sr(OH)_2$ and 0.20 M H_2SO_3

do in class.

a. Write the formula equation

$Sr(OH)_2 + H_2SO_3 \rightarrow SrSO_3 + 2H_2O$

b. Write the complete ionic equation

$Sr^{+2} + 2OH^- + H_2SO_3 \rightarrow SrSO_{3(s)} + 2H_2O(l)$

c. Write the net ionic equation

same

d. Explain why the electrical conductivity of the products is less than that of the reactants

fewer ions as products than of reactants

74. Which statement describes the pH of the equivilence point of a titration of 0.200 M CH_3COOH by 0.200 M KOH?

[✓ A.] The pH > 7 because the KCH_3COO hydrolizes to form OH^-.
B. The pH = 7 because the CH_3COOH and KOH neutralize each other.
C. The pH = 7 because the no hydrolysis of products is possible.
D. The pH < 7 because the KCH_3COO hydrolizes to form H_3O^+.

75. Which statement describes the pH of the equivilence point of a titration of 0.200 M NH_3 by 0.200 M HI?

[✓ A.] The pH < 7 because the NH_4^+ hydrolizes to form H_3O^+.
B. The pH > 7 because the NH_4^+ hydrolizes to form OH^-.
C. The pH = 7 because the no hydrolysis of products is possible.
D. The pH = 7 because the NH_3 and HI neutralize each other.

76. What is the $[H_3O^+]$ at the equivilence point for the titration between HBr and KOH?

A. 0.0 M
B. 1.0 x 10^{-9} M
C. 1.0 x 10^{-5} M
[✓ D.] 1.0 x 10^{-7} M

77. What is the pH at the equivilence point for the titration between HF and KOH?
A. 0.0
B. 6.25
[✓ C. 7.75]
D. 7.00

78. What is the pH at the equivilence point for the titration between HBr and NH_3?
[✓ A. 6.25]
B. 7.75
C. 0.0
D. 7.00

79. At the equivilence point, the titration of HCN with NaOH will form a solution which is
A. acidic with pH < 7
B. basic with pH < 7
[✓ C. basic with pH >7]
D. neutral with pH = 7

80. At the equivilence point, the titration of HCl with NaOH will form a solution which is
A. basic with pH >7
[✓ B. neutral with pH = 7]
C. acidic with pH < 7
D. basic with pH < 7

81. Which of the following titrations would have a pH >7 at the equivilence point?
A. HI with KOH
[✓ B. HCOOH with NaOH]
C. $HClO_4$ with NH_3
D. HCl with $Sr(OH)_2$

82. Which of the following titrations would have a pH < 7 at the equivilence point?
A. HCl with $Sr(OH)_2$
[✓ B. $HClO_4$ with NH_3]
C. HCOOH with NaOH
D. HI with KOH

83. Which of the following titrations always results in pH = 7.0 at the equivilence point?
A. A strong acid is titrated with a weak base.
B. A weak acid is titrated with a weak base.
C. A weak acid is titrated with a strong base.
[✓ D. A strong acid is titrated with a strong base.]

84. The following two experiments were conducted:

Titration A: A strong acid was titrated with a 0.20 M strong base.

Titration B: A weak acid was titrated with a 0.20 M strong base.

a. How does the pH at the equilivence point of Titration B compare with the pH at the equivilence point of Titration A?

Titration A pH = 7 Titration B pH > 7

b. Explain your answer

The salt produced in a strong acid/strong base titration is neutral
the salt produced in a weak acid/strong base is basic.

c. How will the volume of strong base needed to reach equivilence point in Titration B compare with the volume of strong base needed to reach equivilence point in Titration A?

