

Recitation Worksheet 12 – Exam 4 Review

Name:

key

MyID:

Instructions:

1. This recitation worksheet is for the exam 4 review.
2. You **do not need** to submit it to GradeScope.
3. The answer key has been posted with this worksheet to eLC.
4. The **recitation session in the exam week (Apr 17-20) is still mandatory**. The attendance will be recorded.

1. What structural features of a molecule may affect the pKa of an acid?

E

- A. Electronegativity
 B. The strength of the bond to the acidic hydrogen
 C. Inductive effect
 D. Resonance destabilization
 E. All the above

2. Which of the following acidity relationships is true?

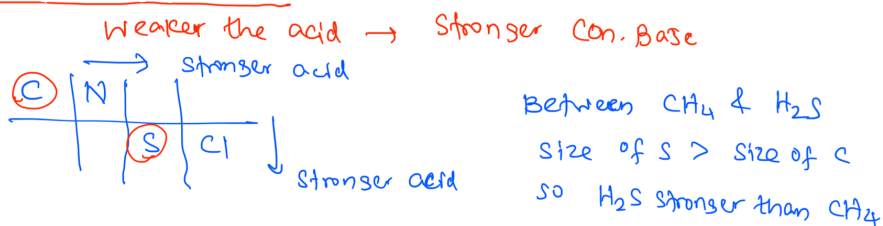
B

- A. $\text{H}_2\text{SO}_3 > \text{H}_2\text{SO}_4$ → $\text{H}_2\text{SO}_3 < \text{H}_2\text{SO}_4$ (# of Oxygens)
 B. $\text{H}_2\text{PO}_4^- > \text{HPO}_4^{2-}$ → Higher the negative charge, Hard to remove H^+
 C. $\text{HF} > \text{HClO}_4$ X
 D. $\text{H}_2\text{CO}_3 > \text{HNO}_3$ X → $\text{H}_2\text{CO}_3 < \text{HNO}_3$ (N is more electronegative than C)
 E. None of these

3. Which acid of the set has the strongest conjugate base?

A

- A. CH_4
 B. NH_3
 C. H_2S
 D. HCl



4. Which of the following acids will be the strongest?

A

- A) H_2SO_4 B) HSO_4^- C) H_2SO_3 D) H_2SeO_4 E) HSO_3^-

A, C, D → neutral molecules
 B & E → negatively charged

} neutral molecules are stronger acids than negatively charged molecules

- Between A, C, and D, choices A & D have a greater no. of oxygens than choice C
 - Between choices A & D sulfur is more electronegative than selenium (oxyacids) ∴ A is the strongest acid

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5. For which of the pairs of acids is the stronger acid listed first? (Select all that apply).

ACD

- A. HI or HCl $I > Cl$ (size)
 B. HF or HBr $Br > F$
 C. HClO₃ or HBrO₃ Cl more electronegative than Br
 D. HOSO₂CF₃ or HOSO₂CH₃
 Highly electronegative than H

6. Which of the following acids will be the strongest?

II

- I. CH₃OH II. CH₃SH III. CH₃PH₂

| | | | |
|--|---|---|--|
| | | O | |
| | P | S | |

The three acid molecules have a similar structure but the ionizable hydrogen is attached to different atoms
 - In this case we will arrange the acids in terms of strength similar to oxy acids (size is more important). In this case, $S > O$ in size & $S > P$ in electronegativity
 \therefore CH₃SH is the strongest acid

7. Predict which one is the stronger acid of each of the following pairs of acids?

H₃PO₄

- A. H₂SiO₃ or H₃PO₄ \rightarrow H₃PO₄ has more Oxygens. P is more EN $>$ Si

H₂CO₃

- B. H₂CO₃ or H₂BO₃ \rightarrow C more electronegative than B. C has higher Oxidation number than B.

HOClO

- C. HOClO or HOBr
 \rightarrow HOClO \rightarrow has more Oxygens
 \rightarrow Cl more electronegative than Br

8. Arrange the following binary compounds in order of increasing acid strength.

III < I < II

- I. H₂Te II. HI III. H₂S
 Strongest weakest

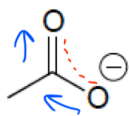
H₂Te & HI
 I more electronegative than Te
 so HI is stronger than H₂Te

H₂Te and H₂S
 Size of Te $>$ Size of S
 so H₂Te stronger than H₂S

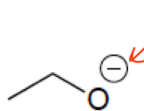
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9. Rank these molecules in terms of **decreasing basicity (strongest to weakest base)**:

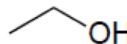
E



I. electron density around negative charge become lower because other 'O' atom withdraws electrons towards itself



II. Strongest Base



III.

A. I > II > III

B. II > III > I

C. III > II > I

D. I > III > II

E. II > I > III

Basic strength (satisfy H^+)

* Cation < Neutral < anion
↑ weakest ← Strongest

10. Which of these species is probably **the weakest acid**?

D

- ~~A.~~ HCl
- ~~B.~~ H_3PO_4
- ~~C.~~ $H_2PO_4^-$
- D. HPO_4^{2-}
- ~~E.~~ HNO_3

Acid strength (donate H^+)

* anion < Neutral < Cation

↑ weakest

↑ Strongest

Higher the negative charge, more like to protons, so harder to remove H^+
 HPO_4^{2-} weaker than $H_2PO_4^-$

11. Given the following acids and K_a values:

E

$HClO_4$

1×10^7

Strongest

$HOAc$

1.76×10^{-5}

HCN

4.93×10^{-10}

HF

3.53×10^{-4}

What is the order of increasing base strength?

A. CN^- , F^- , OAc^- , ClO_4^-

B. CN^- , OAc^- , F^- , ClO_4^-

C. CN^- , ClO_4^- , F^- , OAc^-

D. ClO_4^- , OAc^- , CN^- , F^-

E. ClO_4^- , F^- , OAc^- , CN^-

stronger acid \rightarrow weaker conjugate base

$HClO_4$ - strongest acid \rightarrow ClO_4^- weakest Con. Base

HCN - weakest acid \rightarrow CN^- strongest Con. Base

$ClO_4^- < F^- < OAc^- < CN^-$

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Buffer

salt of conjugate acid

weak base

strong acid

12. A solution is prepared by dissolving 0.32 mol of $\text{CH}_3\text{CH}_2\text{NH}_3\text{Cl}$ in 1.00 L of 1.5 M $\text{CH}_3\text{CH}_2\text{NH}_2$. If 10. mL of 0.11 M HCl is added to this solution, the pH of the solution will slightly _____ because the HCl reacts with the _____ present in the solution.

C

- A. Increase, $\text{CH}_3\text{CH}_2\text{NH}_3^+$
B. Increase, $\text{CH}_3\text{CH}_2\text{NH}_2$
C. Decrease, $\text{CH}_3\text{CH}_2\text{NH}_2$
D. Decrease, $\text{CH}_3\text{CH}_2\text{NH}_3^+$

strong acid is added to buffer. strong acid react with the weak base, and slightly decrease the pH of solution

13.

Identify the false statement regarding a solution that contains 0.20 moles of hypochlorous acid. (weak acid)

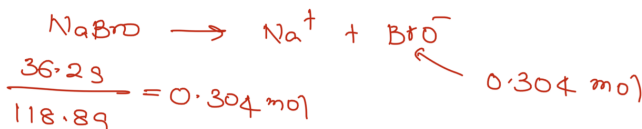
C

- A. Adding NaOH will increase the dissociation of HOCl, and decrease $[\text{H}^+]$
 $\text{NaOH} + \text{H}_3\text{O}^+ \rightarrow \text{Na}^+ + 2\text{H}_2\text{O}$, Decrease $[\text{H}_3\text{O}^+]$, shift \rightarrow
B. Adding HCl will decrease the dissociation of HOCl and decrease $[\text{OCl}^-]$
 $\text{HCl} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{Cl}^-$, Increase $[\text{H}_3\text{O}^+]$, shift \leftarrow
C. Adding NaOCl will increase the dissociation of the original HOCl, and will increase $[\text{H}^+]$
 $\text{NaOCl} \rightarrow \text{Na}^+ + \text{OCl}^-$, Increase $[\text{OCl}^-]$, shift \leftarrow , decrease dissociation, and decrease $[\text{H}^+]$
D. Adding NaCl will not affect either the dissociation of the original HOCl or the solution pH
 $\text{NaCl} \rightarrow \text{Na}^+ + \text{Cl}^-$, Increase $[\text{Na}^+]$ or $[\text{Cl}^-]$ has no effect on equilibrium of HOCl.

14. What volume of 0.80 M HCl will be required to titrate 36.2 grams of NaBrO to the equivalence point?

C

- A. 150 mL
B. 308 mL
C. 381 mL
D. 258 mL
E. None of the above



$$m_1 x_1 = 0.304 \text{ mol}$$

$$V_1 = \frac{0.304 \text{ mol}}{0.80 \text{ M}} = 381 \text{ mL}$$

15. To 60.0 mL of a solution that contains 0.80 M NaF and 0.80 M HF was added 20.0 mL of 0.40 M HCl.

A

- A. 0.056 moles, 0.70 M
B. 0.056 moles, 0.93 M
C. 0.048 moles, 0.60 M
D. 0.040 moles, 0.67 M
E. none are correct.

Calculate the moles of HF and the concentration of HF after addition. Strong acid HCl reacts with weak base NaF

Limiting reactant.

| | HCl | NaF | HF | NaCl |
|---------|---------|---------|---------|---------|
| Initial | 0.0080 | 0.048 | 0.048 | 0 |
| Change | -0.0080 | -0.0080 | +0.0080 | +0.0080 |
| Final | 0 | 0.040 | 0.056 | 0.0080 |

moles of NaF = $0.80 \times 60.0 \times \frac{1}{1000} = 0.048 \text{ mole}$
moles of HF = $0.80 \times 60.0 \times \frac{1}{1000} = 0.048 \text{ mole}$
moles of HCl = $0.40 \times 20.0 \times \frac{1}{1000} = 0.0080 \text{ mole}$

conc of HF after reaction = $\frac{0.056 \text{ mole}}{(60+20) \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}}} = 0.70 \text{ M}$

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16. TRIS $\{(\text{HOCH}_2)_3\text{CNH}_2\}$ is one of the most common buffers used in biochemistry. A solution is prepared by adding enough TRIS and 12 M $\text{HCl}(\text{aq})$ to give 1.00 L of solution with $[\text{TRIS}] = 0.30 \text{ M}$ and $[\text{TRISH}^+] = 0.60 \text{ M}$. What is the pH of this buffered system if the pK_b is 5.92?

C

- A. 5.92
B. 6.22
C. 7.78
D. 8.08

pH of buffer:

$$\text{pH} = \text{pK}_a + \log \frac{[\text{Base}]}{[\text{Acid}]} = 8.08 + \log \frac{0.30}{0.60} = 7.78$$

$\text{pK}_a = 14 - \text{pK}_b = 14.00 - 5.92 = 8.08$

base *acid*

17. All of the following solutions would be considered buffers **except**

B

A. $\text{CH}_3\text{COOH} / \text{CH}_3\text{COO}^-$.

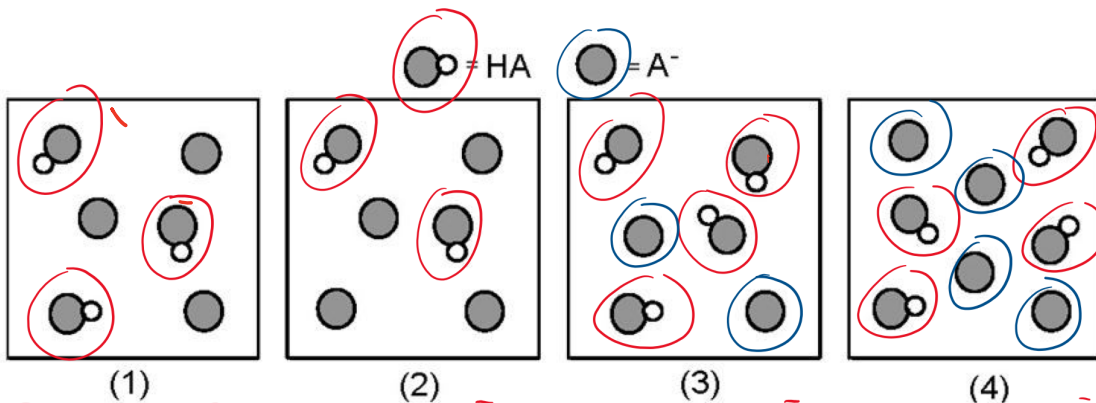
B. NaCl / HCl . → strong acid

C. $\text{H}_2\text{PO}_4^- / \text{HPO}_4^{2-}$.

D. $\text{HPO}_4^{2-} / \text{PO}_4^{3-}$.

buffer contains weak conjugated
acid / base pair.

