

Recitation Worksheet Thirteen : Exam 4 Review

Name:

Key

MyID:

Textbook:

Chemistry & Chemical Reactivity

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Instructions:

- This recitation worksheet is a review for Exam One.
- Exam coverage: 16.4-16.5, 17.1-17.4
- You **do not** need to submit it to Gradescope.
- The answer key has been posted with this worksheet to eLC.
- The **recitation session during the exam week (November 18th– 21st) is still mandatory**. Your attendance will be recorded.
- A periodic table and formula sheet are attached to the end of this worksheet.

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 stabilization
- E. All the above**

2. Which of the following acidity relationships is true?

B

- A. $H_2SO_3 > H_2SO_4$ *H_2SO_3 has less no. of oxygen compared to H_2SO_4*
- B. $H_2PO_4^- > HPO_4^{2-}$** *a negatively charged molecule is basic compared to a neutral molecule*
- C. $HF > HClO_4$
- D. $H_2CO_3 > HNO_3$
- E. None of these

C HF (weak acid) & $HClO_4$ is a strong acid

D

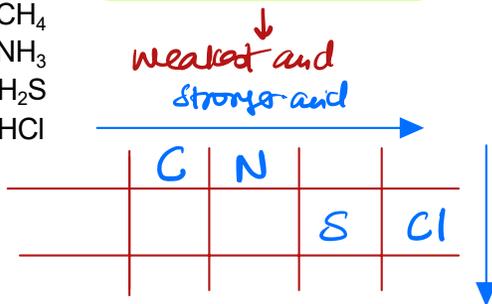
C	N	O
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 nitrogen is more electronegative compared to C \therefore HNO_3 is a stronger acid than H_2CO_3

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

A

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



1) $S < Cl$ in EN $\therefore H_2S$ is a weaker acid
 Also larger in size compared to Cl & N

2) $C < N$ in EN $\therefore CH_4$ is a weaker acid

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 \rightarrow neutral molecules
 B & E \rightarrow negatively charged } neutral molecules are stronger acids than negatively charged molecules

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

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

ACD

- A. HI or HCl *I is larger in size than HCl*
- B. HF or HBr
- C. HClO₃ or HBrO₃ *Cl is more electronegative than Br*
- D. HOSO₂CF₃ or HOSO₂CH₃

↓
Fluorine is electronegative which produces an inductive effect making HOSO₂CF₃ more acidic

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 binary acids (size is more important). In this case, S > O in size & S > P in electronegativity

∴ CH₃SH is the strongest

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

H₃PO₄

- A. H₂SiO₃ or H₃PO₄ *∴ more oxygen*

H₂CO₃

- B. H₂CO₃ or H₂BO₃ *Carbon is more electronegative than Boron*

HOClO

- C. HOClO or HOBr *Cl is more electronegative than Br & HOClO has more oxygen*

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

III < I < II

- I. H₂Te
- II. HI
- III. H₂S

		O
	S	
	Te	I

H₂S < H₂Te < HI

3

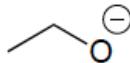
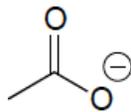
↗ From weaker to strongest acid

① I is more electronegative than Te & larger than Te ∴ HI is the strongest acid

② Te is larger than S ∴ H₂Te is a stronger acid than H₂S

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

E



A strong base will have a strong affinity to a proton

Between the three molecules the ones that are negatively charged will have a stronger affinity to the proton (I & II) than a neutral molecule (III)

I. II.

III.

Can stabilize the negative charge via resonance
∴ a weaker base than II

II > I > III

- A. I > II > III
- B. II > III > I
- C. III > II > I
- D. I > III > II
- E. II > I > III**

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

D

- ~~A.~~ HCl
- ~~B.~~ H₃PO₄
- ~~C.~~ H₂PO₄⁻
- D.** HPO₄²⁻
- ~~E.~~ HNO₃

A & E are strong acids

B, C, E are weaker acids. As the protons become negatively charged, they become more basic (more difficult to lose a proton). The higher the negative charge the more basic the compound becomes

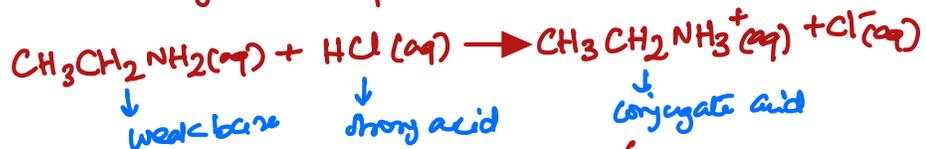


11. A solution is prepared by dissolving 0.32 mol of CH₃CH₂NH₃Cl in 1.00 L of 1.5 M CH₃CH₂NH₂. If 10. mL of 0.11 M HCl is added to this solution, the pH of the solution will slightly _____ because the HCl reacts with the _____ present in the solution.

C

- ~~A.~~ Increase, CH₃CH₂NH₃⁺
- ~~B.~~ Increase, CH₃CH₂NH₂
- C.** Decrease, CH₃CH₂NH₂ ^{base}
- ~~D.~~ Decrease, CH₃CH₂NH₃⁺

If a small amount of a strong acid like HCl is added it will be neutralized by the weak base according to the equation below:

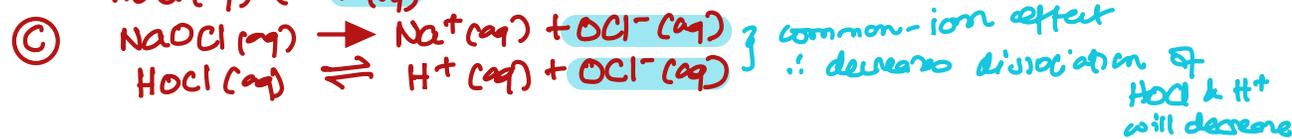
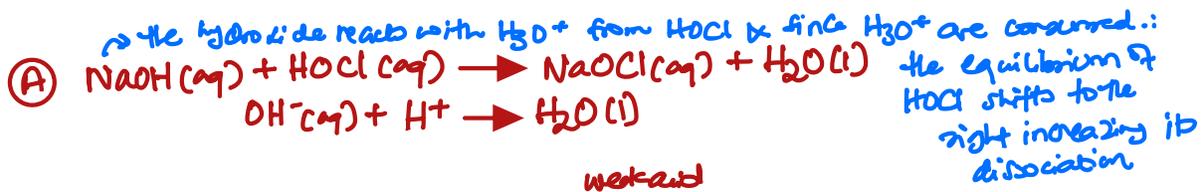