Same for both

85. The following data were recorded when titrating 25.00 mL of CH_3COOH with 0.200 M $Ba(OH)_2$.

Volume of $Ba(OH)_2$ added

	Trial #1	~~Trial #2~~	Trial #3
Initial reading	4.03 mL	~~17.51 mL~~	32.03 mL
Final reading	17.51 mL	~~32.03 mL~~	45.55 mL
	13.48 mL	~~14.52 mL~~	13.52

Ave = 13.50 mL

Calculate the $[CH_3COOH]$

$Ba(OH)_2 + 2CH_3COOH \rightarrow$

A. 0.185 M
[✓ B.] 0.216 M
C. 0.108 M
D. 0.370 M

$$13.50 \times 10^{-3} L \times \frac{.200 \text{ mol}}{L} \text{ of } Ba(OH)_2 \times \frac{2 CH_3COOH}{1 Ba(OH)_2} \times \frac{1}{25.00 \times 10^{-3}}$$

$= .216$ M

86. During a titration, what volume of 0.500 M KOH would be necessary to completely neutralize 10.0 mL of 2.00 M CH_3COOH?

[✓ A. 40.0 mL]
B. 10.0 mL
C. 20.0 mL
D. 25.0 mL

$CH_3COOH + KOH \rightarrow$

$10.0 \times 10^{-3} L \times \frac{2.00 \text{ mol of } CH_3COOH}{L} \times \frac{1 \text{ mole KOH}}{1 \text{ mole } CH_3COOH} \times \frac{L}{.500}$

$= 40$

87. Calculate the volume of 0.500 M NaOH required to completely neutralize 25.0 mL of 0.450 M H_2SO_4.

in class

[✓ A. 45.0 mL]
B. 9.00 mL
C. 11.3 mL
D. 22.5 mL

$2 NaOH + H_2SO_4 \rightarrow$

$25.0 \times 10^{-3} L \times \frac{.450 \text{ mol of } H_2SO_4}{L} \times \frac{2 \text{ mole NaOH}}{1 \text{ mole } H_2SO_4} \times \frac{L}{.500 \text{ mol}}$

$= 45$

88. Calculate the volume of 0.300 M HNO_3 needed to completely neutralize 25.0 mL of 0.250 M $Sr(OH)_2$.

A. 10.4 mL
[✓ B. 41.7 mL]
C. 15.0 mL
D. 20.8 mL

$2HNO_3 + Sr(OH)_2 \rightarrow$

$25.0 \times 10^{-3} L \times \frac{.250 \text{ mol of } Sr(OH)_2}{L} \times \frac{2 \text{ moles } HNO_3}{1 \text{ mole } Sr(OH)_2} \times \frac{L}{.300 \text{ mol}}$

$= 41.66$ mL

89. During a titration, what volume of 0.500 M $Ba(OH)_2$ would be necessary to completely neutralize 10.0 mL of 2.00 M CH_3COOH?

in class

[✓ A. 20.0 mL]
B. 25.0 mL
C. 40.0 mL
D. 10.0 mL

$2CH_3COOH + Ba(OH)_2$

$10.0 \times 10^{-3} L \times \frac{2.00 \text{ mol of } CH_3COOH}{L} \times \frac{1 \text{ mole } Ba(OH)_2}{2 \text{ moles } CH_3COOH} \times \frac{L}{.500 \text{ mol}}$

$= 20$

90. What volume of 0.500 M NaOH is required to neutralize 25.0 mL of a 0.250 M HBr?

[✓ A. 12.5 mL]
B. 5.00 mL
C. 25.0 mL
D. 20.0 mL

$NaOH + HBr \rightarrow$

$25.0 \times 10^{-3} L \times \frac{.250 \text{ mol of HBr}}{L} \times \frac{1 \text{ mole NaOH}}{1 \text{ mole HBr}} \times \frac{L}{.500 \text{ mol}}$

91. The complete neutralization of 15.0 mL of KOH requires 0.0250 moles of H_2SO_4. The [KOH] was

[✓ A. 3.33 M]
B. 1.67 M
C. 0.833 M
D. 3.75×10^{-4} M

$H_2SO_4 + 2KOH \rightarrow$

$.0250 \text{ moles of } H_2SO_4 \times \frac{2 \text{ moles KOH}}{1 \text{ mole } H_2SO_4} \times \frac{1}{15.0 \times 10^{-3} L} =$