18. These pictures represent solutions that contain a weak acid HA ($\text{pK}_a = 5.0$) and its sodium salt NaA. Unshaded spheres represent H atoms and shaded spheres represent A^- ions. (Na^+ , H_3O^+ , OH^- , and solvent H_2O molecules have been omitted for clarity.)



Which solution has the greatest buffer capacity?

D

- A. (1)
B. (2)
C. (3)
D. (4)

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20. Which titration curve corresponds to an initial pH of 10.7 and an equivalence point at pH = 4.5?

D

- A. A strong acid to which strong base is added
- B. A strong base to which strong acid is added
- C. A weak acid to which strong acid is added
- D. A weak base to which strong acid is added**
- E. A weak base to which strong base is added

starting with a basic pH → acidic at equivalence point. pH < 7, so the ion in salt is acidic (weak acid). So, before titration, it was the conjugated weak base

21. What is the percent dissociation of glycine if the solution has a pH = 8.60 and a pKa = 9.60?

B

- A. 50%
- B. 9%**
- C. 5%
- D. 1%

Assume concentration of HA is (M) before ionization.
 $[H_3O^+] = 10^{-pH} = 10^{-8.60} = 2.51 \times 10^{-9} M$
 $K_a = 10^{-pK_a} = 10^{-9.60} = 2.51 \times 10^{-10}$
 $HA + H_2O \rightleftharpoons H_3O^+ + A^-$

| | | | |
|---|---------------------------|------------------------|------------------------|
| I | (M) | | |
| C | -2.51×10^{-9} | $+2.51 \times 10^{-9}$ | $+2.51 \times 10^{-9}$ |
| E | $M - 2.51 \times 10^{-9}$ | 2.51×10^{-9} | 2.51×10^{-9} |

$$K_a = \frac{(2.51 \times 10^{-9})^2}{M - 2.51 \times 10^{-9}} = 2.51 \times 10^{-10} \Rightarrow M = 2.76 \times 10^{-8}$$

$$\% \text{ ionization} = \frac{2.51 \times 10^{-9}}{2.76 \times 10^{-8}} \times 100\% = 9\%$$

22. Twenty-five milliliters of 0.10 M HCl(aq) is titrated with 0.10 M NaOH(aq). What is the pH after 15 mL of NaOH(aq) has been added?

E

- A. 1.4
- B. 1.2
- C. 1.0
- D. 2.0
- E. 1.6**

$\left\{ \begin{array}{l} \text{mole HCl} = 25 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times 0.10 \text{ M} = 2.5 \times 10^{-3} \text{ mole} \\ \text{mole NaOH} = 15 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times 0.10 \text{ M} = 1.5 \times 10^{-3} \text{ mole} \end{array} \right.$
 limiting reactant
 $HCl + NaOH \rightarrow NaCl + H_2O$

| | | | |
|---|-----------------------|-----------------------|-----------------------|
| I | 2.5×10^{-3} | 1.5×10^{-3} | 0 |
| C | -1.5×10^{-3} | -1.5×10^{-3} | $+1.5 \times 10^{-3}$ |
| F | 1.0×10^{-3} | 0 | 1.5×10^{-3} |

 After reaction, there are 1.0×10^{-3} mole HCl and 1.5×10^{-3} mole NaCl in the solution. NaCl is neutral, pH = 7.
 $\text{Conc of HCl} = \frac{1.0 \times 10^{-3} \text{ mole}}{(25 + 15) \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}}} = 0.025 \text{ M} = [H_3O^+]$
 $pH = -\log [H_3O^+] = -\log (0.025 \text{ M}) = 1.6$

23. In a titration experiment, it was determined that a 50.0 mL sample of HNO₃ required 66.0 mL of 0.80 M NaOH to reach the equivalence point. What was the molarity of the HNO₃?

C

- A. 0.61 M
- B. 0.86 M
- C. 1.06 M**
- D. 1.24 M
- E. none of these are correct

$HNO_3 + NaOH \rightarrow NaNO_3 + H_2O$
 mole of HNO₃ = mole of NaOH
 $50.0 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times M_{HNO_3} = 66.0 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times 0.80 \text{ M}$
 $M_{HNO_3} = 1.06 \text{ M}$

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24. HN_3 has $K_a = 2.0 \times 10^{-5}$. What is the concentration of N_3^- in a solution that is 0.50 M in HN_3 and 0.40 M in HNO_3 ?

B

A. 3.5×10^{-3}

B. 2.5×10^{-5}

C. 2.0×10^{-6}

D. 8.0×10^{-6}

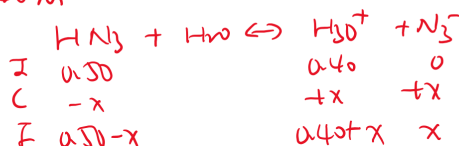
E. None of these are correct

$$\left\{ \begin{array}{l} \frac{0.40}{K_a} > 100 \text{ ignore } +x \\ \frac{0.50}{K_a} > 100 \text{ ignore } -x \end{array} \right.$$



$[\text{H}_3\text{O}^+]$ from $\text{HNO}_3 = 0.40 \text{ M}$

then HN_3 ionize in H_2O :



$$K_a = \frac{(0.40+x)x}{0.50-x} = 2.0 \times 10^{-5}$$

$$\frac{0.40 \times x}{0.50} = 2.0 \times 10^{-5}$$

$$x = 2.5 \times 10^{-5}$$

25. Which combination will give a pH of lower than 7.00 at the equivalence point?

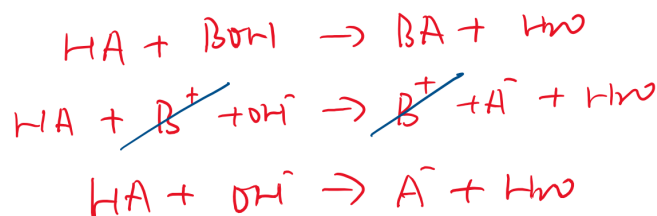
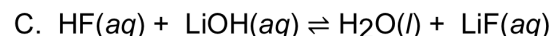
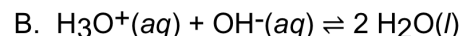
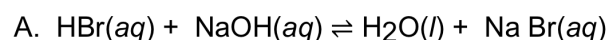
A



E. None of these are correct

26. Which is a net ionic equation for the neutralization of a weak acid with a strong base?

D



27. What is the hydronium ion concentration in a solution prepared by mixing 50.00 mL of 0.10 M HCN with 50.00 mL of 0.010 M NaCN ? Assume that the volumes of the solutions are additive and that $K_a = 4.9 \times 10^{-10}$ for HCN .

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C

- A. 4.9×10^{-11} M
- B. 4.9×10^{-10} M
- C. 4.9×10^{-9} M
- D. 7.0×10^{-6} M

buffer.

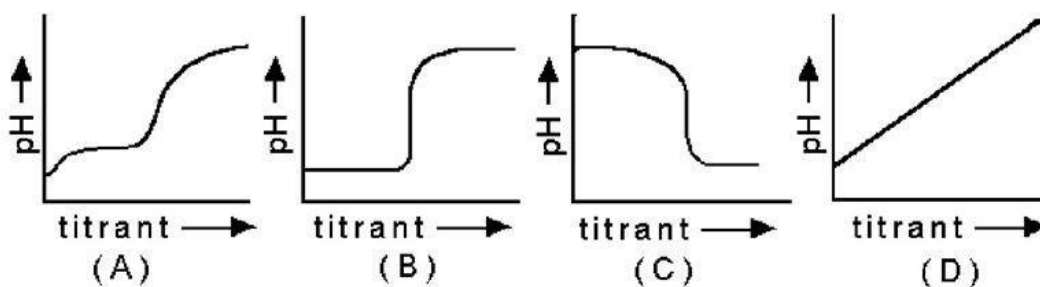
$$\text{pH} = \text{pK}_a + \log \frac{[\text{Base}]}{[\text{Acid}]} = -\log(K_a) + \log \frac{[\text{Base}]}{[\text{Acid}]}$$

$$= -\log(4.9 \times 10^{-10}) + \log \frac{50.0 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times 0.010 \text{ M} \times \frac{1}{(50.0 + 50.0) \text{ mL}}}{50.0 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times 0.10 \text{ M} \times \frac{1}{(50.0 + 50.0) \text{ mL}}}$$

$$= 8.31$$

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-8.31} = 4.9 \times 10^{-9}$$

28.



I. What is the characteristic pH-titrant curve for the titration of a strong acid by a strong base?

B

II. What is the characteristic pH-titrant curve for the titration of a strong base by a strong acid?

C

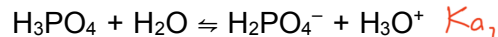
III. What is the characteristic pH-titration curve for the titration of a weak acid by a strong base?

A

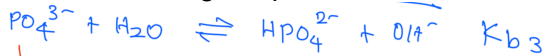
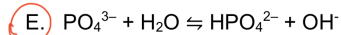
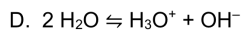
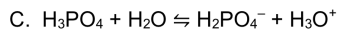
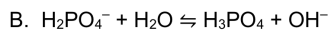
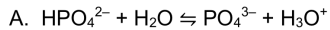
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29. Phosphoric acid is a triprotic acid, ionizing in sequential steps:

E

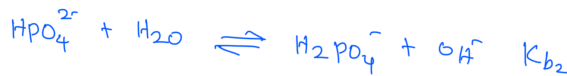


Which equilibrium is most important in determining the pH of a solution of sodium phosphate?



↓

$$K_b(\text{PO}_4^{3-}) = \frac{K_w}{K_a(\text{HPO}_4^{2-})}$$



$$K_{a1} \gg K_{a2} \gg K_{a3}$$

$$\text{so } K_{b1} \ll K_{b2} \ll K_{b3}$$

30. What is the pH of a solution prepared by mixing 50.00 mL of 0.10 M NH_3 with 20.00 mL of 0.10 M NH_4Cl ?

$K_b = 1.8 \times 10^{-5}$ for NH_3 .

9.65

Buffer.

$$\text{pH} = \text{pK}_a + \log \left(\frac{[\text{Base}]}{[\text{Acid}]} \right) = 9.25 + \log \left(\frac{50.00 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times 0.10 \text{ M} \times \frac{1}{51.20}}{20.00 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times 0.10 \text{ M} \times \frac{1}{51.20}} \right)$$

$$= 9.25 + 0.398 = 9.65$$

31. Solutions of sodium salts of the acids in the table are prepared with an initial concentration of 0.500 M.

Which solution will have the highest pH and be the most basic?

D

Acid **pK_a**

HA 4.00

HB 7.00

HC 10.00

HD 11.00

Stronger the acid

↓

Weaker the conjugate base

A. NaA

B. NaB

C. NaC

D. NaD

E. All will have the same pH because the concentrations are the same.

acid: $\text{HA} > \text{HB} > \text{HC} > \text{HD}$

Base: $\text{A}^- < \text{B}^- < \text{C}^- < \text{D}^-$

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32. Which of the groups, A–D, consist of salts that all form basic solutions in water?