more conjugate acid is produced
∴ pH will decrease

12. Identify the **false** statement regarding a solution that contains 0.20 moles of hypochlorous acid.

C

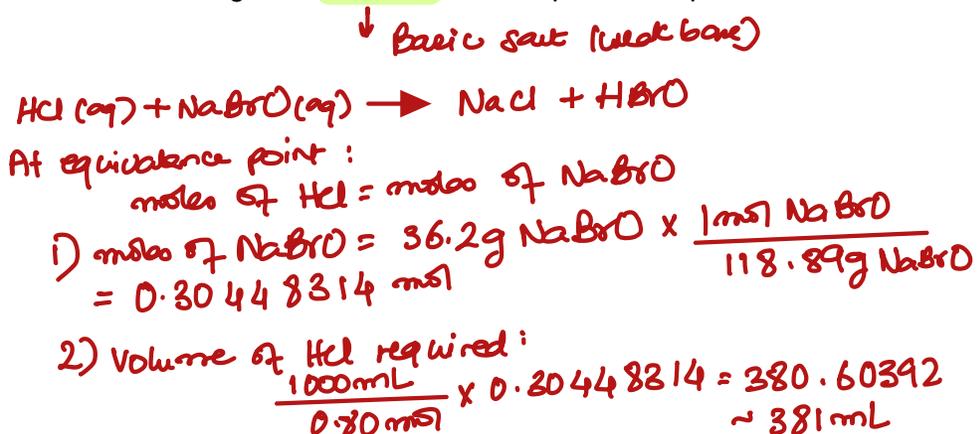
- A. Adding NaOH ^{strong base} will increase the dissociation of HOCl, and decrease $[H^+]$
- B. Adding HCl will decrease the dissociation of HOCl and decrease $[OCl^-]$
- C. Adding NaOCl will increase the dissociation of the original HOCl, and will increase $[H^+]$
- D. Adding NaCl will not affect either the dissociation of the original HOCl or the solution pH



13. 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



14. 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.

Calculate the moles of HF and the concentration of HF after addition.

Buffer

↓ strong acid ∴ neutralized by basic salt

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.



Initial	0.0080	0.048	0	0.048
Change	-0.0080	-0.0080	+0.0080	+0.0080
Final	0	0.040	0.0080	0.056

② $[\text{HF}] = \frac{0.056 \text{ mol}}{(60.0 + 20.0) \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 0.70 \text{ M}$

15. 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 HCl(aq) to give 1.0 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 $\text{p}K_b$ is 5.92?

C

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

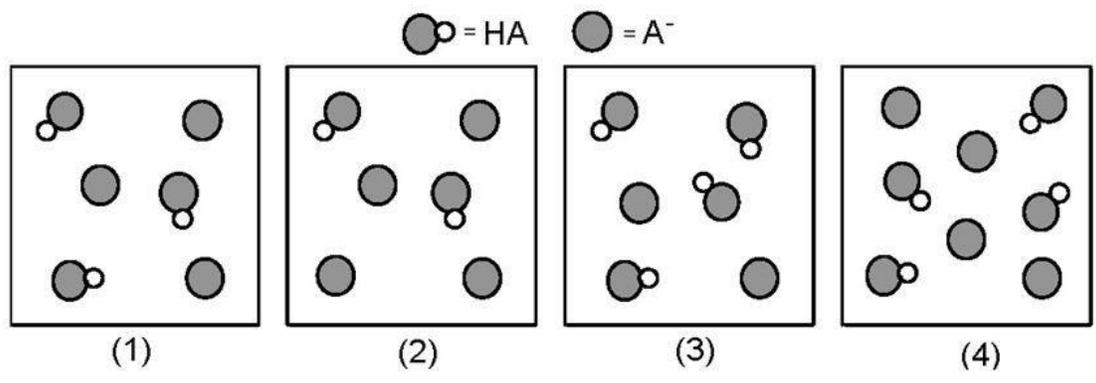
$$\begin{aligned} \text{pOH} &= \text{p}K_b + \log \frac{[\text{TRISH}^+]}{[\text{TRIS}]} \\ \therefore \text{pOH} &= 5.92 + \log \frac{[0.60]}{[0.30]} \\ &= 6.221029996 \\ \text{pH} &= 14.00 - 6.22 = 7.78 \end{aligned}$$

16. All of the following solutions would be considered buffers **except** → weak acid to its conjugate base salt or weak base to its conjugate acid

B

- A. CH₃COOH/CH₃COO⁻.
- B.** NaCl/HCl. → strong acid to its conjugate salt ∴ cannot be a buffer
- C. H₂PO₄⁻/HPO₄²⁻.
- D. HPO₄²⁻/PO₄³⁻.

17. These pictures represent solutions that contain a weak acid HA (pK_a = 5.0) and its sodium salt NaA. Buffers
 Unshaded spheres represent H atoms and shaded spheres represent A⁻ ions. (Na⁺, H₃O⁺, OH⁻, and solvent H₂O molecules have been omitted for clarity.)



$$\frac{[A^-]}{[HA]}$$

$$\frac{3}{3} = 1$$

$$\frac{4}{2} = 2$$

$$\frac{2}{4} = \frac{1}{2}$$

$$\frac{4}{4} = 1$$

Which solution has the greatest buffer capacity? → Greatest buffer capacity is when the ratio of [A⁻]/[HA] ~ 1

D

- E. (1)
- F. (2)
- G. (3)
- H. (4)

pH of a weak base

18. 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

↓
acidic salt produced
∴ weak base & a strong acid titration

19. 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%

HA



Initial	y	-	0	0
Change	-2.5×10^{-9}	-	$+2.5 \times 10^{-9}$	$+2.5 \times 10^{-9}$
Equilibrium	$y - 2.5 \times 10^{-9}$	-	2.5×10^{-9}	2.5×10^{-9}