92. What volume of 0.100 M H_2SO_4 is needed to titrate 25.0 mL of 0.200 M NaOH?

A. 100.0 mL
[✓ B. 25.0 mL]
C. 12.5 mL
D. 50.0 mL

$25.0 \times 10^{-3} L \times \frac{.200 \text{ mol of NaOH}}{L} \times \frac{1 \text{ mole } H_2SO_4}{2 \text{ moles NaOH}} \times \frac{L}{.100 \text{ mol}}$

$= 25$

$Ba(OH)_2 + 2HNO_3$

93. How many moles of $Ba(OH)_2$ are required to react completely with 100.0 mL of 0.250 M HNO_3

A. 0.0250 moles
B. 0.0500 moles
C. 1.25 moles
[✓ D. 0.0125 moles]

$$100.0\times10^{-3}\,L \times \frac{.250\text{ moles of }HNO_3}{L} \times \frac{1\text{ mole }Ba(OH)_2}{2\text{ moles }HNO_3} = .0125\text{ moles}$$

94. A 10.0 mL sample of H_2SO_3 is completely neutralized by titration with 18.6 mL of 0.10 M NaOH. Calculate the concentration of the acid.

A. 0.74 M
[✓ B. 0.093 M]
C. 0.19 M
D. 0.37 M

$$H_2SO_3 + 2NaOH \rightarrow$$

$$18.6\times10^{-3}\,L \times \frac{.10\text{ moles of }NaOH}{L} \times \frac{1\text{ mole }H_2SO_3}{2\text{ moles }NaOH} \times \frac{1}{10.0\times10^{-3}\,L} = .093$$

95. During a titration, 25.0 mL of H_3PO_4 (aq) is completely neutralized by 42.6 mL of 0.20 M NaOH. Calculate the concentration of the acid.

A. 1.0 M
B. 0.17 M
[✓ C. 0.11 M]
D. 0.34 M

$$H_3PO_4 + 3NaOH \rightarrow$$

$$42.6\times10^{-3}\,L \times \frac{.20\text{ mol of }NaOH}{L} \times \frac{1\text{ mole }H_3PO_4}{3\text{ mole }NaOH} \times \frac{1}{25.0\times10^{-3}} = .1136\text{ M}$$

96. A 20.0 mL sample of HCl is titrated with 25.0 mL of 0.20 M $Sr(OH)_2$. What is the concentration of the acid?

A. 0.20 M
B. 0.13 M
[✓ C. 0.50 M]
D. 0.25 M

$$Sr(OH)_2 + 2HCl \rightarrow$$

$$25.0\times10^{-3}\,L \times \frac{.20\text{ mol of }Sr(OH)_2}{L} \times \frac{2\text{ moles }HCl}{1\text{ mole }Sr(OH)_2} \times \frac{1}{20.\ \times10} = .50\text{M}$$

97. A 25.0 mL sample of H_2SO_3 is titrated with 20.0 mL of 0.150 M NaOH. Calculate the concentration of the H_2SO_3.

[✓ A. 0.0600 M]
B. 0.120 M
C. 0.240 M
D. 0.00300 M

$$H_2SO_3 + 2NaOH \rightarrow$$

$$20.0\times10^{-3}\,L \times \frac{.150\text{ mol of }NaOH}{L} \times \frac{1\text{ mole }H_2SO_3}{2\text{ mole }NaOH} \times \frac{1}{25.00\times10} = .060\text{ M.}$$