D

A. NaNO_3 , NH_4CN , NaOOCCH_3 , NH_4Cl
Neutral depends Basic acidic

B. NaHCO_3 , NaF , NH_4Cl , Na_2SO_3
depends Basic acidic Basic

C. Na_2CO_3 , KCl , NaOOCCH_3 , NH_4Cl
Basic neutral Basic acidic

D. Na_2CO_3 , NaF , NaOOCCH_3 , NaCN
Basic Basic Basic Basic

E. All of the above.

* All of them are con. Base of Weak acid
 ↓
 Basic solution

33. Which one of these is correct?

A

A. K_2SO_3 is a stronger base than KHSO_3 .
 * SO_3^{2-} is the Conj. base of HSO_3^-
 ← so SO_3^{2-} is more basic than HSO_3^-

B. Na_2HPO_4 is a weaker base than NaH_2PO_4 .

C. K_2CO_3 is a weaker base than KHCO_3 .

D. NaHSO_3 is a stronger acid than NaHSO_4 .

E. All of these statements are correct.

* $\text{HPO}_4^{2-} \rightarrow \text{H}_2\text{PO}_4^-$
 Con. Base Con. acid

* $\text{CO}_3^{2-} \rightarrow \text{HCO}_3^-$
 Con. Base acid

* H_2SO_4 is a stronger acid than H_2SO_3 .
 so, HSO_3^- is stronger base than HSO_4^-

34. An aqueous solution of an unknown acid had a pH of 3.70. Titration of a 25.0 mL aliquot of the acid solution required 21.7 mL of 0.104 M aqueous sodium hydroxide for complete reaction. Assuming that the acid is monoprotic, what is its ionization constant?

4.4×10^{-7}

$$25.0 \times M_A = 21.7 \times 0.104$$

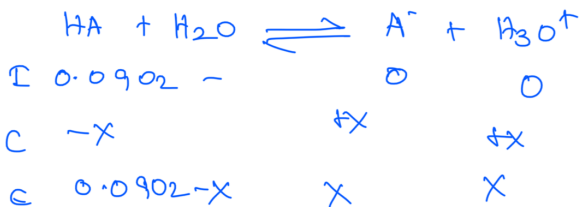
$$M_A = 0.0902 \text{ M}$$

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-3.70} = 0.0001995$$

$$K_a = \frac{x^2}{0.0902}$$

$$K_a = \frac{(0.0001995)^2}{0.0902}$$

$$K_a = 4.41 \times 10^{-7}$$



35. Which of these mixtures would result in a buffered solution when 1.0 L of each of the two solutions are mixed.

C

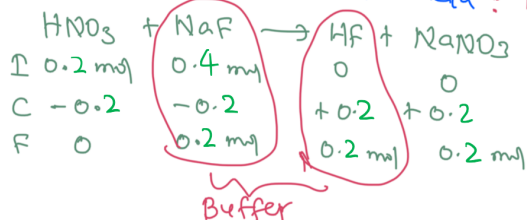
A. 0.2 M HNO_3 and 0.2 M NaNO_3 ← Strong acid and Conjugate base of Strong acid
 No Buffer

B. 0.2 M HNO_3 and 0.4 M HF ← Strong acid and weak acid : No Buffer

C. 0.2 M HNO_3 and 0.4 M NaF

D. 0.2 M HNO_3 and 0.4 M NaOH

↑
 Strong acid + Strong Base
 No Buffer



Recitation Worksheet 12 – Exam 4 Review

36. The pH of a solution of $\text{NH}_4\text{C}_2\text{H}_3\text{O}_2$ is approximately 7. The best explanation is:

D

A. This salt does not react with water.

B. Ammonium acetate is a weak electrolyte.

C. All salts of weak acids and weak bases are neutral.

D. Aqueous ammonia and acetic acid have approximately equal ionization constants.

E. The salt is a product of a strong acid and a strong base.



$$K_a \text{ of } \text{NH}_4^+ = K_b \text{ of } \text{CH}_3\text{COO}^-$$

$$K_b \text{ of } \text{CH}_3\text{COO}^- = K_a \text{ of } \text{CH}_3\text{COOH}$$

$$K_a (\text{NH}_4^+) \approx K_b (\text{CH}_3\text{COO}^-)$$

$$\text{pH} \approx 7$$

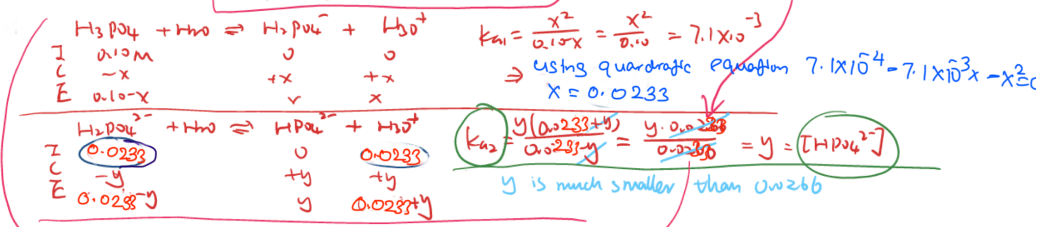
37. What is the $[\text{HPO}_4^{2-}]$ of a solution labeled "0.10 M phosphoric acid"?

$$6.3 \times 10^{-8} \text{ M}$$

$$[K_{a1} = 7.1 \times 10^{-3}; K_{a2} = 6.3 \times 10^{-8}; K_{a3} = 4.2 \times 10^{-13}]$$

value of x does not matter. you can directly use $[\text{HPO}_4^{2-}] = K_{a2}$

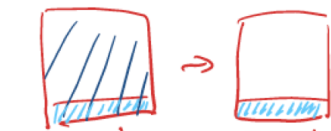
$$\text{Conclusion: } [\text{HPO}_4^{2-}] = K_{a2} = 6.3 \times 10^{-8}$$



x canceled, no need to calculate anything!!

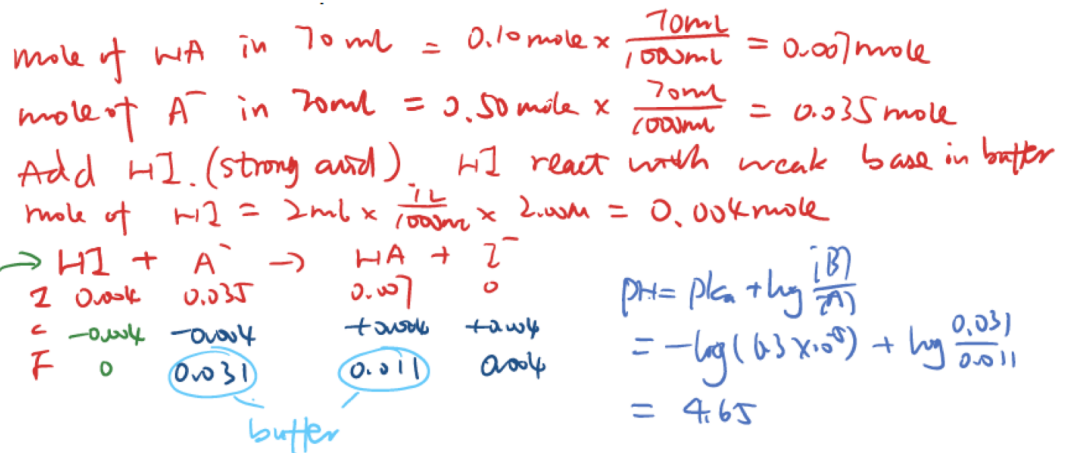
38. A pH 4.88 buffer was prepared by dissolving 0.10 mol of benzoic acid ($K_a = 6.3 \times 10^{-5}$) and 0.50 mol of sodium benzoate in sufficient pure water to form a 1.00 L solution. To a 70.0 mL aliquot of this solution was added 2.00 mL of 2.00 M aqueous HI solution. What was the pH of the new 72.0 mL solution?

$$4.65$$



1.00 L
0.1 mole HA
0.5 mole A^-

70 mL
? mole HA
? mole A^-

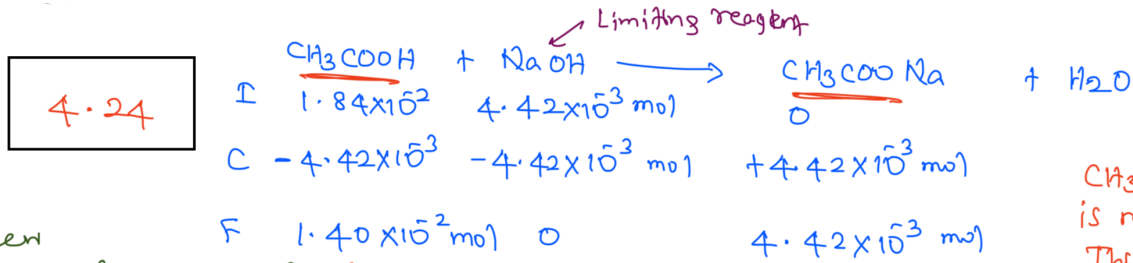


Recitation Worksheet 12 – Exam 4 Review

39. What is the pH of the solution when 52.60 mL of 0.35 M acetic acid is added to 22.08 mL of 0.20 M NaOH? The K_a for acetic acid is 1.8×10^{-5} .

$$1.84 \times 10^{-2} \text{ mol}$$

$$4.42 \times 10^{-3} \text{ mol}$$



$$[\text{CH}_3\text{COOH}] = \frac{1.40 \times 10^{-2} \text{ mol}}{0.07468 \text{ L}} = 0.187 \text{ M}$$

$$[\text{CH}_3\text{COONa}] = \frac{4.42 \times 10^{-3} \text{ mol}}{0.07468 \text{ L}} = 0.0592 \text{ M}$$

$$\text{pH} = \text{p}K_a + \log \frac{[\text{Base}]}{[\text{acid}]}$$

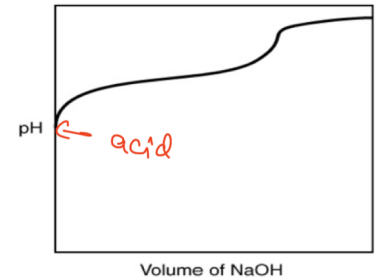
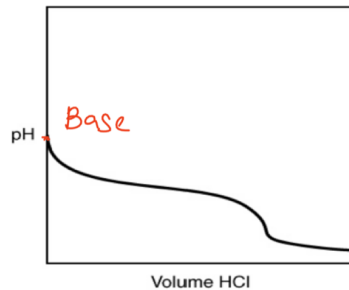
$$\text{pH} = 4.74 + \log \left(\frac{0.0592}{0.187} \right)$$

$$\text{pH} = 4.24$$

CH_3COOH & CH_3COONa is remaining.
This is a buffer

40. A 0.500 g sample of an unknown substance was titrated with a 0.1 M HCl solution. Another 0.500 g sample of it was titrated with a 0.1 M NaOH solution. The resulting titration curves are illustrated here. Given the following possibilities, what is the sample?

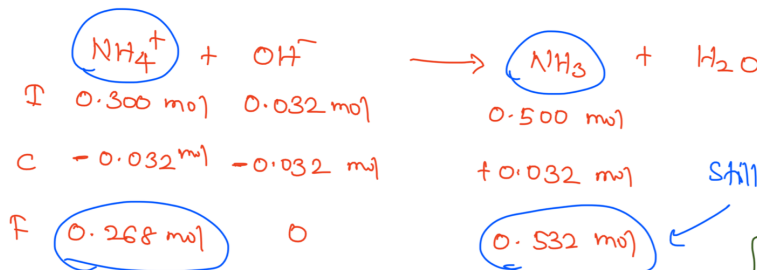
- C
- A. $\text{Na}_2\text{CO}_3 \leftarrow \text{Base}$
B. $\text{CO}_2 + \text{H}_2\text{O} \leftrightarrow \text{H}_2\text{CO}_3 \leftarrow \text{acid}$
C. $\text{NaHCO}_3 \rightarrow \text{Na}^+ + \text{HCO}_3^-$
D. $\text{H}_2\text{CO}_3 \leftarrow \text{acid}$ amphoteric
E. There is no way to tell.



