Recitation Worksheet 11: Acid-Base Equilibria (15.1-15.4 and 15.6)

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→ A buffer solution is a solution of a weak acid & salt of its conjugate base or a weak base & the salt of its conjugate acid.

1. Which of the aqueous solutions below are buffer solutions? Select all that apply. Insert letters without spaces in the answer box, example **ABCD.**

CEF

A. 0.100 M KBr
B. 0.200 M NaCl and 0.200 M NH_4Cl
C. 0.100 M CH_3NH_2 and 0.150 M $CH_3NH_3^+Cl^-$
D. 0.100 M HCl and 0.050 M $NaNO_2$
E. 0.100 HCl and 0.200 M $NaCH_3COO$
F. 0.100 M Na_2HPO_4 and 0.100 M NaH_2PO_4
G. 0.100 M CH_3COOH and 0.100 M $NaCH_3CH_2COO$

(A) KBr is a pH-neutral salt.

(B) NaCl is a pH-neutral salt & $NH_4^+Cl^-$ is an acidic salt

(C) CH_3NH_2 is a weak base & $CH_3NH_3^+Cl^-$ the salt of its conjugate acid

(D) $HCl + NaNO_2 \rightarrow HNO_2 + NaCl$

	HCl	$NaNO_2$	HNO_2	$NaCl$
I	0.100 mol	0.050 mol	0	0
C	− 0.050 mol	0.050 mol	+ 0.050	+ 0.050
F	0.050 mol	0 mol	0.050 mol	0.050 mol

From the ICF table HCl (strong acid), HNO_2 (weak acid), & NaCl (pH-neutral salt) which non of these components from a buffer

(E) $HCl + CH_3COONa \rightarrow CH_3COOH + NaCl$

(CH_3COONa: conjugate salt of CH_3COOH; CH_3COOH: weak acid)

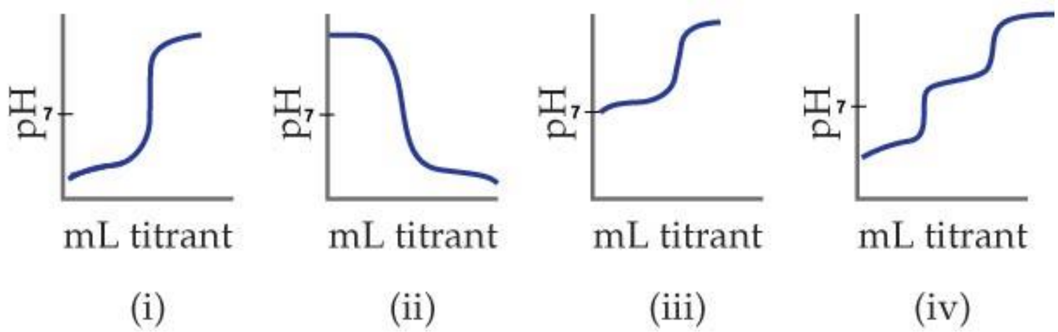
	HCl	CH_3COONa	CH_3COOH	$NaCl$
I	0.100 mol	0.200 mol	0	0
C	− 0.100	− 0.100 mol	+ 0.100 mol	+ 0.100 mol
F	0	0.100 mol	0.100 mol	0.100 mol

Buffer mixture

(F) 0.100 M Na_2HPO_4 and 0.100 M NaH_2PO_4 ($H_2PO_4^-$ is a weak acid & HPO_4^{2-} is it's conjugate base)

(G) Although CH_3COOH is a weak acid CH_3CH_2COONa is not the salt of it's conjugate base

2. Match the titrations curves (i-iv) to the appropriate description (A-D).



ii A. Titration of a strong base with a strong acid.

iii B. Titration of a weak acid with a strong base.

i C. Titration of a strong acid with a strong base.

iv D. Titration of a polyprotic acid with a strong base.

K_a of $NH_4^+ = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$
$\therefore pK_a = -\log(5.6 \times 10^{-10}) = 9.26$

3. In the laboratory, you were asked to prepare a buffer solution with pH = 10.50. How many **grams** of NH_4Cl (molar mass = 53.5 g/mol) would you add to 625 mL of 0.258 M NH_3 to prepare a buffer with pH = 10.50? Assume that the solution's volume remains constant. K_b of $NH_3 = 1.8 \times 10^{-5}$. Keep your answer to three significant figures.