98. What volume of 0.250M KOH is required to titrate 2.30×10^{-3} mol of the weak acid $H_2C_2O_4$?

[✓ A. 18.4 mL]
B. 1.15 mL
C. 4.60 mL
D. 9.20 mL

$$H_2C_2O_4 + 2KOH \rightarrow$$

$$2.30\times10^{-3}\text{ moles of }H_2C_2O_4 \times \frac{2\text{ moles }KOH}{1\text{ mole }H_2C_2O_4} \times \frac{L}{.250\text{ mol}} = .0184\text{ L}$$

99. How many moles of KOH are necessary to completely neutralize 42.0 mL of 3.00 M HCl?

[✓ A. 0.126 moles]
B. 0.252 moles
C. 3.00 moles
D. 0.0140 moles

$$KOH + HCl \rightarrow$$

$$42.0\times10^{-3}\,L \times \frac{3.00\text{ mol of }HCl}{L} \times \frac{1\text{ mole }KOH}{1\text{ mole }HCl} \approx .126$$

100. A 25.0 mL sample of a diprotic weak acid is titrated with 20.2 mL of 0.10 M NaOH. What is the concentration of the acid?

A. 0.12 M

B. 0.080 M

[✓ C. 0.040 M]

D. 0.16 M

$H_2X + 2NaOH \rightarrow$

$20.2 \times 10^{-3} L \times \frac{.10 \text{ moles of NaOH}}{L} \times \frac{1 \text{ mole } H_2X}{2 \text{ moles NaOH}} \times \frac{\quad}{25.0 \times 1}$

$= .0404$ M

101. A titration was performed by adding 0.115 M HCl to a 25.00 mL sample of $Ba(OH)_2$. Calculate the $[Ba(OH)_2]$ from the following data:

	Trial #1	~~Trial #2~~	Trial #3
Inital volume of HCl (mL)	4.00	~~22.45~~	3.45
Final volume of HCl (mL)	22.45	~~42.85~~	22.00
	18.45	~~20.4~~	18.55

Ave = 18.50 mL

$2HCl + Ba(OH)_2 \rightarrow 2H_2O + BaCl_2$

$18.50 \times 10^{-3} L \times \frac{.115 \text{ mol of HCl}}{L} \times \frac{1 \text{ mole } Ba(OH)_2}{2 \text{ moles HCl}} \times \frac{1}{25.00 \times 10^{-3} L}$

$= .04255$ M

$= .0426$ M

do in class

change to $H_2C_2O_4 \cdot 2H_2O$

102. A solution of NaOH $_{(aq)}$ was standardized by titration using oxalic acid ($H_2C_2O_{4\,(s)}$) as the primary standard. The following data was collected:

Mass of $H_2C_2O_{4\,(s)}$ used = 1.02 g
Volume of NaOH $_{(aq)}$ used = 40.0 mL

use ~~mass~~ mass for $H_2C_2O_4 \cdot 2H_2O$

$H_2C_2O_4 + 2NaOH \rightarrow 2H_2O + Na_2C_2O_4$

Calculate the concentration of the NaOH $_{(aq)}$.

$$1.02\,g \text{ of } H_2C_2O_4 \cdot 2H_2O \times \frac{1 \text{ mole } H_2C_2O_4 \cdot 2H_2O}{126.0\,g \text{ of } H_2C_2O_4} \times \frac{2 \text{ moles } H_2C_2O_4}{1 \text{ mole } H_2C_2O_4} \times \frac{1}{40.0 \times 10^{-3}}$$

$= .4047619\ M$

$= .405\ M$

$\therefore [H_3O^+] = [HCl] = 1.00 \times 10^{-2} M$

103. A 250.0 mL sample of HCl with a pH of 2.000 is completely neutralized with 0.200 M NaOH.

a. What volume of NaOH is required to reach the ~~stoichiometric~~ point? change to equivalence pt

$$250.0 \times 10^{-3} L \times \frac{1.00 \times 10^{-2} \text{ moles of HCl}}{L} \times \frac{1 \text{ mole NaOH}}{1 \text{ mole HCl}} \times \frac{L}{.200 \text{ mol of NaOH}}$$

$= .0125\ L$ = 12.5 mL of NaOH

b. Write the net ionic equation for the neutralization.

c. If the HCl were titrated with 0.200 M $NH_{3\,(aq)}$ instead of 0.200 M NaOH, how would the volume of base required to reach the equivilence point compare with the volume calculated in part a) ? Explain your answer.

volume would be the same because in a titration there is no difference between strong & weak bases.