Must be an amphoteric species

41. Consider 1.00 L of a solution initially containing 0.500 mol ammonia (NH_3) and 0.300 mol of ammonium ion (NH_4^+). What is the pH after addition of 40 mL of 0.800M NaOH to this solution? ($\text{NH}_4^+ K_a = 5.6 \times 10^{-10}$)?

$$9.55$$



$$\frac{[\text{NH}_3]}{[\text{NH}_4^+]}$$

Still you have a buffer

$$[\text{NH}_4^+] = \frac{0.268 \text{ mol}}{1.04 \text{ L}} = 0.258 \text{ M}$$

$$[\text{NH}_3] = \frac{0.532 \text{ mol}}{1.04 \text{ L}} = 0.512 \text{ M}$$

$$\text{pH} = \text{p}K_a + \log \frac{[\text{Base}]}{[\text{acid}]}$$

$$\text{pH} = 9.25 + \log \left(\frac{0.512}{0.258} \right)$$

$$\text{pH} = 9.55$$

Recitation Worksheet 12 – Exam 4 Review

42. Of the following substances, which one(s) will form basic solutions (select all that apply).
 $(\text{H}_2\text{CO}_3 \text{ } K_{a1} = 4.4 \times 10^{-7}, K_{a2} = 4.7 \times 10^{-11}; \text{H}_2\text{S} \text{ } K_{a1} = 1.1 \times 10^{-7}, K_{a2} = 1.0 \times 10^{-19})$

E

(D) (B)
NaHS

(A) (N)
 $\text{Cu}(\text{NO}_3)_2$

(N) (B)
 KHCO_3

(N) (B)
NaF

$\text{H}_2\text{S} \text{ } K_{a1} = 1.1 \times 10^{-7}$
 $\text{HS}^- \text{ } K_{a2} = 1.0 \times 10^{-19}$

A) NaHS, $\text{Cu}(\text{NO}_3)_2$

B) KHCO_3 , NaHS

C) NaF only

D) NaF, KHCO_3

(E) NaHS, KHCO_3 and NaF

$\text{NaHS} \leftarrow \text{Na}^+ \leftrightarrow \text{NaOH}$ strong base $\Rightarrow \text{Na}^+$ neutral

$\text{HS}^- \leftarrow$ as acid: $\text{HS}^- + \text{H}_2\text{O} \leftrightarrow \text{H}_2\text{S} + \text{OH}^-$, $K_a \text{ of } \text{HS}^- = K_{a2} \text{ of } \text{H}_2\text{S} = 1.0 \times 10^{-19}$
 as base: $\text{HS}^- + \text{H}_2\text{O} \leftrightarrow \text{H}_2\text{S} + \text{OH}^-$, $K_b \text{ of } \text{HS}^- = \frac{K_w}{K_a \text{ of } \text{H}_2\text{S}} = \frac{1 \times 10^{-14}}{1.1 \times 10^{-7}} = 9.1 \times 10^{-8}$ } basic

$\text{Cu}^{2+} + 2\text{H}_2\text{O} \leftrightarrow \text{Cu}(\text{OH})_2 + 2\text{H}^+$ weak base $\Rightarrow \text{Cu}^{2+}$ acidic
 $\text{NO}_3^- \leftrightarrow \text{HNO}_3$ strong acid $\Rightarrow \text{NO}_3^-$ neutral } acidic

$\text{H}_2\text{CO}_3 \leftarrow \text{K}^+ \leftrightarrow \text{KOH}$ strong base $\Rightarrow \text{K}^+$ neutral

$\text{HCO}_3^- \leftarrow$ as acid: $\text{HCO}_3^- + \text{H}_2\text{O} \leftrightarrow \text{CO}_3^{2-} + \text{H}_3\text{O}^+$, $K_a \text{ of } \text{HCO}_3^- = K_{a2} \text{ of } \text{H}_2\text{CO}_3 = 4.7 \times 10^{-11}$
 as base: $\text{HCO}_3^- + \text{H}_2\text{O} \leftrightarrow \text{H}_2\text{CO}_3 + \text{OH}^-$, $K_b \text{ of } \text{HCO}_3^- = \frac{K_w}{K_a \text{ of } \text{H}_2\text{CO}_3} = \frac{1 \times 10^{-14}}{4.4 \times 10^{-7}} = 2.3 \times 10^{-8}$ } basic

$\text{NaF} \leftarrow \text{Na}^+ \leftrightarrow \text{NaOH}$ strong base $\Rightarrow \text{Na}^+$ neutral

$\text{F}^- + \text{H}_2\text{O} \leftrightarrow \text{HF} + \text{OH}^-$ weak acid $\Rightarrow \text{F}^-$ basic } basic

43. An aqueous solution of NaF is prepared by dissolving 0.350 mol of NaF in sufficient water to yield 1.0 L of

2.1 $\times 10^{-10}$

$\text{NaF} \leftarrow \text{Na}^+ \leftrightarrow \text{NaOH}$ strong base $\Rightarrow \text{Na}^+$ neutral \Rightarrow does not change pH \Rightarrow ignore Na^+
 $\text{F}^- + \text{H}_2\text{O} \leftrightarrow \text{HF} + \text{OH}^-$ weak acid $\Rightarrow \text{F}^-$ basic

$$[\text{NaF}] = \frac{0.350 \text{ mole}}{1 \text{ L}} = 0.350 \text{ M} = [\text{F}^-]$$

$$\text{pOH} = 14 - \text{pH} = 14 - 8.93 = 5.07$$

$$[\text{OH}^-] = 10^{-\text{pOH}} = 10^{-5.07} = 8.51 \times 10^{-6} \text{ M}$$

$$\text{F}^- + \text{H}_2\text{O} \leftrightarrow \text{HF} + \text{OH}^-$$

| | | | |
|---|-------------------------------|------------------------|------------------------|
| I | 0.350 | 0 | 0 |
| C | -8.51×10^{-6} | $+8.51 \times 10^{-6}$ | $+8.51 \times 10^{-6}$ |
| E | $0.350 - 8.51 \times 10^{-6}$ | 8.51×10^{-6} | 8.51×10^{-6} |

$$K_b = \frac{(8.51 \times 10^{-6}) \times (8.51 \times 10^{-6})}{0.350 - 8.51 \times 10^{-6}} = 2.069 \times 10^{-10} = 2.1 \times 10^{-10}$$

44. The acid-dissociation constant, K_a , for an unknown acid HA is 4.57×10^{-3} . What is the base-dissociation constant, K_b , for the unknown anion A^- ?

2.19 $\times 10^{-12}$

$$K_b \text{ of } \text{A}^- = \frac{K_w}{K_a \text{ of } \text{HA}} = \frac{1 \times 10^{-14}}{4.57 \times 10^{-3}} = 2.19 \times 10^{-12}$$

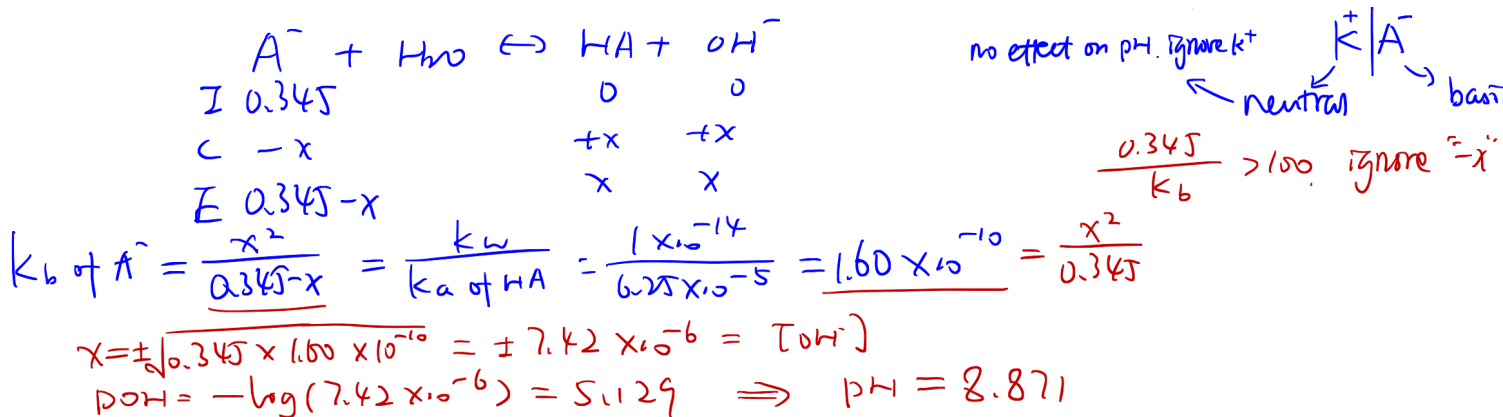
Recitation Worksheet 12 – Exam 4 Review

45. A certain acid, HA, has a K_a given by:

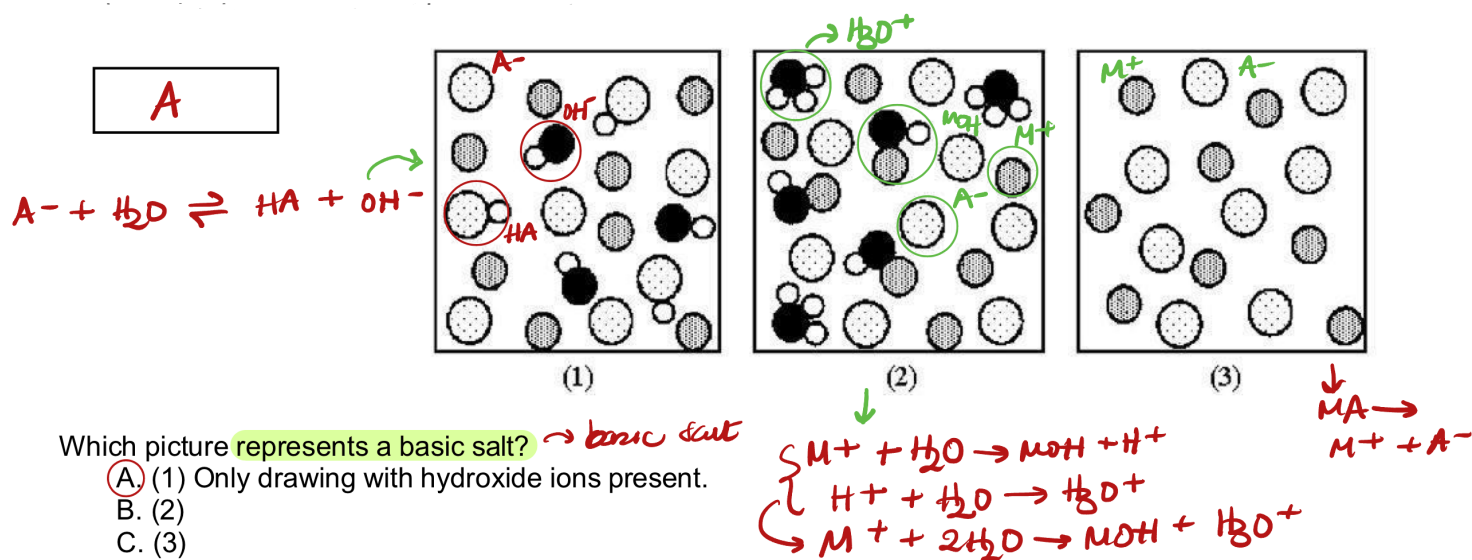


What is the pH of a 0.345 M aqueous solution of the acid's potassium salt, KA, which undergoes the hydrolysis reaction?

8.871



46. The pictures represent solutions of three salts MA; water molecules have been omitted for clarity. Dotted spheres represent A^- ions; gray spheres represent M^+ ions; black spheres represent oxygen atoms; and unshaded spheres represent hydrogen atoms.