0.491	g

① Using Henderson Hassel-balch equation to determine the ratio of base to acid

$pH = pK_a + \log\frac{[NH_3]}{[NH_4Cl]} \Rightarrow 10.50 = 9.26 + \log\frac{[0.258]}{[NH_4Cl]}$

$\therefore 1.25 = \log\frac{[0.258]}{[NH_4^+]}$ → taking the inverse of log on both sides

$17.56820923 = \frac{0.258}{[NH_4^+]} \Rightarrow \therefore [NH_4^+] = 0.014685617\ M$

② Use the molarity of NH_4Cl to determine the grams of NH_4Cl added to 625mL of 0.258 M NH_3

$625\ mL \times \frac{0.014685617\ mol\ NH_4^+}{1000\ mL} \times \frac{1\ mol\ NH_4Cl}{1\ mol\ NH_4^+} \times \frac{53.5\ g\ NH_4Cl}{1\ mol\ NH_4Cl}$

$= 0.491\ g$

4. If 0.100 M of the solutions below were provided to you, which **two solutions** would you use to prepare a buffer with pH = 3.50? Select all that apply. Insert letters without spaces in the answer box, example **AB.**

AE

A. Formic acid (HCOOH, $pK_a = 3.74$)
B. Acetic acid, (CH_3COOH, $pK_a = 4.74$)
C. Phosphoric acid (H_3PO4, $pK_{a1} = 2.15$)
D. Sodium acetate ($NaCH_3COOH$)
E. Sodium formate (NaHCOO)
F. Sodium dihydrogen phosphate (NaH_2PO_4)

- to prepare a buffer with pH = 3.50 you need to choose an acid with a pka close to the pH of the buffer & in this case this is formic acid
- the conjugate base salt of formic acid is sodium formate

5. Calculate the final **pH** in each of the titration scenarios below:
   A. The titration of 25.00 mL of 0.160 M HCl with 15.00 mL of 0.242 M NaOH. Keep your answers to two decimal places.

2.03

(strong acid) (strong base)

① $HCl + NaOH \longrightarrow NaCl + H_2O$

	HCl	NaOH	NaCl	H_2O
I	0.00400	0.00363	0	—
C	−0.00363	−0.00363	+0.00363	+0.00363
F	0.00037	0	0.00363	0.00363

0.00037 → left over HCl

0.00363 → pH-neutral salt ∴ Na^+ & Cl^- ions are spectator ions

② Calculate the new molarity of HCl.

$\frac{0.00037\ mol}{25.00 + 15.00\ mL} \times \frac{1000\ mL}{1\ L} = 0.00925\ M$

③ pH of solution will be the pH of the remaining HCl
$\therefore pH = -\log[0.00925] = 2.03$

B. The titration of 25.00 mL of 0.100 M CH_3COOH (K_a of CH_3COOH = 1.7 x 10^{-5}) with 12.5 mL of 0.200 M NaOH. Keep your answers to two decimal places.

8.80

weak acid — strong base

(1)

$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$

	CH_3COOH	NaOH	CH_3COONa	H_2O
I	0.00250 mol	0.00250	0	–
C	–0.00250 mol	–0.00250	+0.00250 mol	–
F	0	0	0.00250 mol	–

salt with basic pH

(2) Molarity of CH_3COONa

$$\frac{0.00250\ mol}{25.00 + 12.5\ mL} \times \frac{1000\ mL}{1L} = 0.0667M$$

(3) Calculate the pH of $CH_3COO^-Na^+$ (Na^+ is pH-neutral)

$CH_3COO^-(aq) + H_2O(l) \rightleftharpoons CH_3COOH + OH^-$

	CH_3COO^-	H_2O	CH_3COOH	OH^-
I	0.0667	–	0	0
C	– x	–	+x	+x
E	0.0667–x	–	x	x

solving for a base ∴ determine K_b ∴ $K_b = \frac{1.0 \times 10^{-14}}{1.7 \times 10^{-5}} = 5.9 \times 10^{-10}$ ∴ $5.9 \times 10^{-10} = \frac{[x][x]}{[0.0667 - x]}$

$\frac{C}{K} \gg 100$

$x^2 = 6.9 \times 10^{-10} \times 0.0677$ — take square root of both sides

$x = \pm 6.2638 \times 10^{-6}$

(4) $pOH = -\log[6.2638 \times 10^{-6}]$
$= 5.20$

∴ pH + pOH = 14.00
∴ pH = 14.00 – 5.20
= 8.80

6. What is the pH of a mixture of 0.012 M of C_6H_5COOH (K_a = 6.3 x 10^{-5}) and 0.033 M NaC_6H_5COOH? Keep your answers to two decimal places.