→ $K_a = 10^{-9.60} = 2.51188643 \times 10^{-10}$

③ % dissociation
 $= \frac{2.5 \times 10^{-9}}{2.8 \times 10^{-8}} \times 100$
 $= 9.090909085$
 $\approx 9\%$

② Find the initial conc. of glycine

$2.5 \times 10^{-10} = \frac{[2.5 \times 10^{-9}]^2}{[y - 2.5 \times 10^{-9}]}$

↪ $10^{-8.60} = 2.51188643 \times 10^{-10}$

∴ $y = 2.763075047 \times 10^{-8} \approx 2.8 \times 10^{-8} M$

20. 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



Initial	0.0025	0.0015	0	-
Change	-0.0015	-0.0015	$+0.0015$	-
Final	0.0010	0	0.0015	-

↓
strong acid

↓
Neutral salt

② New molarity of HCl = $\frac{0.0010 \text{ mol}}{(25 + 15) \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}}$
 $= 0.025 M$

③ $pH = -\log [0.025] = 1.60$

21. In a titration experiment, it was determined that a 50.0 mL sample of HNO_3 required 66.0 mL of 0.80 M NaOH to reach the equivalence point. What was the molarity of the HNO_3 ?

C

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

At equivalence point:

$$\text{moles of } \text{HNO}_3 = \text{moles of NaOH}$$

1) moles of NaOH:

$$\frac{0.80 \text{ mol}}{1000 \text{ mL}} \times 66.0 \text{ mL} = 0.0528 \text{ mol}$$

2) Molarity of HNO_3 :

$$\frac{0.0528 \text{ mol}}{50.0 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 1.056 \approx 1.06 \text{ M}$$

→ Acid

22. 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



Initial	0.50	-	0	0.40
Change	-x	-	+x	+x
Equilibrium	0.50-x	-	x	0.40+x

② $K_a = \frac{[\text{N}_3^-][\text{H}_3\text{O}^+]}{[\text{HN}_3]} \therefore 2.0 \times 10^{-5} = \frac{[x][0.40+x]}{[0.50-x]}$
 $\therefore x = 2.5 \times 10^{-5} = \text{N}_3^-$ (C) >> K < 100

<7

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

A

- A. $\text{HClO}_4 + \text{NaF}$ strong acid + weak base \rightarrow basic salt (<7) (or basic salt)
- B. $\text{HNO}_3 + \text{KOH}$ strong acid + strong base (=7)
- C. $\text{NH}_4\text{Cl} + \text{NaOH}$ acidic salt + strong base (>7)
- D. $\text{HF} + \text{NaOH}$ weak acid + strong base (>7)
- E. None of these are correct

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

D

- A. $\overset{\text{strong acid}}{\text{HBr(aq)}} + \overset{\text{strong base}}{\text{NaOH(aq)}} \rightleftharpoons \text{H}_2\text{O(l)} + \text{NaBr(aq)}$
 B. $\text{H}_3\text{O}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightleftharpoons 2 \text{H}_2\text{O(l)}$ \rightarrow net ionic of strong acid & strong base
 C. $\text{HF(aq)} + \text{LiOH(aq)} \rightleftharpoons \text{H}_2\text{O(l)} + \text{LiF(aq)}$ total ionic
 D. $\text{HF(aq)} + \text{OH}^-(\text{aq}) \rightleftharpoons \text{H}_2\text{O(l)} + \text{F}^-(\text{aq})$

25. 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.

Buffer

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

$$\text{pH} = \text{pK}_a + \log \frac{[\text{CN}^-]}{[\text{HCN}]}$$

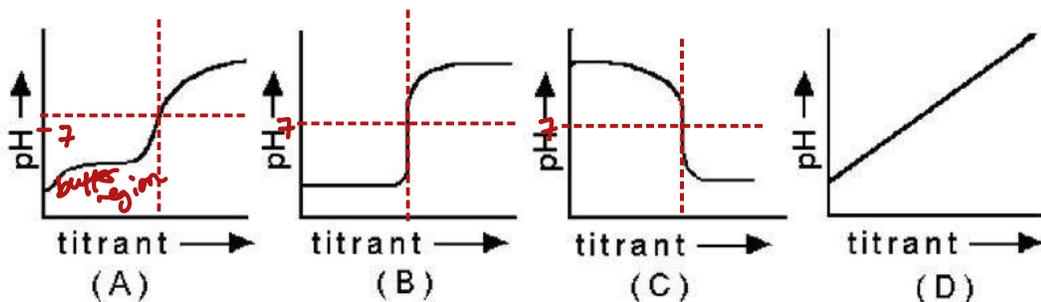
$$[\text{HCN}] = \frac{0.10 \text{ mol}}{1000 \text{ mL}} \times 50.00 \text{ mL} = \frac{0.0050 \text{ mol}}{100.00 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 0.050 \text{ M}$$

$$[\text{CN}^-] = \frac{0.010 \text{ mol}}{1000 \text{ mL}} \times 50.00 \text{ mL} = \frac{0.00050}{100.00 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 0.0050 \text{ M}$$

$$\therefore \text{pH} = -\log(4.9 \times 10^{-10}) + \log \frac{[0.0050]}{[0.050]}$$

$$\text{pH} = 8.30980392$$

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



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?

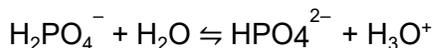
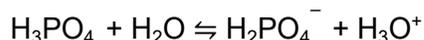
C

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

A

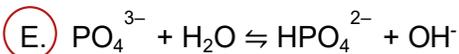
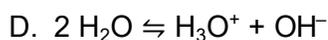
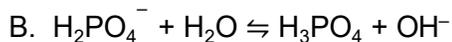
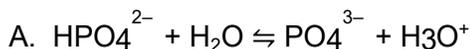
27. Phosphoric acid is a triprotic acid, ionizing in sequential steps:

E



Handwritten notes:
 K_{a1} comes from NaOH \therefore weak acid \therefore does not affect pH (pH neutral)
 K_{a2}
 K_{a3}
 Na_3PO_4 (salt)
 PO_4^{3-} is the conjugate base of HPO_4^{2-} (weak acid) \therefore undergoes hydrolysis in H_2O

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



Handwritten: how to determine K_b for this reaction?