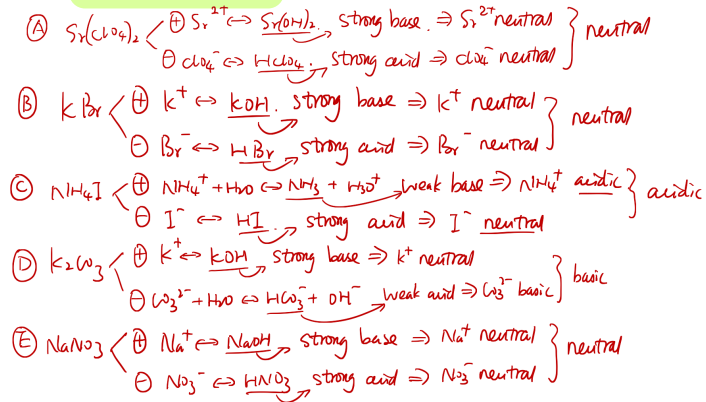


Recitation Worksheet 12 – Exam 4 Review

47. Which of the following salts will produce an acidic solution?

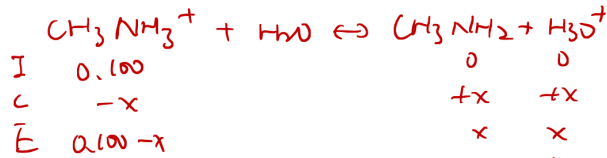
C

- A) $\text{Sr}(\text{ClO}_4)_2$
 B) KBr
 C) NH_4I
 D) K_2CO_3
 E) NaNO_3



48. Calculate the pH of a 0.100 M $\text{CH}_3\text{NH}_3\text{Cl}$ solution. K_b for methylamine, CH_3NH_2 , is 3.7×10^{-4} .

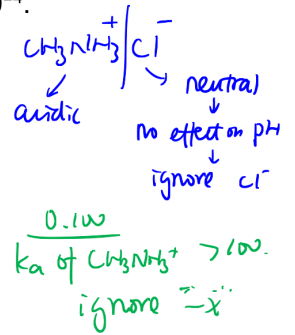
5.784



$$K_a \text{ of } \text{CH}_3\text{NH}_3^+ = \frac{x^2}{0.100-x} = \frac{K_w}{K_b \text{ of } \text{CH}_3\text{NH}_2} = \frac{1 \times 10^{-14}}{3.7 \times 10^{-4}} = 2.7 \times 10^{-11} = \frac{x^2}{0.100}$$

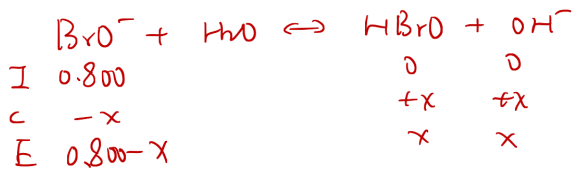
$$x = \pm \sqrt{0.100 \times 2.7 \times 10^{-11}} = 1.643 \times 10^{-6} = [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log(1.643 \times 10^{-6}) = 5.784$$



49. Calculate the pH of a 0.800 M KBrO solution. K_a for hypobromous acid, HBrO , is 2.0×10^{-9} .

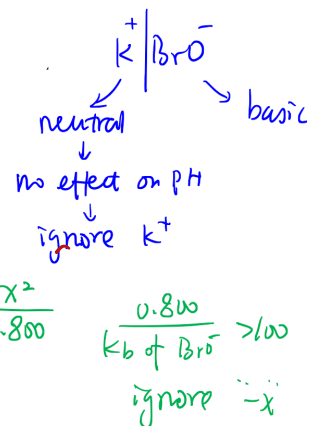
11.301



$$K_b \text{ of } \text{BrO}^- = \frac{x^2}{0.800-x} = \frac{K_w}{K_a \text{ of } \text{HBrO}} = \frac{1 \times 10^{-14}}{2.0 \times 10^{-9}} = 5.0 \times 10^{-6} = \frac{x^2}{0.800}$$

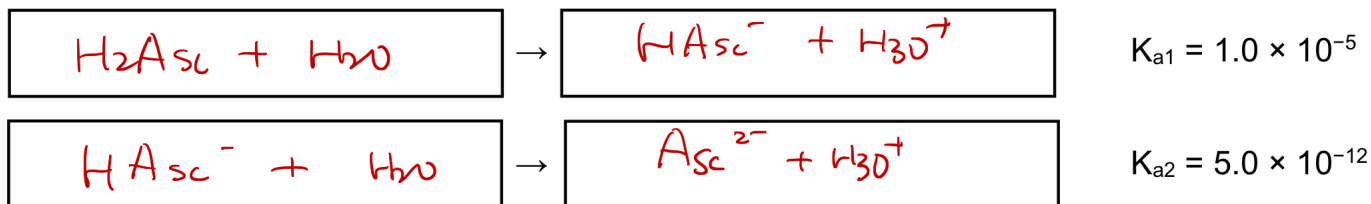
$$x = \pm \sqrt{0.800 \times 5.0 \times 10^{-6}} = \pm 2.00 \times 10^{-3} = [\text{OH}^-]$$

$$\text{pOH} = -\log(2.00 \times 10^{-3}) = 2.699 \Rightarrow \text{pH} = 11.301$$



Recitation Worksheet 12 – Exam 4 Review

50. Ascorbic acid, $\text{H}_2\text{C}_6\text{H}_6\text{O}_6$ is a diprotic acid, with $K_{a1} = 1.0 \times 10^{-5}$ and $K_{a2} = 5.0 \times 10^{-12}$. It is often abbreviated as **H2Asc**. Using this abbreviation to write out the equilibria of this acid with water. If you type your answers in the pdf, please use “^” for superscripts and “_” for subscripts.



51. Predict whether each of the following salt solutions will be **A**) acidic, **B**) basic, or **C**) neutral. Remember that ionic compounds dissociate completely in water.

$$K_a \text{ for } \text{HNO}_2 = 4.0 \times 10^{-4}$$

$$K_b \text{ for } \text{NH}_3 = 1.8 \times 10^{-5}$$

$$\text{H}_3\text{PO}_4: K_{a1} = 7.5 \times 10^{-3}$$

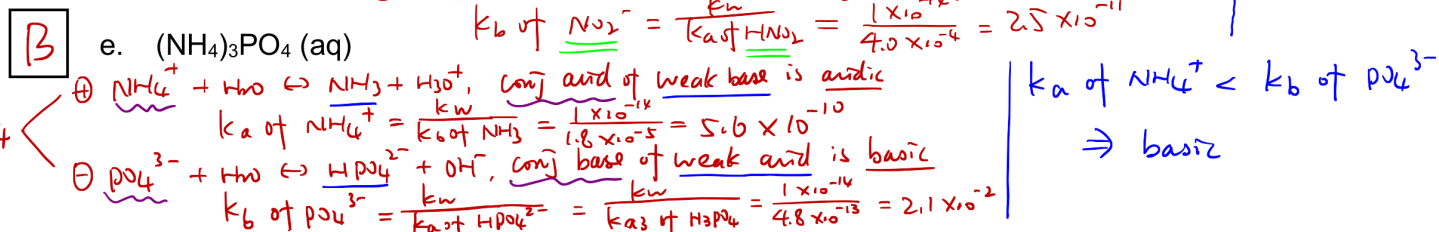
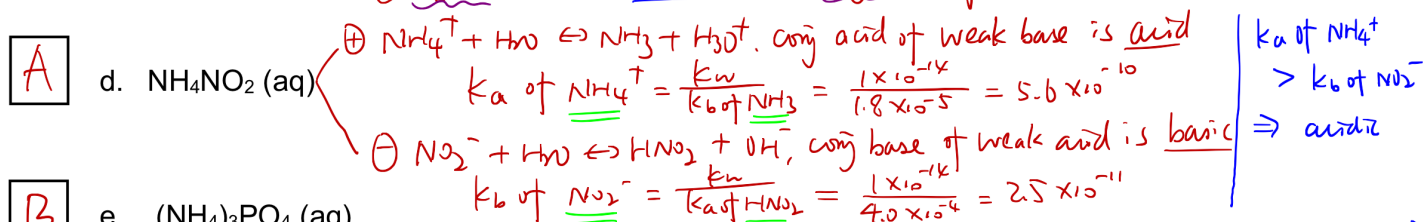
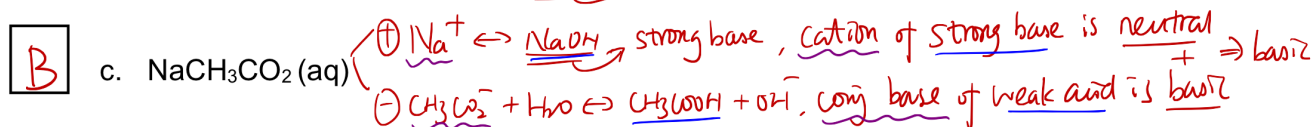
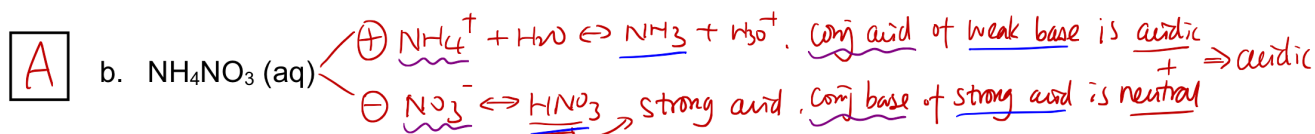
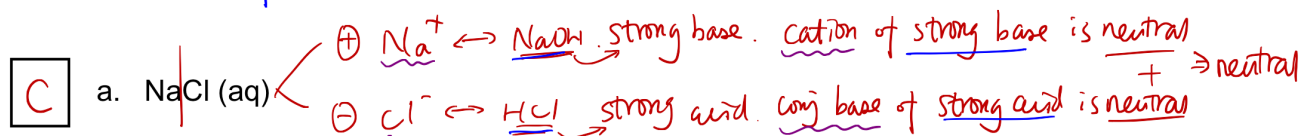
$$K_{a2} = 6.2 \times 10^{-8}$$

$$K_{a3} = 4.8 \times 10^{-13}$$

$$(K_a \text{ of } \text{H}_3\text{PO}_4)$$

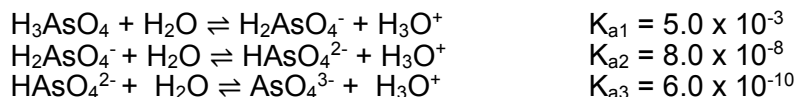
$$(K_a \text{ of } \text{H}_2\text{PO}_4^-)$$

$$(K_a \text{ of } \text{HPO}_4^{2-})$$



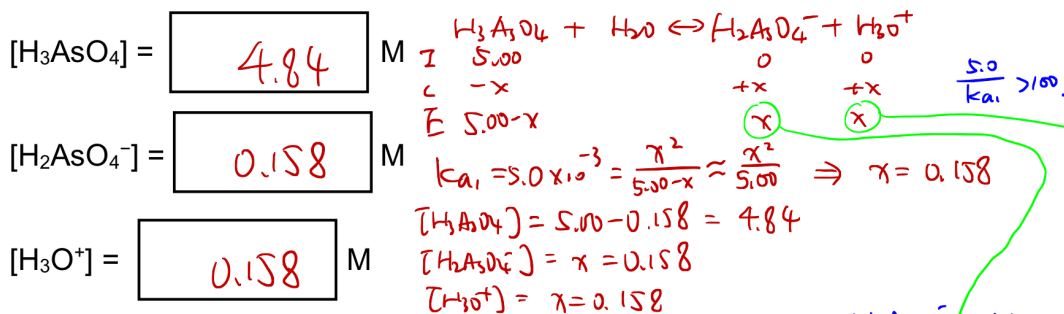
Recitation Worksheet 12 – Exam 4 Review

52. Arsenic acid, H_3AsO_4 is a triprotic acid. It has three equilibrium associated with its reaction with water:



If we have a 5.00 M solution of H_3AsO_4 , what are the concentrations of all of the species present?

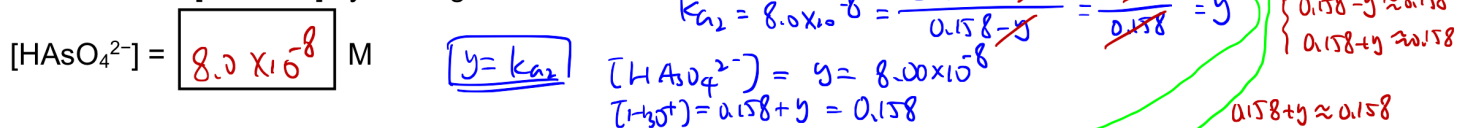
a. First, calculate the following concentrations using the first ionization.