R6: Review

weak base strong base

104. The property common to both 0.10 M NH_3 and 0.10 M NaOH is that both solutions
   - A. react with magnesium to produce hydrogen gas
   - B. turn blue litmus paper ~~red~~
   - [✓ C.] have a pH > 7
   - D. disso~~ciate~~ 100%

105. A substance which absorbs hydrogen ions in solution is a definition of which of the following?
   - [✓ A.] ~~a Bronsted-Lowry base~~
   - B. a Bronsted-Lowry acid
   - C. an Arrhenius base
   - D. an Arrhenius acid

106. Consider the following Bronsted-Lowry equilibrium:

$$\overset{\text{base}}{C_6H_5NH_{2\,(aq)}} + \overset{\text{acid}}{H_2PO_4^-{}_{(aq)}} \leftrightarrows \overset{\text{acid}}{C_6H_5NH_3^+{}_{(aq)}} + \overset{\text{base.}}{HPO_4^{-2}{}_{(aq)}}$$

The substances acting as acids and bases from left to right are:
   - A. acid, base, base, acid
   - [✓ B.] base, acid, acid, base
   - C. base, acid, base, acid
   - D. acid, base, acid, base

107. Which of the following will have the lowest electrical conductivity?
   - A. 1.00 M NaCN salt
   - B. 1.00 M $NaHCO_3$ salt
   - [✓ C.] 1.00 M $H_2C_2O_4$ acid weak
   - D. 1.00 M $HClO_4$ acid

108. In a 1.0 M HF solution, the concentrations of HF, F^- and OH^-, from highest to lowest is
   - A. $[\cancel{OH}^-] > [F^-] > [HF]$
   - B. $[F^-] > [HF] > [OH^-]$
   - [✓ C.] $[HF] > [F^-] > [OH^-]$
   - D. $[\cancel{OH}^-] > [HF] > [F^-]$

$HF + H_2O \rightleftharpoons F^- + H_3O^+$ $\therefore [H_3O^+] > [OH^-]$

1.0 5% .05M .05M 2.0×10^{-13} M

109. The strength of the ions $HC_2O_4^-$, HSO_3^- and $H_2PO_4^-$ from weakest to strongest **base** is
   - A. $H_2PO_4^- < HSO_3^- < HC_2O_4^-$
   - B. $HC_2O_4^- < H_2PO_4^- < HSO_3^-$
   - [✓ C.] $HC_2O_4^- < HSO_3^- < H_2PO_4^-$
   - D. $HSO_3^- < H_2PO_4^- < HC_2O_4^-$

use pg 6

110. Which of the following is the strongest acid that can exist in an aqueous solution?

A. $HClO_4$ [✓ B. H_3O^+] C. NH_2^- D. O^{-2}

111. Consider the following equilibrium:

$$H_3AsO_4 + HSeO_4^- \rightleftarrows H_2AsO_4^- + H_2SeO_4$$

Products are favored in this equilibrium. Which of the following describes the relative strengths of the acids and bases?