Answer: Using the K_a of HPO_4^{2-} (K_{a2})

$K_w = K_{a2} \times K_b \quad \therefore K_b(\text{PO}_4^{3-}) = \frac{K_w}{K_{a2}(\text{HPO}_4^{2-})}$

Butter

28. What is the pH of a solution prepared by mixing 50.00 mL of 0.10 M NH₃ with 20.00 mL of 0.10 M NH₄Cl?

K_b = 1.8 × 10⁻⁵ for NH₃.

9.65

$$[\text{NH}_3] = \frac{50.00 \text{ mL} \times 0.10 \text{ mol/L}}{(50.00 + 20.00) \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 0.07142857143 \text{ M}$$

$$[\text{NH}_4\text{Cl}] = \frac{20.00 \text{ mL} \times 0.10 \text{ mol/L}}{(50.00 + 20.00) \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 0.02857142857 \text{ M}$$

$$\text{pOH} = \text{pK}_b + \log \frac{[\text{NH}_4^+]}{[\text{NH}_3]} = -\log(1.8 \times 10^{-5}) + \log \frac{[0.029]}{[0.071]}$$

$$= 4.346787486$$

$$\therefore \text{pH} = 14.00 - 4.35 = 9.65$$

29. 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	pKa
←	HA	4.00
	HB	7.00
	HC	10.00
←	HD	11.00

weakest CB (green) strongest acid (blue) ← HA

strongest CB (green) weakest acid (blue) ← HD

○ the lower the pKa the stronger the acid ∴ produces a weak conjugate base

○ the higher the pKa the weaker the acid ∴ produces a strong conjugate base

A. NaA

B. NaB

C. NaC

D. NaD

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

↓ produces the most basic salt

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

D

- A. NaNO₃, NH₄CN, NaOOCCH₃, NH₄Cl
neutral depends basic acidic
- B. NaHCO₃, NaF, NH₄Cl, Na₂SO₃
depends basic acidic basic
- C. Na₂CO₃, KCl, NaOOCCH₃, NH₄Cl
basic neutral basic acidic
- D. Na₂CO₃, NaF, NaOOCCH₃, NaCN
basic basic basic basic

E. All of the above.

31. Which one of these is correct?

A

- (A) K_2SO_3 is a stronger base than $KHSO_3$. *K⁺ is pH neutral & SO_3^{2-} is the conjugate base of HSO_3^-*
- B. Na_2HPO_4 is a weaker base than NaH_2PO_4 . *HPO_4^{2-} is the conjugate base of $H_2PO_4^-$*
- C. K_2CO_3 is a weaker base than $KHCO_3$. *CO_3^{2-} is the conjugate base of HCO_3^- (acid)*
- D. $NaHSO_3$ is a stronger acid than $NaHSO_4$. *H_2SO_4 is a stronger acid because has more oxygens than H_2SO_3*
- E. All of these statements are correct.

32. 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? \rightarrow weak acid

4.4×10^{-7}

$M_1 V_1 = M_2 V_2$
 $M_1 \times 25.0 \text{ mL} = 0.104 \times 21.7$
 $\therefore M_1 = 9.0272 \times 10^{-2} \text{ M}$

$HA(aq) + H_2O(l) \rightleftharpoons A^-(aq) + H_3O^+(aq)$

Initial	9.03×10^{-2}	-	0	0
Change	-2.0×10^{-4}	-	2.0×10^{-4}	2.0×10^{-4}
Equilibrium	9.01×10^{-2}	-	2.0×10^{-4}	2.0×10^{-4}

data mined from pH
 $\therefore [H_3O^+]_{eq} = 10^{-3.70}$
 $= 1.995262315 \times 10^{-4} \text{ M}$

$\therefore K_a = \frac{[A^-][H_3O^+]}{[HA]} = \frac{[2.0 \times 10^{-4}][2.0 \times 10^{-4}]}{[9.01 \times 10^{-2}]}$
 $= 4.419853857 \times 10^{-7} \approx 4.4 \times 10^{-7}$

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

C

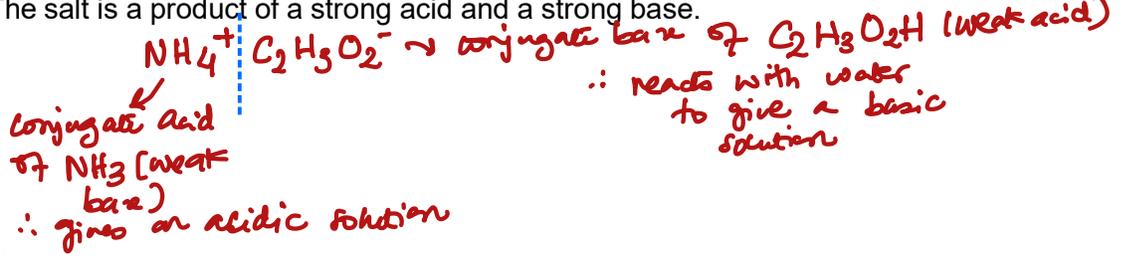
- A. 0.2 M HNO_3 and 0.2 M $NaNO_3$. *strong acid salt of conjugate base \therefore not a buffer*
- B. 0.2 M HNO_3 and 0.4 M HF. *strong acid weak acid \therefore not a buffer*
- C. 0.2 M HNO_3 and 0.4 M NaF. *strong acid basic salt (weak base) \therefore Buffer*
- D. 0.2 M HNO_3 and 0.4 M NaOH. *strong acid strong base \therefore not a buffer*
- limiting*
 $HNO_3(aq) + NaOH(aq) \rightarrow NaNO_3(aq) + H_2O(l)$
- | | | | | |
|----------|------|------|------|-----|
| Limiting | 0.2 | 0.4 | 0 | 0 |
| F | -0.2 | -0.2 | +0.2 | 0.2 |
| I | 0 | 0.2 | 0.2 | 0.2 |
| C | 0 | 0 | 0 | 0 |
- basic salt conjugate acid neutral salt*