4.64

$$pH = pKa + \log\frac{[NaC_6H_5COOH]}{[C_6H_5COOH]}$$

$$= -\log(6.3 \times 10^{-5}] + \log\frac{[0.033]}{[0.012]}$$

$= 4.63999$

~ 4.64

Buffer solution: HClO, NaClO

7. A solution is prepared by dissolving 0.23 mol of hypochlorous acid and 0.27 mol of sodium hypochlorite in water sufficient to yield 1.00 L of solution. The addition of 0.05 mol of HCl to this buffer solution causes the pH to drop slightly. The pH does not decrease drastically because the HCl reacts with the _________ present in the buffer solution. The K_a of hypochlorous acid is 1.36 × 10^{-3}.

C

A. H_2O
B. H_3O^+
C. Hypochlorite ion
D. Hypochlorous acid
E. This is a buffer solution. The pH does not change upon addition of acid or base.

limiting reactant

$HCl + NaClO \rightarrow HClO + NaCl$

	HCl	NaClO	HClO	NaCl
I	0.05	0.27	0.23	0
C	–0.05	–0.05	+0.05	+0.05
F	0	0.22	0.28	0.05

Moles of CH_3COOH = 0.300 L × $\frac{0.250 \text{ mol}}{1 \text{ L}}$ = 0.075 mol Buffer Moles of $NaCH_3COO$ = 0.300 L × $\frac{0.560 \text{ mol}}{1 \text{ L}}$ = 0.168 mol

8. To a 0.300 L buffer solution consisting of 0.250 M CH_3COOH and 0.560 M $NaCH_3COO$, 0.0060 mol HCl is added. What are the moles and concentration of CH_3COOH after the addition of HCl? Assume that the volume of the buffer does not change upon the addition of HCl.

strong acid → should react with base component in the buffer (CH_3COO^-)

C

A. 0.0060 mol, 0.020 M
B. 0.162 mol, 0.54 M
C. 0.081 mol, 0.27 M
D. 0.075 mol, 0.250 M
E. 0.168 mol, 0.560 M

① $HCl + CH_3COO^- \rightarrow CH_3COOH + Cl^-$

	HCl	CH_3COO^-	CH_3COOH	Cl^-
I	0.0060	0.168 mol	0.075 mol	0
C	−0.0060	−0.0060	+0.0060	+0.0060
F	0	0.162	0.081	0.0060

② Molarity of CH_3COOH = $\frac{0.081 \text{ mol}}{0.300 \text{ L}}$ = 0.27 M

9. In which of the solutions given below would nitrous acid (HNO_2) ionize *less* than it does in pure water?

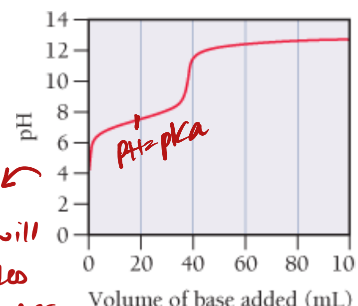
C

A. NaCl
B. KNO_3
C. KNO_2
D. $NaClO_4$

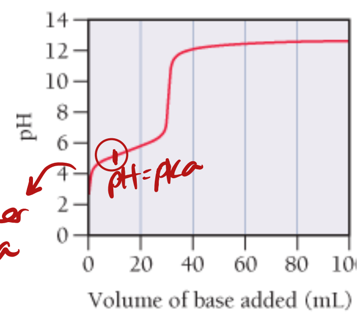
↓ common-ion effect

$HNO_2 + H_2O \rightleftharpoons NO_2^- + H_3O^+$
$KNO_2(aq) \rightarrow NO_2^- + K^+$
↳ common-ion ∴ shifts the equilibrium position of HNO_2 to the left ∴ HNO_2 ionizes less

10. You are provided with the titration curves I and II for two weak acids titrated with 0.100 M NaOH. Refer to the titration curves to answer the following questions:



I



II

For the more conc. acid it will take more moles of NaOH & greater volume to reach the equivalence point

lower pKa

A. Which acid is more concentrated? **I**

B. Which acid has the larger K_a? **II**

pH = pKa at half equivalence point; the ↓ the pKa the larger the Ka & the stronger the acid