	Stronger Acid	Stronger Base
A.	$H_2\cancel{Se}O_4$	$HSeO_4^-$
B.	$H_{\cancel{2}}\cancel{Se}O_4$	$H_2\cancel{As}O_4^-$
[✓ C.]	H_3AsO_4	$HSeO_4^-$
D.	H_3AsO_4	$\cancel{H_2AsO_4^-}$

typo error

112. When comparing 0.1 0 M HPO_4^{-2} and 0.10 M $HC_2O_4^-$ as stronger acids, which of the following is true?

A. HPO_4^{-2} is weaker and its pH is smaller ✗

B. $HC_2O_4^-$ is weaker and its pH is larger

C. HPO_4^{-2} is stronger and its pH is larger ✗

[✓ D.] $HC_2O_4^-$ is stronger and its pH is smaller

113. Consider the following equilibrium:

$$2H_2O_{(l)} + \overrightarrow{\text{energy}} \leftrightarrows H_3O^+_{(aq)}\downarrow + OH^-_{(aq)}\uparrow$$

A few drops of NaOH are added and a new equilibrium is established. The new equilibrium can be described by

A. pH = pOH and $K_w \cancel{<} 1.0 \times 10^{-14}$

B. $pH \cancel{<} pOH$ and $K_w = 1.0 \times 10^{-14}$

[✓ C.] $pH > pOH$ and $K_w = 1.0 \times 10^{-14}$

D. pH = pOH and $K_w \cancel{>} 1.0 \times 10^{-14}$

114. At 20°C the ionization constant of water (K_w) is 6.76×10^{-15}. Calculate the pOH and pH of water at 20°C.

	pH	pOH
A.	7.000	7.000
B.	7.085	6.915
[✓ C.]	7.085	7.085
D.	6.915	7.085

pH= pOH and neither are 7.00

115. Which of the following solutions will have a pH = 1.00?

I. 0.10 M HCl ✓
II. 0.10 M HNO_2 ✗
III. 0.10 M NaOH ✗

A. I, II and III B. III only [✓ C. I only] D. I and II only

116. Which of the following solutions would have a pH = 2.00?

[✓ A.] 0.010 M HCl
B. 0.010 M NaOH ✗
C. 0.010 M HCN ✗
D. 0.010 M H_2SO_4 ✗ you have H_3O^+ from both H_2SO_4 and HSO_4^-

117. Using calculations, show why the electrical conductivity of 1.0 M H_2CO_3 will be less than that for 0.10 M HCl. change to .0020M for next year.

$HCl \rightarrow H^+ + Cl^-$ ∴ [ions] = .20 M

.10M .10M .10M

$H_2CO_3 + H_2O \rightleftharpoons HCO_3^- + H_3O^+$

H_2CO_3	H_2O	HCO_3^-	H_3O^+
1.0		0	0
−x		+x	+x
1.0−x		x	x

$$Ka = \frac{[HCO_3^-][H_3O^+]}{[H_2CO_3]}$$

$$4.3x10^{-7} = \frac{x^2}{1.0-x}$$

Assume $1.0 - x \approx 1.0$

$4.3x10^{-7} = x^2$

$x = [HCO_3^-] = [H_3O^+] = 6.5574x10^{-4} M$

∴ [ions] = $1.3x10^{-3} M$

∴ way fewer ions than for HCl.

118. Calculate the pH of a 0.010 M NH_4CN solution.

NH_4^+ $Ka = 5.6x10^{-10}$

CN^- $Kb = \frac{1.0x10^{-14}}{4.9x10^{-10}} = 2.0x10^{-5}$ ✓

$CN^- + H_2O \rightleftharpoons HCN + OH^-$

CN^-	H_2O	HCN	OH^-
.010		0	0
−x		+x	+x
.010−x		x	x

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

$$\frac{1.0x10^{-14}}{4.9x10^{-10}} = \frac{x^2}{.010-x}$$

Assume $.010 - x \approx .010$

$$\frac{1.0x10^{-14}}{4.9x10^{-10}} = \frac{x^2}{.010}$$

$x^2 = 2.0408x10^{-7}$

$x = [OH^-] = 4.5175x10^{-4} M$

pOH = 3.35

pH = 10.65

119. A 0.20 M solution of a weak acid HA has a pH = 4.91. Use calcaulations and the table of "Relative Strengths of Bronsted-Lowry Acids and Bases" from the Data Booklet to determine the identity of the acid.