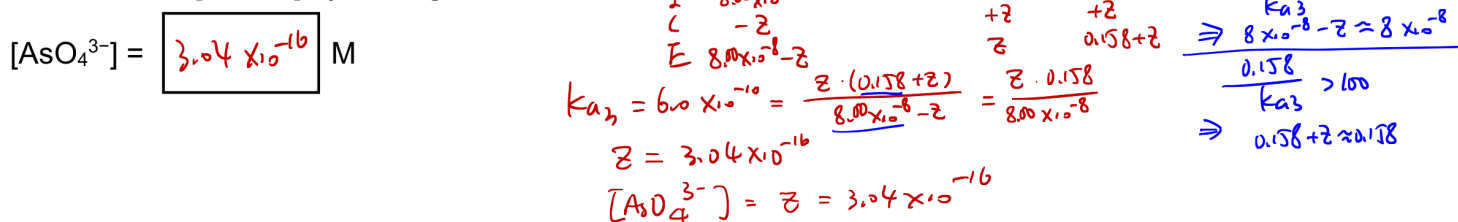
b. What is the pH of this solution?

$$\text{pH} = \boxed{0.801} \quad \text{pH} = -\log(0.158) = 0.801$$

c. Calculate the $[\text{HAsO}_4^{2-}]$ by solving for K_{a2} .



d. Calculate the $[\text{AsO}_4^{3-}]$ by solving for K_{a3} .



Recitation Worksheet 12 – Exam 4 Review

53. For which of these salts does the solubility increase at low pH? *more acidic, more H_3O^+*

B

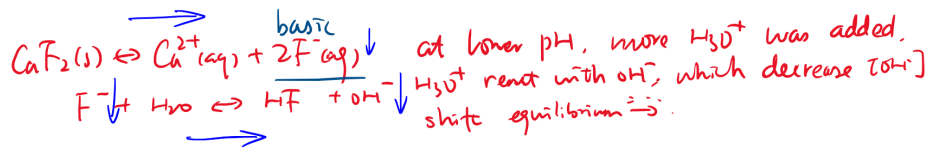
A. $Ca(NO_3)_2$ *neutral*

B. CaF_2 *basic*

C. $CaCl_2$ *neutral*

D. $CaBr_2$ *neutral*

E. CaI_2 *neutral*



54. Copper(II) fluoride has a solubility of 0.0020 mol/L. What is the value of K_{sp} ?

C

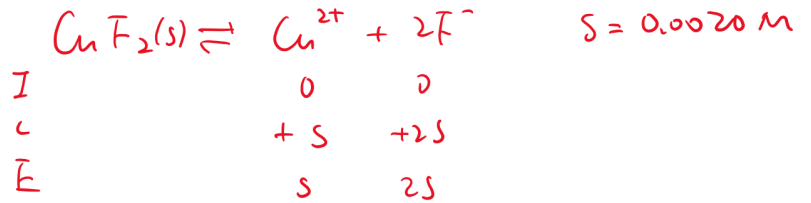
A. 1.8×10^{-7}

B. 4.0×10^{-6}

C. 3.2×10^{-8}

D. 8.0×10^{-9}

E. None of these are correct.



$$S = 0.0020 M$$

$$K_{sp} = S \cdot (2S)^2 = 4S^3 = 4 \times (0.0020)^3 = 3.2 \times 10^{-8}$$

55. The K_{sp} of $AgCl$ is 1.7×10^{-10} . How many moles of $MnCl_2$ can be dissolved in one liter of an aqueous solution in which $[AgNO_3] = 3.4 \times 10^{-4} M$ before a precipitate appears?

E

A. $5.0 \times 10^{-7} mol$

B. $5.8 \times 10^{-14} mol$

C. $2.4 \times 10^{-7} mol$

D. $2.0 \times 10^{-6} mol$

E. $2.5 \times 10^{-7} mol$

$$K_{sp} = [Ag^+][Cl^-] = 3.4 \times 10^{-4} \times [Cl^-] = 1.7 \times 10^{-10}$$

$$[Cl^-] = 5.0 \times 10^{-7} M$$

$$[MnCl_2] = [Cl^-] \times \frac{1 MnCl_2}{2 Cl^-} = 5.0 \times 10^{-7} \times \frac{1}{2} = 2.5 \times 10^{-7} M$$

Recitation Worksheet 12 – Exam 4 Review

56. When equal volumes of the indicated aqueous solutions are mixed, precipitation should occur only for:

| <u>Salt</u> | <u>K_{sp}</u> |
|--------------------|-----------------------|
| barium fluoride | 1.0×10^{-6} |
| calcium carbonate | 2.8×10^{-9} |
| calcium fluoride | 5.3×10^{-9} |
| magnesium fluoride | 3.7×10^{-9} |
| silver carbonate | 8.5×10^{-12} |

B

- A. $2 \times 10^{-5} \text{ M Ag}^+ + 2 \times 10^{-5} \text{ M CO}_3^{2-}$
 B. $2 \times 10^{-4} \text{ M Ca}^{2+} + 2 \times 10^{-4} \text{ M CO}_3^{2-}$
 C. $2 \times 10^{-5} \text{ M Ca}^{2+} + 2 \times 10^{-3} \text{ M F}^-$
 D. $2 \times 10^{-5} \text{ M Mg}^{2+} + 2 \times 10^{-6} \text{ M F}^-$
 E. $2 \times 10^{-3} \text{ M Ba}^{2+} + 2 \times 10^{-3} \text{ M F}^-$

When equal volume of solutions are mixed, their concentrations are halved.

$$A. Q_{sp} = [Ag^+]^2 [CO_3^{2-}] = \left(\frac{2 \times 10^{-5}}{2}\right)^2 \times \left(\frac{2 \times 10^{-5}}{2}\right) = 2 \times 10^{-15} < K_{sp}(Ag_2CO_3)$$

$$B. Q_{sp} = [Ca^{2+}] [CO_3^{2-}] = \left(\frac{2 \times 10^{-4}}{2}\right) \times \left(\frac{2 \times 10^{-4}}{2}\right) = 1 \times 10^{-8} > K_{sp}(CaCO_3)$$

precipitation

$$C. Q_{sp} = [Ca^{2+}] [F^-]^2 = \left(\frac{2 \times 10^{-5}}{2}\right) \times \left(\frac{2 \times 10^{-3}}{2}\right)^2 = 1 \times 10^{-11} < K_{sp}(CaF_2)$$

$$D. Q_{sp} = [Mg^{2+}] [F^-]^2 = \left(\frac{2 \times 10^{-5}}{2}\right) \times \left(\frac{2 \times 10^{-6}}{2}\right)^2 = 1 \times 10^{-17} < K_{sp}(MgF_2)$$

$$E. Q_{sp} = [Ba^{2+}] [F^-]^2 = \left(\frac{2 \times 10^{-3}}{2}\right) \times \left(\frac{2 \times 10^{-3}}{2}\right)^2 = 1 \times 10^{-9} < K_{sp}(BaF_2)$$

Recitation Worksheet 12 – Exam 4 Review

57. Calculate the concentration of copper(II) ion in $\text{Cu}(\text{OH})_2$, in ppm, that can remain at equilibrium in a solution having a $[\text{OH}^-] = 1.0 \times 10^{-4} \text{ M}$. $K_{\text{sp}} = 1.6 \times 10^{-19}$. Density of solution is 1.0 g/mL .

$$1.0 \times 10^{-6}$$



$$K_{\text{sp}} = [\text{Cu}^{2+}][\text{OH}^-]^2$$

$$[\text{Cu}^{2+}] = \frac{K_{\text{sp}}}{[\text{OH}^-]^2} = \frac{1.6 \times 10^{-19}}{(1.0 \times 10^{-4})^2} = 1.6 \times 10^{-11} \text{ M}$$

Assume 1 L of solution.

$$\text{g of solute } \text{Cu}^{2+} = 1.6 \times 10^{-11} \text{ M} \times 1 \text{ L} \times 63.55 \frac{\text{g}}{\text{mol}} = 1.02 \times 10^{-9} \text{ g}$$

$$\text{g of solution} = 1 \text{ L} \times \frac{1000 \text{ mL}}{1 \text{ L}} \times 1.00 \text{ g/mL} = 1000 \text{ g} \quad \text{ppm} = \frac{1.02 \times 10^{-9} \text{ g}}{1000 \text{ g}} \times 10^6 = 1.0 \times 10^{-1}$$

58. To a concentrated buffer of pH 9.0 was added an equal volume of an aqueous solution that was 0.20 M in each of the ions Ca^{2+} , Cd^{2+} , and Cu^{2+} . The expected precipitate would consist of:

salt: calcium hydroxide cadmium hydroxide copper(II) hydroxide

K_{sp} : 4.0×10^{-6} 2.0×10^{-14} 1.8×10^{-19}

D

- A. only $\text{Ca}(\text{OH})_2$
- B. only $\text{Cd}(\text{OH})_2$
- C. only $\text{Cu}(\text{OH})_2$
- ☒ D. only $\text{Cd}(\text{OH})_2$ and $\text{Cu}(\text{OH})_2$
- E. $\text{Ca}(\text{OH})_2$, $\text{Cd}(\text{OH})_2$, and $\text{Cu}(\text{OH})_2$

$\text{pH} = 9.0$ $\text{pOH} = 14.00 - 9.0 = 5.0$ $[\text{OH}^-] = 10^{-\text{pOH}} = 10^{-5.0} = 1 \times 10^{-5}$

$\text{Ca}(\text{OH})_2$: $Q = [\text{Ca}^{2+}][\text{OH}^-]^2 = 0.20 \times (1 \times 10^{-5})^2 = 2.5 \times 10^{-12} < K_{\text{sp}} \text{Ca}(\text{OH})_2$

$\text{Cd}(\text{OH})_2$: $Q = [\text{Cd}^{2+}][\text{OH}^-]^2 = 0.20 \times (1 \times 10^{-5})^2 = 2.5 \times 10^{-12} > K_{\text{sp}} \text{Cd}(\text{OH})_2$

$\text{Cu}(\text{OH})_2$: $Q = [\text{Cu}^{2+}][\text{OH}^-]^2 = 0.20 \times (1 \times 10^{-5})^2 = 2.5 \times 10^{-12} > K_{\text{sp}} \text{Cu}(\text{OH})_2$

59. A small amount of solid magnesium hydroxide is shaken vigorously in a test tube almost full of water until no further change occurs and most of the solid settles out. The resulting solution is:

B

- A. concentrated and saturated
- ☒ B. dilute and saturated
- C. dilute and unsaturated
- D. dilute and supersaturated
- E. concentrated and supersaturated

\rightarrow precipitation. saturated solution
 $\text{Mg}(\text{OH})_2$ has very small solubility (considered as insoluble). \Rightarrow diluted.

60. What is the molar solubility of $\text{Mg}(\text{OH})_2(\text{s})$ in a basic aqueous solution with a pH of 12.50? K_{sp} for $\text{Mg}(\text{OH})_2(\text{s})$ is 5.6×10^{-12} .

B

- A. $1.8 \times 10^{-10} \text{ mol/L}$
- ☒ B. $5.6 \times 10^{-9} \text{ mol/L}$
- C. $2.4 \times 10^{-6} \text{ mol/L}$
- D. $1.1 \times 10^{-4} \text{ mol/L}$

$\text{pH} = 12.50$ $\text{pOH} = 14.00 - 12.50 = 1.50$

$[\text{OH}^-] = 10^{-\text{pOH}} = 10^{-1.50} = 3.16 \times 10^{-2}$

$\text{Mg}(\text{OH})_2 \rightleftharpoons \text{Mg}^{2+} + 2\text{OH}^-$

| | | |
|---|----|--------------------------------|
| I | 0 | 3.16×10^{-2} |
| C | +S | +2S |
| E | S | $(3.16 \times 10^{-2} + 2S)^2$ |

$K_{\text{sp}} = (S) \cdot (3.16 \times 10^{-2} + 2S)^2 = 5.6 \times 10^{-12}$

$S = 5.6 \times 10^{-9} \text{ M}$

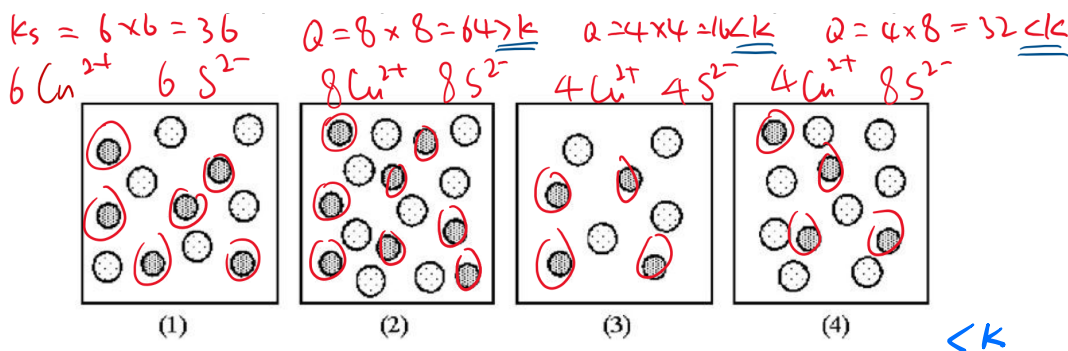
Recitation Worksheet 12 – Exam 4 Review

61. The dissolution of barium hydroxide in water is an endothermic process with a K_{sp} of 5.00×10^{-4} at 298 K. Which of these will increase the solubility?