11. Which solution has the **greatest** buffering capacity?

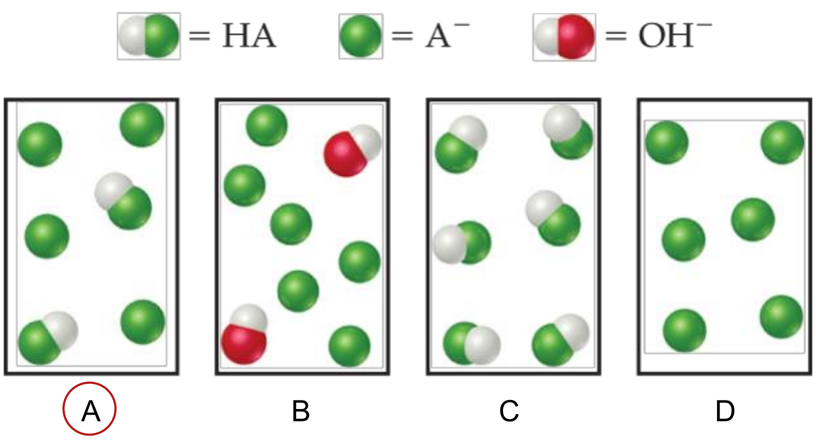
C

A. 0.335M $HC_2H_3O_2$ and 0.497 M $NaC_2H_3O_2$
B. 0.520 M $HC_2H_3O_2$ and 0.116 M $NaC_2H_3O_2$
C. 0.820 M $HC_2H_3O_2$ and 0.715 M $NaC_2H_3O_2$
D. 0.120 M $HC_2H_3O_2$ and 0.115 M $NaC_2H_3O_2$

the greatest buffer capacity is when the buffer components are large & equal

12. A strong base such as KOH is mixed in a specific proportion with the weak acid HA to make a buffer. Which of the diagrams below is a correct representation of the buffer solution?

A



KOH + HA → KA + H_2O

the weak acid & it's conjugate base salt should be present in order for the solution to be a buffer solution

13. What change will be caused by addition of a small amount of HCl to a solution containing fluoride ions and hydrogen fluoride?

strong acid

HF/F^-

D

A. The concentration of hydronium ions will increase significantly.
B. The concentration of fluoride ions will increase as will the concentration of hydronium ions. Conc. of F^- ↓ & H_3O^+ ↑
C. The concentration of hydrogen fluoride will decrease, and the concentration of fluoride ions will increase. HF↑ & F^-↓
D. The concentration of fluoride ion will decrease, and the concentration of hydrogen fluoride will increase.
E. The fluoride ions will precipitate out of solution as its acid salt.

HCl + F^- → HF + Cl^-

more HF is produced ∴ the pH of the buffer will increase

14. A 1.00 L buffer solution is 0.150 M in $HC_7H_5O_2$ and 0.250 M in $LiC_7H_5O_2$. Calculate the pH of the solution after the addition of 100.0 mL of 1.00 M HCl. The Ka for $HC_7H_5O_2$ is 6.5×10^{-5}.

3.97

A. 4.19
B. 5.03
C. 4.41
D. 3.34
(E.) 3.97

① $HCl + C_7H_5O_2^- \rightarrow C_7H_5O_2H + Cl^-$

	HCl	$C_7H_5O_2^-$	$C_7H_5O_2H$	Cl^-
I	0.100 mol	0.250 mol	0.150 mol	0
C	−0.100 mol	−0.100 mol	+0.100 mol	+0.100 mol
F	0	0.150	0.250 mol	

② Using Henderson-Hasselbalch equation to calculate the pH of the buffer

$$pH = pKa + \log \frac{[C_7H_5O_2^-]}{[HC_7H_5O_2]}$$

$$= -\log(6.5\times10^{-5}) + \log \frac{[0.150]}{[0.250]}$$

$$= 3.96523$$

$$\approx 3.97$$

15. You are working in the lab with three acidic solutions. **Solution 1** is 0.1 M of a **weak monoprotic acid**, **solution 2** is 0.1 M of a **strong monoprotic acid** and **solution 3** is 0.1 M of a **weak diprotic acid**. Each of the former solutions has been titrated with a 0.2 M KOH solution. Which *quantity* is the same for all the three solutions?

B

A. The volume required to reach the final equivalence point.
(B.) The volume required to reach the first equivalence point.
C. The pH at the first equivalence point.
D. The pH at one-half the first equivalence point.
E. None of the quantities is the same for the three solutions.

At equivalence point, the mol of the acid = mol of the base. Since all the acids have the same conc ($\frac{0.1 mol}{1L}$) & they are titrated with 0.2M KOH
∴ the volume to reach the first equivalence point

16. Which of the following statements is **<u>accurately</u>** describes the common-ion effect? → decreased

C

~~A.~~ The extent of ionization of a weak electrolyte is increased by adding to the solution a strong electrolyte that has an ion in common with the weak electrolyte.
~~B.~~ The solubility of a slightly soluble salt is increased by the presence of a second solute that provides a common ion to the system. → decreased
(C.) The common ion effect occurs when a solubility equilibrium is shifted due to a second compound that contains an ion common with the first.
~~D.~~ The common ion effect is that common ions precipitate all counter-ions.
~~E.~~ None of the above statements accurately describe the common-ion effect.

Choice C is the definition of the common-ion effect