$$x = [H_3O^+] = 10^{-4.91} = 1.23\times10^{-5}$$

HA	$+H_2O$	$\rightleftharpoons A^-$	$+H_3O^+$
$.20$		0	0
$-x$		$+x$	$+x$
$.20-1.23\times10^{-5}$		1.23×10^{-5}	1.23×10^{-5}

$$K_a = \frac{[A^-][H_3O^+]}{[HA]}$$

$$K_a = \frac{(1.23\times10^{-5})^2}{(.20\cdot1.23\times10^{-5})}$$

$$= 7.568\times10^{-10}$$

$$= H_3BO_3$$

120. What concentration of NH_3 would be required to provide a solution with a pH of 11.35? basic!

$$pH = 11.35$$

$$pOH = 2.65 \quad [OH^-] = x = 2.2387\times10$$

NH_3	$+H_2O$	$\rightleftharpoons NH_4^+$	$+OH^-$
Q		0	0
$-x$		$+x$	$+x$
$Q-2.2387\times10^{-3}$		2.2387×10^{-3}	2.2387×10^{-3}

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

$$\frac{1.0\times10^{-14}}{5.6\times10^{-10}} = \frac{(2.2387\times10^{-3})^2}{(Q-2.2387\times10^{-3})}$$

$$Q-2.2387\times10^{-3} = \frac{(2.2387\times10^{-3})^2\times5.6\times10^{-10}}{1.0\times10^{-14}} = .28066485$$

$$Q = .28066485 + 2.2387\times10^{-3} = .28|29\ M$$

$$= .28\ M$$

121. Consider the following equilibrium:

$$2H_2O_{(l)} \overset{\longleftarrow}{\leftrightarrows} H_3O^{+}{}^{\uparrow}_{(aq)} + OH^{-}{}^{\downarrow}_{(aq)}$$

What changes occur to $[H_3O^+]$ and pH when $NaHSO_3$ is added?

Ka > Kb so it is acidic

	$[H_3O^+]$	pH
A.	~~decreases~~	decreases
[✓ B.]	increases	decreases
C.	~~decreases~~	~~increases~~
D.	increases	~~increases~~

122. In an aqueous solution of $Fe(NO_3)_3$, the pH is

acidic

[✓ A.] less than 7 and the solution is acidic
B. ~~greater~~ than 7 and the solution is acidic
C. ~~equal~~ to 7 and the solution is neutral
D. ~~greater~~ than 7 and the solution is basic

123. The $HCO_3^-{}_{(aq)}$ ion will act as

Kb > Ka

A. a ~~acid~~ since $K_a \cancel{>} K_b$
B. a base since $K_a \cancel{>} K_b$
[✓ C.] a base since $K_a < K_b$
D. a ~~acid~~ since $K_a < K_b$

124. Consider the following indicator equilibrium:

$$HIn + H_2O \leftrightarrows H_3O^+ + In^-$$

Which of the following is true about the transition point of this indicator?

A. moles of H_3O^+ = moles of In^-
B. pH = 7.0
[✓ C.] $[HIn] = [In^-]$
D. $[HIn] > [In^-]$

pH = 1.0

yellow ← 7.3

125. What color would 0.10 M HCl be in an indicator mixture consisting of phenol red and bromcresol green?

yellow ← 4.6

[✓ A] yellow
B. blue
C. green
D. purple

126. When the indicator thymol blue is added to 0.010 M solution of an unknown acid, the solution is orange. The acid could be

A. ~~H_2S~~

B. ~~HCN~~

[✓ C. HNO_3]

D. ~~HF~~

∴ pH = 2.0

∴ it is an acid

pH = 2.0 ∴ $[H_3O^+]$ = .010 M

∴ it is a strong acid

127. Consider the following equilibrium for the chemical indicator phenol red, HInd, at a pH = 7.3 (orange).

$$\underset{\text{yellow}}{HInd} + H_2O \overset{\longrightarrow}{\leftrightarrows} H_3O^+\downarrow + \underset{\text{red}}{Ind^-}$$

When some Na_2CO_3 (base) is added, what stress is imposed on the equilibrium and what colour change occurs?