34. 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. *it is reactive (look at the explanation below)*
- B. Ammonium acetate is a weak electrolyte. *$\text{NH}_4\text{C}_2\text{H}_3\text{O}_2$ is a strong electrolyte*
- C. All salts of weak acids and weak bases are neutral. *they are reactive*
- 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. *→ salt is a product of weak acid & weak base*

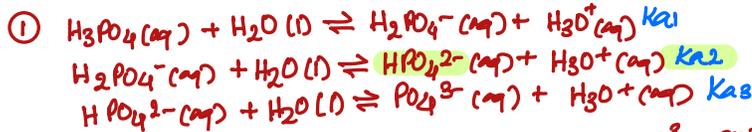
If both cation & anion are reactive (salt of a weak base & weak acid) ∴ $K_a \approx K_b$ should be compared if $K_a \approx K_b$ ∴ the salt is neutral



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

6.3×10^{-8}

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



② *If the difference between K_{a1} & K_{a2} is greater than a factor of 10^8 ∴ most H_2O comes from the first step & according to Le Chatelier's principle the dissociation of the second step is suppressed*
 $\therefore [\text{H}_2\text{PO}_4^-] \approx [\text{H}_3\text{O}^+] \therefore K_{a2} = \frac{[\text{HPO}_4^{2-}][\text{H}_3\text{O}^+]}{[\text{H}_2\text{PO}_4^-]}$
 $\therefore K_{a2} = \frac{[\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]}$

36. 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?

4.66

Strong acid ∴ neutralized by the weak base in the buffer
 $\text{NaB} =$ sodium benzoate
 $\text{BH} =$ benzoic acid



Initial	0.036	0.00400	0	0.0072
Change	-0.00400	-0.00400	+0.00400	+0.00400
Final	0.032	0	0.00400	0.0112

* Moles of NaB in 72.0 mL solution:
 $\frac{0.50 \text{ mol}}{1000 \text{ mL}} \times 72.0 \text{ mL} = 0.036 \text{ mol}$

* Moles of BH in 72.0 mL solution:
 $\frac{0.10 \text{ mol}}{1000 \text{ mL}} \times 72.0 \text{ mL} = 0.0072$

* Moles of HI in 2.00 mL solution:
 $\frac{2.00 \text{ mol}}{1000 \text{ mL}} \times 2.00 \text{ mL} = 0.00400$

② $\text{pH} = \text{p}K_a + \log \frac{[\text{NaB}]}{[\text{BH}]}$
 $= -\log(6.3 \times 10^{-5}) + \log \frac{[0.032]}{[0.0112]}$
 $= 4.656591406$
 ≈ 4.66

the pH of the buffer is lower due to the addition of the strong acid. when the strong acid is neutralized by the weak base ∴ more weak acid is produced & the pH of the buffer is lower.

37. 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} .

Weak acid

Strong base

4.24



Initial	0.018	0.0044	0	-
Change	-0.0044	-0.0044	+0.0044	-
Final	0.014	0	0.0044	-

Buffer

* Moles of acetic acid:

$$52.60 \text{ mL} \times \frac{0.35 \text{ mol}}{1000 \text{ mL}} = 0.01841 \text{ moles}$$

* Moles of NaOH:

$$22.08 \text{ mL} \times \frac{0.20 \text{ mol}}{1000 \text{ mL}} = 0.004416$$

② $\text{pH} = \text{p}K_a + \log \frac{[\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} = -\log(1.8 \times 10^{-5}) + \log \frac{[0.0044 \text{ mol}]}{[0.014 \text{ mol}]}$
 $= 4.24381469 \approx 4.24$

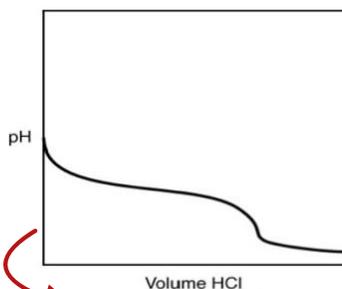
38. A 0.500 g sample of an unknown substance was titrated with a 0.1 M HCl solution. Another 0.500 g sample of it was titrated with a 0.1 M NaOH solution. The resulting titration curves are illustrated here. Given the following possibilities, what is the sample? *Strong base*

Strong acid

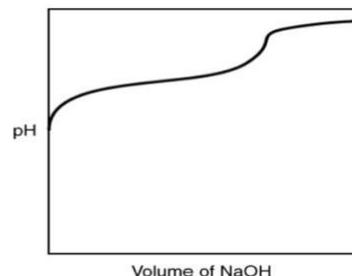
C

- A. $\text{Na}_2\text{CO}_3^{2-}$ *basic only*
- B. CO_2 *Acidic*
- C. NaHCO_3^-
- D. H_2CO_3 *Weak acid*
- E. There is no way to tell.

Na^+ ions from NaOH (strong base) \therefore neutral



Acts as a base



Acts as an acid like



Buffer

39. 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 0.800 M NaOH to this solution? (NH_4^+ $K_a = 5.6 \times 10^{-10}$)

Strong base

9.55



limiting

Initial	0.300	0.032	0.500	-
Change	-0.032	-0.032	+0.032	-
Final	0.268	0	0.532	-

Acid

conjugate base
Buffer

② $\text{pH} = \text{p}K_a + \log \frac{[\text{NH}_3]}{[\text{NH}_4^+]}$
 $= -\log(5.6 \times 10^{-10}) + \log \frac{[0.532]}{[0.268]}$
 $= 9.549588811 \approx 9.550$

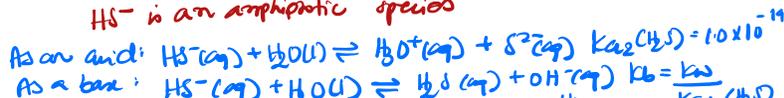
40. Of the following substances, which one(s) will form basic solutions (select all that apply).
 (H_2CO_3 $K_{a1} = 4.4 \times 10^{-7}$, $K_{a2} = 4.7 \times 10^{-11}$; H_2S $K_{a1} = 1.1 \times 10^{-7}$, $K_{a2} = 1.0 \times 10^{-19}$)