D

- Same compound, no change*
 $Ba(OH)_2(s) \rightarrow Ba^{2+} + 2OH^{-}$
increase OH^{-} shift \leftarrow
 A. Barium hydroxide is added to the solution.
For most of solids, solubility increase with temperature.
 B. Sodium hydroxide (NaOH) is added to the solution.
adding H^{+} would decrease OH^{-} , shift \rightarrow
 C. The temperature is decreased.
 D. HCl is added to the mixture.

62. The pictures represent solutions of CuS and may also contain ions other than Cu^{2+} and S^{2-} which are not shown. Gray spheres represent Cu^{2+} ions and dotted spheres represent S^{2-} ions.



If solution (1) is a saturated solution of CuS, which of solutions (2)-(4) are unsaturated?

D

- A. only (2)
 B. only (3)
 C. only (4)
 D. (3) and (4)

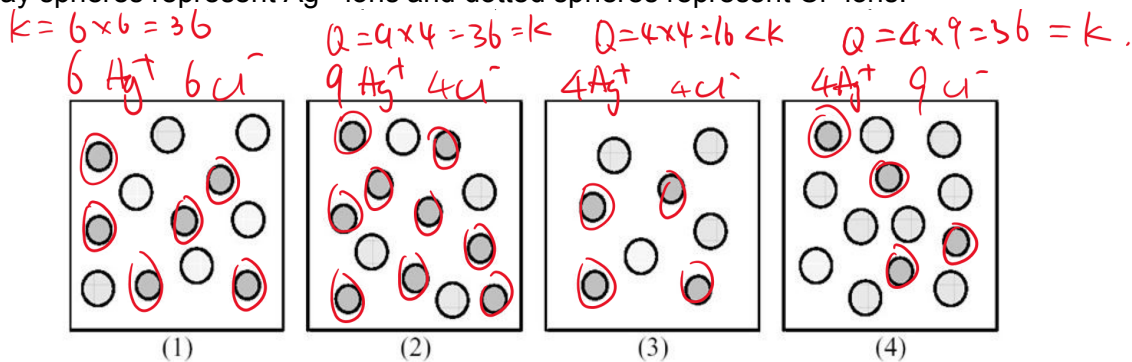
63. In which of the following pairs is the oxidation number for the underlined element **incorrect**?

B

- A. $\underline{Mn}O_4^{-} / (+7)$ (A) MnO_4^{-} $Mn + (-2)4 = -1 \rightarrow Mn = +7$
 B. $\underline{S}O_4^{2-} / (+4)$ (B) SO_4^{2-} $S + (-2)4 = -2 \rightarrow S = +6$
 C. $\underline{N}H_4^{+} / (-3)$ (C) NH_4^{+} $N + (+1)4 = +1 \rightarrow N = -3$
 D. $\underline{N}O_3^{-} / (+5)$ (D) NO_3^{-} $N + (-2)3 = -1 \rightarrow N = +5$
 E. $\underline{Cr}_2O_7^{2-} / (+6)$ (E) $Cr_2O_7^{2-}$ $2Cr + (-2)7 = -2 \rightarrow Cr = +6$

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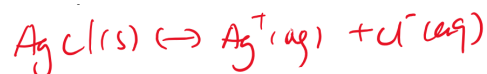
64. The pictures represent solutions of AgCl and may contain ions other than Ag^+ and Cl^- that are not shown. Gray spheres represent Ag^+ ions and dotted spheres represent Cl^- ions.



If solution (1) is a saturated solution of AgCl, which of solutions (1)-(4) represents the solution after a small amount of HNO_3 is added and equilibrium is restored?

A

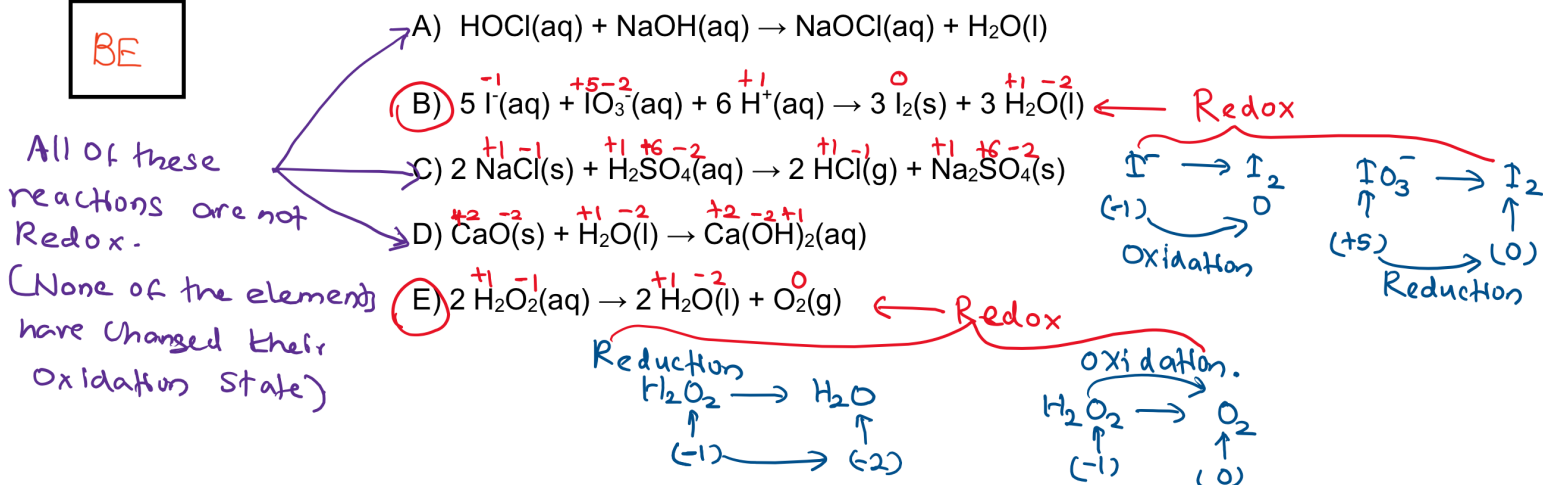
- (A) (1)
B. (2)
C. (3)
D. (4)



HNO_3 does not react with $\text{Ag}^+(aq)$ or $\text{Cl}^-(aq)$ or $\text{AgCl}(s)$.
So there is no reaction. thus, there is no change
(HNO_3 also does not contain common ions as AgCl)

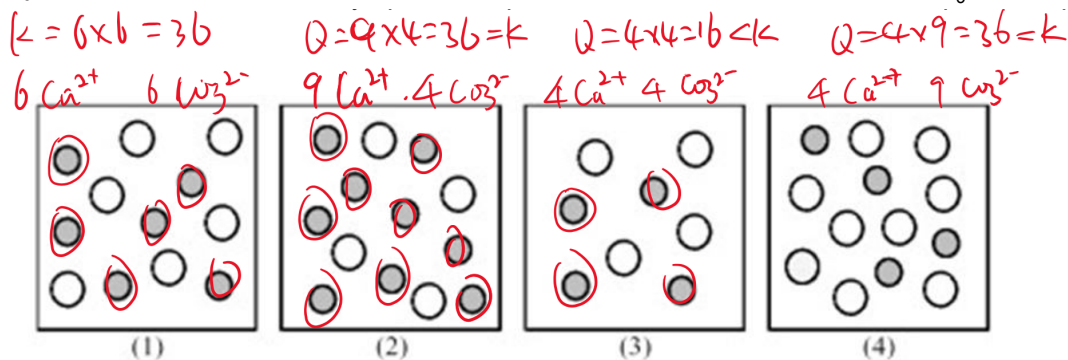
65. Identify the reaction(s) that are redox reactions:

BE



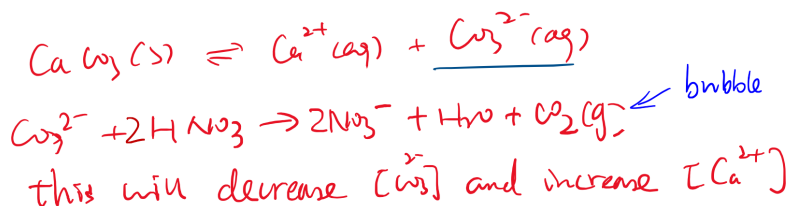
Recitation Worksheet 12 – Exam 4 Review

66. The pictures represent solutions of CaCO_3 . The solution may contain ions other than Ca^{2+} and CO_3^{2-} which are not shown. Gray spheres represent Ca^{2+} ions and unshaded spheres represent CO_3^{2-} ions.

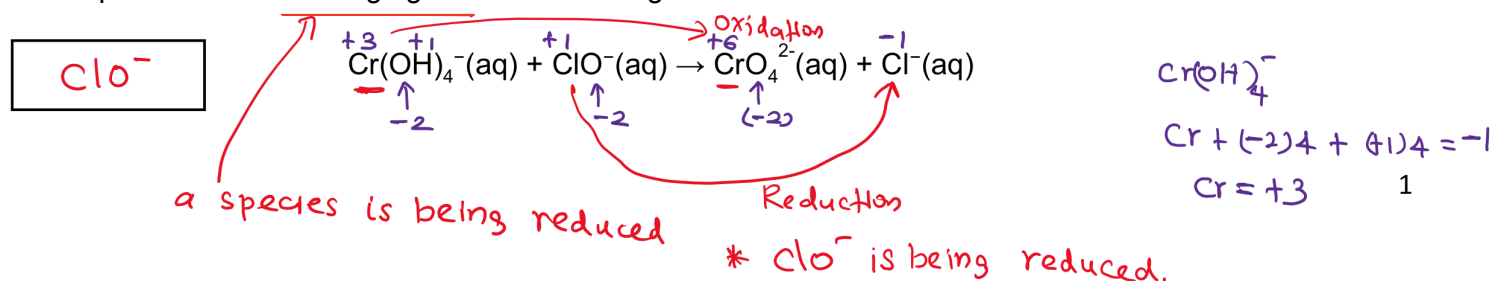


If solution (1) is a saturated solution of CaCO_3 , which of solutions (1)-(4) represents the solution after a small amount of HNO_3 is added and equilibrium is restored?