	Stress	Indicator Colour Change
A.	decreased $[H_3O^+]$	turns ~~yellow~~
B.	~~increased~~ $[H_3O^+]$	turns red
[✓ C.]	decreased $[H_3O^+]$	turns red
D.	~~increased~~ $[H_3O^+]$	turns ~~yellow~~

128. The indicator phenolphthalein can be described by the following equilibrium equation:

$$\underset{\text{colorless}}{HIn} + H_2O \overset{\longrightarrow}{\leftrightarrows} H_3O^+\downarrow + \underset{\text{pink}}{In^-}$$

NaCN (base) is added to a slighly pink sample of this indicator. After equilibrium has been re-established, how do the $[H_3O^+]$ and the colour of the solution compare with the original equilibrium?

	$[H_3O^+]$	Colour of solution
A.	~~increases~~	turns ~~colourless~~
B.	decreases	turns ~~colourless~~
[✓ C.]	decreases	turns more pink
D.	~~increases~~	turns more pink

129. A chemical indicator has a $K_a = 1.6 \times 10^{-4}$. What is the pH at the transition point and what is the identity of the indicator?

	pH	Indicator
A.	10.2	thymophthalien
B.	10.2	phenophthalien
C.	3.8	bromcresol green
[✓ D.]	3.8	methyl orange

$pH = -\log(1.6 \times 10^{-4})$

sig figs!

130. At 45.0 °C, $K_w = 4.00 \times 10^{-14}$ for pure water.

a. Calculate the pH of water at 45.0 °C.

$K_w = [H_3O^+][OH^-] = [H_3O^+]^2 = 4.00 \times 10^{-14}$

$[H_3O^+] = 2.00 \times 10^{-7}$

pH = 6.70

yellow 8.8 ←

6.0 red →

b. A mixture of the indicators Thymol Blue and Chlorophenol Red is added to the water. What is the resulting colour of the mixture? Explain.

Resulting color orange

Explaination: thymol blue is yellow

Chlorophenol Red is red

131. A 20.0 mL sample of H_2SO_4 is titrated with 25.0 mL of 0.20 M $Sr(OH)_2$. What is the concentration of the acid?

$H_2SO_4 + Sr(OH)_2 \rightarrow 2H_2O + SrSO_4$

$25.0 \times 10^{-3} L \times \frac{.20 \text{ moles of } Sr(OH)_2}{L} \times \frac{1 \text{ mole } H_2SO_4}{1 \text{ mole } Sr(OH)_2} \times \frac{1}{20.0 \times 10^{-3} L}$

A. 0.50 M
[✓ B. 0.25 M]
C. 0.20 M
D. 0.13 M

132. At the equivilence point, the titration of HCl with $Ba(OH)_2$ will form a solution which is

A. acidic with pH < 7
[✓ B. neutral with pH = 7]
C. basic with pH < 7
D. basic with pH >7

133. What is always true about the pH at the equivilence point when a weak acid is titrated with a strong base?

A. the pH does not change anymore, even if more strong base is added
B. pH = 7.0
[✓ C. pH > 7.0]
D. pH < 7.0

134. Which of the following pairs of chemicals could be used to make a buffer solution?

A. $NaClO_4$ and $HClO_4$
B. HNO_3 and $NaNO_3$
[✓ C.] HCN and NaCN
D. HI and NaI

135. Consider the following buffer equilibrium:

$$HF_{(aq)} + H_2O_{(l)} \leftrightarrows H_3O^+_{(aq)} + F^-_{(aq)}$$

What would limit the buffering action if KCH_3COO were added?

A. [HF]
B. $[H_2O]$
C. $[H_3O^+]$
[✓ D.] $[F^-]$