E

NaHS $\text{Cu}(\text{NO}_3)_2$ KHCO_3 NaF

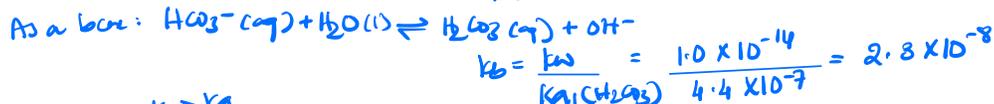
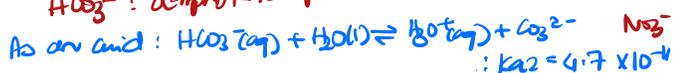
- A. NaHS, $\text{Cu}(\text{NO}_3)_2$
- B. KHCO_3 , NaHS
- C. NaF only
- D. NaF, KHCO_3
- E. NaHS, KHCO_3 and NaF**

(A) NaHS:
 Na^+ comes from NaOH (strong base) \therefore pH neutral
 HS^- is an amphoteric species



(B) KHCO_3^- : K^+ comes from KOH (strong base) \therefore K^+ is pH neutral

HCO_3^- : amphoteric species



$K_b > K_a$
 \therefore basic

$\text{Cu}(\text{NO}_3)_2$: Cu^{2+} a small highly charged metal cation \therefore forms an acidic solution

NO_3^- : conjugate base of HNO_3 (strong acid) \therefore pH neutral

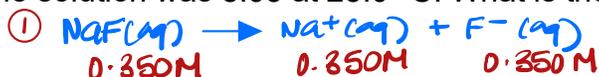
(C) NaF: Na^+ (pH neutral), F^- conjugate base of HF (weak acid) \therefore basic
 $\text{F}^-(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{HF}(\text{aq}) + \text{OH}^-(\text{aq})$

41. An aqueous solution of NaF is prepared by dissolving 0.350 mol of NaF in sufficient water to yield 1.0 L of solution. The pH of the solution was 8.93 at 25.0 °C. What is the K_b of F^- ?

2.1

× 10

-10



	$[\text{F}^-]$	$[\text{H}_2\text{O}]$	$[\text{HF}]$	$[\text{OH}^-]$
Initial	0.350	-	0	0
Change	-x	-	+x	+x
Equilibrium	0.350-x	-	x	x

(3) $[\text{OH}^-] = 10^{-\text{pOH}}$
 $= 10^{-5.07}$
 $= 8.511380382 \times 10^{-6} \text{ M}$
 $= x$

(2) Find $[\text{OH}^-]$

$\text{pH} + \text{pOH} = 14.00$

$\text{pOH} = 14.00 - 8.93 = 5.07$

$K_b = \frac{[\text{HF}][\text{OH}^-]}{[\text{F}^-]} = \frac{[8.5 \times 10^{-6}][8.5 \times 10^{-6}]}{[0.350]}$

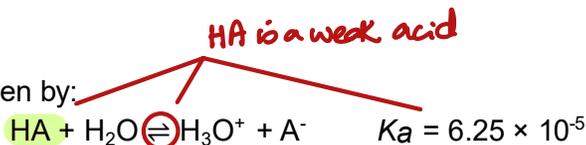
$= 2.06986737 \times 10^{-10}$
 $\sim 2.1 \times 10^{-10}$

42. 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^- ?

2.19×10^{-12}

$K_w = K_a \times K_b$
 $1.0 \times 10^{-14} = 4.57 \times 10^{-3} \times K_b$
 $\therefore K_b = 2.19 \times 10^{-12}$

43. 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

$K^+A^- \rightarrow$ potassium salt
 \therefore conjugate base of HA \therefore reactive
 From strong base (KOH) ① $A^-(aq) + H_2O(l) \rightleftharpoons HA^-(aq) + OH^-(aq)$

Initial	0.345	-	0	0
Change	-x	-	+x	+x
Equilibrium	0.345-x	-	x	x

② solve for K_b
 $K_w = K_a \times K_b \therefore K_b = \frac{K_w}{K_a}$
 $= \frac{1.0 \times 10^{-14}}{6.25 \times 10^{-5}} = 1.6 \times 10^{-10}$

③ $1.6 \times 10^{-10} = \frac{[x][x]}{[0.345-x]} \rightarrow \frac{C}{K} \gg 100$

$\therefore x^2 = 1.6 \times 10^{-10} [0.345] = 5.52 \times 10^{-11}$
 $\therefore x = \sqrt{5.52 \times 10^{-11}} = \pm 7.429670248 \times 10^{-6}$
 & discard the negative value
 $\therefore [OH^-] = x = 7.43 \times 10^{-6} M$
 $\therefore pOH = -\log [7.43 \times 10^{-6}] = 5.129$
 $\therefore pH = 14.00 - 5.129 = 8.871$

44. 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.

A

Basic

(1)

acidic

(2)

neutral

(3)

MA is a strong electrolyte
 $\therefore MA(aq) \rightarrow M^+(aq) + A^-(aq)$
6 spectator ions

$A^-(aq) + H_2O(l) \rightleftharpoons HA(aq) + OH^-(aq)$

MA is a strong electrolyte
 $\therefore MA(aq) \rightarrow M^+(aq) + A^-(aq)$
6 spectator ions

$M^+(aq) + H_2O(l) \rightleftharpoons MOH(aq) + H^+(aq)$
 $H^+(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq)$

 $M^+(aq) + 2H_2O(l) \rightleftharpoons MOH(aq) + H_3O^+(aq)$

$MA(aq) \rightarrow M^+(aq) + A^-(aq)$

Which picture represents a basic salt?