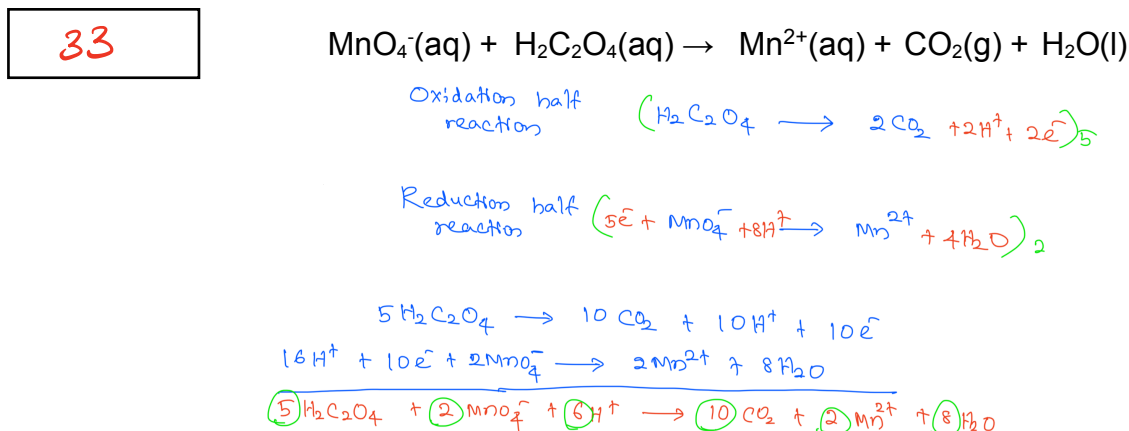
- ☐ A. (1)
☒ B. (2)
☐ C. (3)
☐ D. (4)



67. What species is the oxidizing agent in the following redox reaction?



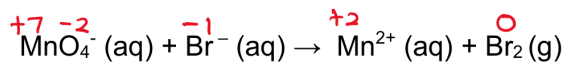
68. Balance the oxidation-reduction reaction in acidic solution. What is the sum of the coefficients?



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69. What is the coefficient of the permanganate ion in the balanced equation between bromide and permanganate ions in acidic solution?

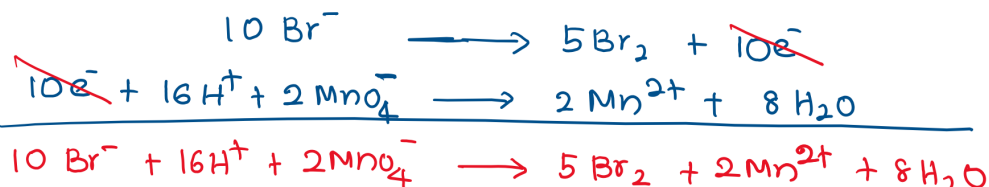
2



Oxidation half rxn



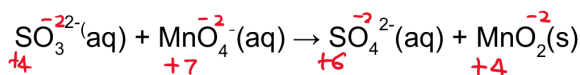
Reduction half rxn



70. Balance the following redox reaction if it occurs in basic solution. What are the coefficients in front of SO_3^{2-} and MnO_4^- in the balanced reaction?

SO_3^{2-}

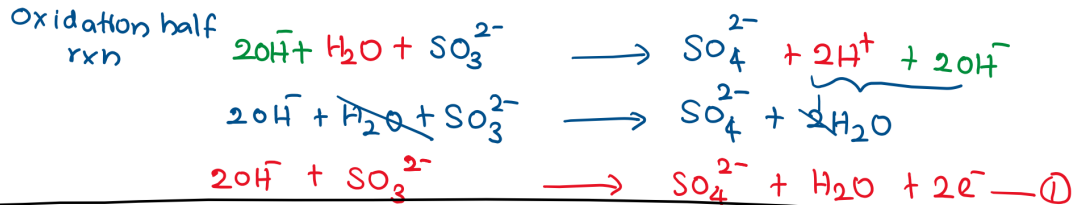
3



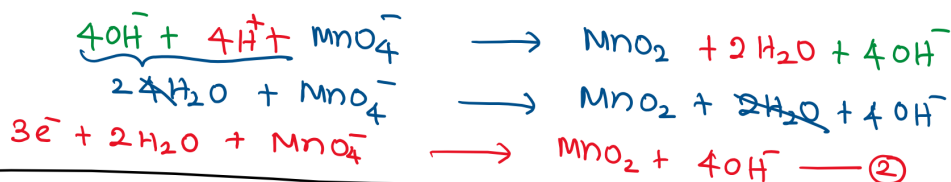
MnO_4^-

2

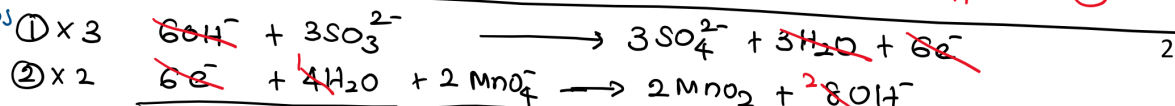
Oxidation half rxn



Reduction half rxn



Balancing e⁻s

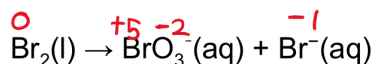


Adding together

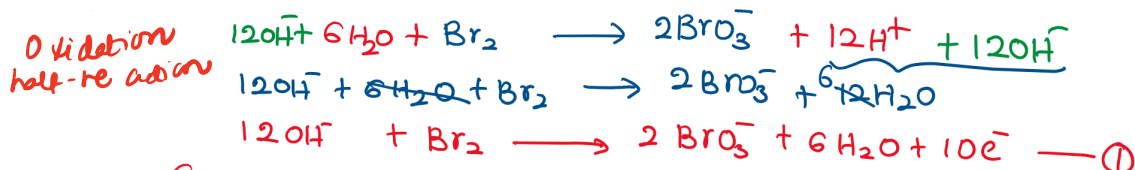


71. Balance the following redox reaction if it occurs in basic solution. What is the sum of the coefficients in the balanced reaction?

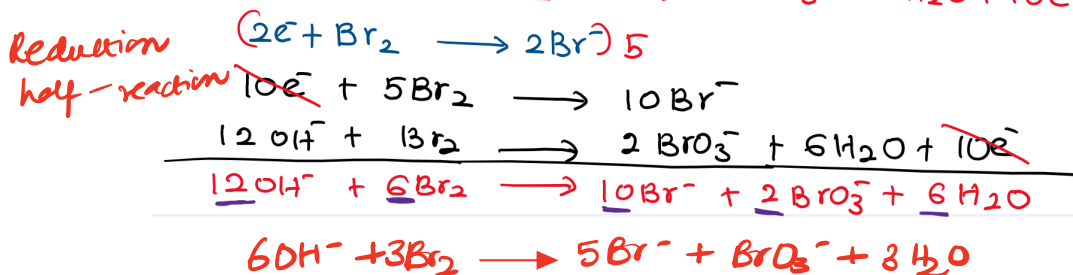
18



Oxidation half-reaction



Reduction half-reaction

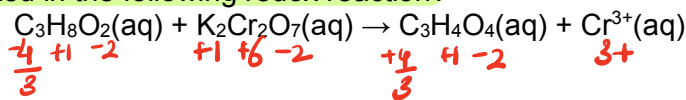


→ Oxidation state increases

72. What element is being oxidized in the following redox reaction?

C

- A. C
B. H
C. O
D. K



73. Which statement is true about redox reactions?

D

- A. A half-reaction can occur by itself. *the oxidation half reaction has to occur simultaneously with the reduction*
B. A redox reaction in base can include excess H^+ after it has been balanced. *balanced by OH^-*
C. Two oxidations can occur instead of one oxidation and one reduction.
D. At least 2 atoms must have their oxidation states change during a redox reaction.
E. None of these statements are true.

74. A homeowner in Boston becomes concerned that there may be appreciable amounts of lead in her drinking water due to 150-year old water pipes in her house. Consequently, she takes a sample of her drinking water in to be analyzed. The laboratory technician, who is new on the job, has been told to analyze by precipitating the lead ion out as lead iodide ($K_{\text{sp}} = 7.1 \times 10^{-9}$ for $\text{PbI}_2(\text{s})$) by slowly adding small portions of 1.00 M $\text{NaI}(\text{aq})$ solution.

If we assume that the concentration of lead in the solution is 1.00 mg/liter or approximately 4.8×10^{-6} M (this would be 1.0 part per million) is it possible to detect the lead in the drinking water by adding a total of no more than 10.0 mL of $\text{NaI}(\text{aq})$ solution to a 100 mL sample of drinking water? What is Q_{sp} ?

Yes

$$Q = [\text{Pb}^{2+}][\text{I}^-]^2 = \left(4.8 \times 10^{-6} \times \frac{100 \text{ mL}}{(100+10) \text{ mL}}\right) \times \left(1.00 \text{ M} \times \frac{10 \text{ mL}}{(100+10) \text{ mL}}\right)^2$$

$$= 3.61 \times 10^{-8} > K_{\text{sp}}$$

$$K_a = \frac{K_w}{K_b} = \frac{10^{-14}}{1.8 \times 10^{-5}} = 5.56 \times 10^{-10} \quad \text{p}K_a = -\log K_a = 9.255$$

Recitation Worksheet 12 – Exam 4 Review

75. When 100 mL each of 2.0×10^{-6} M $\text{Ag}^+(\text{aq})$ and 2.0×10^{-3} M $\text{Br}^-(\text{aq})$ are mixed, what is the remaining $\text{Ag}^+(\text{aq})$ ion concentration in the solution and is precipitation complete? The solubility product constant of $\text{AgBr}(\text{s})$ is 5.0×10^{-13} .

B

A. 7.1×10^{-7} M, no

B. 5.0×10^{-10} M, yes

C. 1.0×10^{-3} M, no

D. 5.0×10^{-13} M, yes

E. 2.5×10^{-10} M, yes

Method 1:



$$Q = [\text{Ag}^+][\text{Br}^-] = 2.0 \times 10^{-6} \times 2.0 \times 10^{-3} = 4.0 \times 10^{-9} \gg K_{sp}$$

so, precipitation is formed.

$$[\text{Br}^-] = 2.0 \times 10^{-3} \gg [\text{Ag}^+] = 2.0 \times 10^{-6}, \text{ and } K_{sp} \text{ is so small}$$

thus, $[\text{Br}^-]$ almost does not change in the precipitation.

Use original $[\text{Br}^-]$ in Q calculation. Remember, volume doubled, conc halved

$$Q = K_{sp} = [\text{Ag}^+][\text{Br}^-] = [\text{Ag}^+] \times (2.0 \times 10^{-3} \times \frac{1}{2}) = 5.0 \times 10^{-13}$$

$$[\text{Ag}^+] = 5.0 \times 10^{-10}$$

Method 2:

$Q > K_{sp}$, reaction shifts \leftarrow .



I $\frac{1}{2} 2.0 \times 10^{-6}$ $\frac{1}{2} 2.0 \times 10^{-3}$

C $-x$ $-x$

E $(1.0 \times 10^{-6} - x)$ $(1.0 \times 10^{-3} - x)$

$$K_{sp} = (1.0 \times 10^{-6} - x)(1.0 \times 10^{-3} - x) = 5.0 \times 10^{-13}$$

$$\Rightarrow x_1 = 1.0 \times 10^{-6}$$

$$x_2 = 9.99499 \times 10^{-7}$$

$$[\text{Ag}^+]_{eq} = 1.0 \times 10^{-6} - x = 1.0 \times 10^{-6} - 9.99499 \times 10^{-7} = 5.01 \times 10^{-10}$$

76. The solubility product constant of calcium chlorate ($\text{Ca}(\text{ClO}_3)_2$) of water is 7.1×10^{-7} at 25°C . How many grams of $\text{Ca}(\text{ClO}_3)_2$ is dissolved in 750 mL of saturated solution?

0.872

g



$$K_{sp} = [\text{Ca}^{2+}][\text{ClO}_3^-]^2$$

$$K_{sp} = S(2S)^2 = 4S^3$$

$$7.1 \times 10^{-7} = 4S^3 \rightarrow S = 5.620 \times 10^{-3} \text{ mol/L}$$

grams

$$\frac{5.620 \times 10^{-3} \text{ mol}}{\text{L}} \times \frac{206.98 \text{ g}}{1 \text{ mol}} = 1.163 \text{ g/L}$$

$$S \text{ is } 750 \text{ mL} = 1.163 \text{ g/L} \times 750 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.872 \text{ g}$$

77. Concentrated aqueous solutions of copper (I) nitrate, sodium sulfate and silver acetate are mixed together. The precipitate which forms is _____.

C

A. copper (I) acetate

B. silver nitrate

C. silver sulfate

D. sodium acetate

E. nothing precipitates



Most of SO_4^{2-} salts are soluble, except

Ca^{2+} , Sr^{2+} , Ba^{2+} , Ag^+ and Pb^{2+}