- A. (1) Only drawing with hydroxide ions present.
- B. (2)
- C. (3)
- D. none of these

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

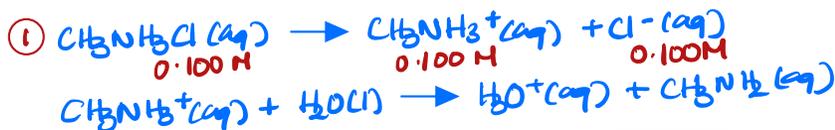
C

- A) $Sr(ClO_4)_2$ *neutral*
- B) KBr *neutral*
- C) NH_4I *acidic*
- D) K_2CO_3 *basic*
- E) $NaNO_3$ *neutral*

Acidic salt

46. 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



	$[\text{CH}_3\text{NH}_3^+]$	$[\text{H}_2\text{O}]$	$[\text{H}_3\text{O}^+]$	$[\text{CH}_3\text{NH}_2]$
Initial	0.100	-	0	0
Change	-x	-	+x	+x
Equilibrium	0.100-x	-	x	x

$$\textcircled{2} K_a = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{3.7 \times 10^{-4}} = 2.7 \times 10^{-11}$$

$$\textcircled{3} K_a = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{NH}_2]}{[\text{CH}_3\text{NH}_3^+]}$$

$$2.7 \times 10^{-11} = \frac{[x][x]}{[0.100-x]}$$

$\frac{0.100}{x} \gg 100$ Basic salt

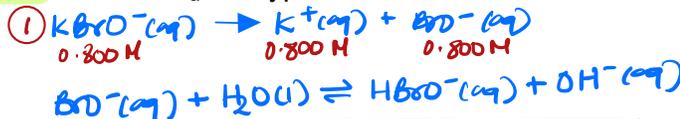
$$\textcircled{4} x^2 = 2.70 \times 10^{-12}$$

$$\therefore x = 1.64 \times 10^{-6}$$

$$\therefore \text{pH} = -\log [1.64 \times 10^{-6}] = 5.784$$

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

11.301



	$[\text{BrO}^-]$	$[\text{H}_2\text{O}]$	$[\text{HBrO}]$	$[\text{OH}^-]$
Initial	0.800	-	0	0
Change	-x	-	+x	+x
Equilibrium	0.800-x	-	x	x

$$\textcircled{2} \text{ Calculate } K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{2.0 \times 10^{-9}} = 5.0 \times 10^{-6}$$

$$\textcircled{3} K_b = \frac{[\text{HBrO}][\text{OH}^-]}{[\text{BrO}^-]}$$

$$5.0 \times 10^{-6} = \frac{[x][x]}{[0.800-x]}$$

$\frac{0.800}{5.0 \times 10^{-6}} \gg 10$

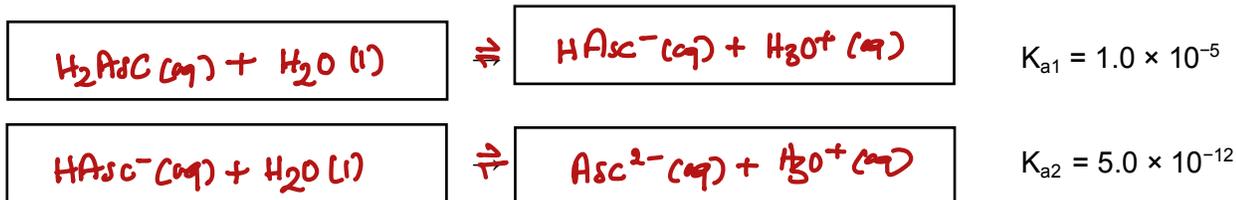
$$x^2 = 4.00 \times 10^{-6}$$

$$\therefore x = 2.00 \times 10^{-3}$$

$$\textcircled{4} \text{pOH} = -\log [2.00 \times 10^{-3}] = 2.669$$

$$\therefore \text{pH} = 14.000 - 2.669 = 11.301$$

48. Ascorbic acid, $H_2C_6H_6O_6$ 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.



49. 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 for $HNO_2 = 4.0 \times 10^{-4}$

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

H_3PO_4 :

$K_{a1} = 7.5 \times 10^{-3}$

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

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

C

a. $NaCl(aq)$
From NaOH (strong base) + HCl (strong acid) → pH = 7 (neutral)
conjugate base of HCl (strong acid) → pH = 7 (neutral)

A

b. $NH_4NO_3(aq)$
conjugate acid of NH_3 (weak base) ∴ pH < 7
conjugate base of HNO_2 (strong acid) ∴ pH = 7 (neutral)
The salt is overall acidic

B

c. $NaCH_3CO_2(aq)$
conjugate base of CH_3CO_2H (weak acid) ∴ overall basic salt
pH = 7 (neutral)

A

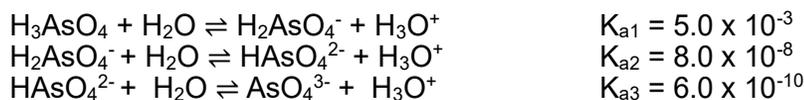
d. $NH_4NO_2(aq)$
conjugate acid of NH_3 (weak base) ∴ pH < 7
conjugate base of HNO_2 (weak acid) ∴ pH > 7
to determine whether the solution will be acidic, basic or neutral compare K_a of NH_4^+ & K_b of NO_2^-

$K_a(NH_4^+) = \frac{K_w}{K_b(NH_3)} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$
 $K_a > K_b$ ∴ the solution is acidic

B

e. $(NH_4)_3PO_4(aq)$
conjugate acid of NH_3 (weak base) ∴ pH < 7
 $K_a = 5.6 \times 10^{-10}$
 $PO_4^{3-}(aq) + H_2O(l) \rightleftharpoons HPO_4^{2-}(aq) + OH^-(aq)$
conjugate base of $H_2PO_4^-$ (weak acid)
 $\therefore K_b = \frac{K_w}{K_{a3}} = \frac{1.0 \times 10^{-14}}{4.8 \times 10^{-13}} = 2.1 \times 10^{-2}$
 $K_b > K_a$ ∴ solution is basic

50. 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.

$$[\text{H}_3\text{AsO}_4] = \boxed{4.84} \text{ M}$$

$$[\text{H}_2\text{AsO}_4^-] = \boxed{0.158} \text{ M}$$

$$[\text{H}_3\text{O}^+] = \boxed{0.158} \text{ M}$$



Initial	5.00	-	0	0
Change	-x	-	+x	+x
Equilibrium	5.00-x	-	x	x

$$\begin{aligned} \therefore K_{a1} &= \frac{[x][x]}{[5.00-x]} \quad \therefore 5.0 \times 10^{-3} = \frac{x^2}{[5.00-x]} \quad \frac{C}{F} \gg 100 \\ \therefore x &= \sqrt{5.0 \times 10^{-3} [5.00]} = 0.158 = [\text{H}_3\text{O}^+] = [\text{H}_2\text{AsO}_4^-] \\ \therefore [\text{H}_3\text{AsO}_4]_{\text{eq}} &= 5.00 - 0.158 = 4.84 \end{aligned}$$

b. What is the pH of this solution?

$$\text{pH} = \boxed{0.801}$$

If the difference between K_{a1} & K_{a2} is $> 10^2$ \therefore pH is calculated using the first step = $\text{pH} = -\log [0.158] = 0.801$



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

$$[\text{HASO}_4^{2-}] = \boxed{8.0 \times 10^{-8}} \text{ M}$$

Initial	0.158	-	0	0.158
Change	-y	-	+y	+y
Equilibrium	0.158-y	-	y	0.158+y

$$K_{a2} = \frac{[y][0.158+y]}{[0.158-y]} \quad \frac{C}{F} \gg 100$$

$$\begin{aligned} \text{Assuming } [\text{H}_2\text{AsO}_4^-] &\approx [\text{H}_3\text{O}^+] \\ \therefore K_{a2} &= \frac{[\text{HASO}_4^{2-}][\text{H}_3\text{O}^+]}{[\text{H}_2\text{AsO}_4^-]} = 8.0 \times 10^{-8} \end{aligned}$$

$$\begin{aligned} \therefore 8.0 \times 10^{-8} &= \frac{y[0.158]}{[0.158]} \\ \therefore y &= 8.0 \times 10^{-8} \end{aligned}$$

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

$$[\text{AsO}_4^{3-}] = \boxed{8.04 \times 10^{-16}} \text{ M}$$



Initial	8.0×10^{-8}	-	0	0.158
Change	-z	-	+z	+z
Equilibrium	$8.0 \times 10^{-8} - z$	-	z	0.158+z

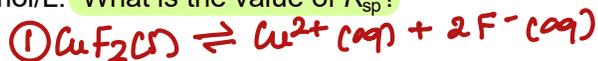
$$\therefore K_{a3} = \frac{[\text{AsO}_4^{3-}][\text{H}_3\text{O}^+]}{[\text{HASO}_4^{2-}]}$$

$$6.0 \times 10^{-10} = \frac{[z][0.158+z]}{[8.0 \times 10^{-8} - z]} \quad \frac{C}{F} \gg 100 \quad \therefore z = 8.04 \times 10^{-16} = [\text{AsO}_4^{3-}]$$

51. 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



	$[Cu^{2+}]$	$[F^{-}]$
Initial	0	0
Change	+s	+2s
Equilibrium	s	2s

② $\therefore K_{sp} = [Cu^{2+}][F^{-}]^2 = [s][2s]^2 = 4s^3$
 $\therefore K_{sp} = 4(0.0020)^3 = 3.2 \times 10^{-8}$

insoluble ionic compound

52. 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

↳ precipitate

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

B

- ~~A.~~ $\underline{Mn}O_4^{-} / (+7)$
- B. $\underline{S}O_4^{2-} / (+4)$
- ~~C.~~ $\underline{N}H_4^{+} / (-3)$
- ~~D.~~ $\underline{N}O_3^{-} / (+5)$
- ~~E.~~ $\underline{Cr}_2O_7^{2-} / (+6)$

A) $Mn + (4 \times 2-) = 1-$
 $\therefore Mn = +7$

B) $S + (4 \times 2-) = 2-$
 $\therefore S = 6+$

C) $N + (4 \times 1+) = 1+$
 $\therefore N = 3-$

D) $N + (3 \times 2-) = 1-$
 $\therefore N = 5+$

E) $2Cr + (7 \times 2-) = 2-$
 $2Cr = 12$
 $\therefore Cr = 6+$

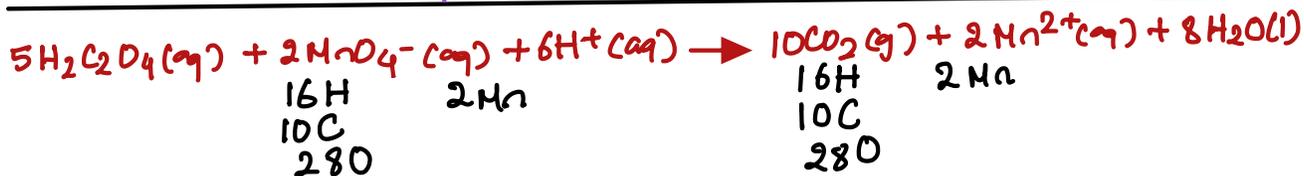
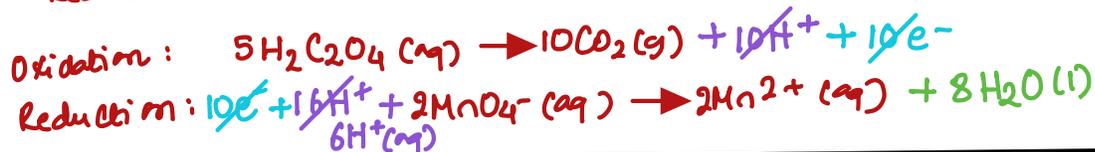
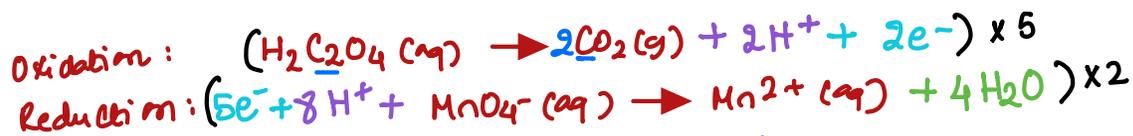
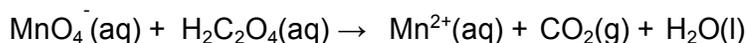
54. Identify the reaction(s) that are redox reactions:

BE

- A. $\overset{+1}{\text{H}}\overset{-1}{\text{Cl}}(\text{aq}) + \overset{+1}{\text{Na}}\overset{-1}{\text{OH}}(\text{aq}) \rightarrow \overset{+1}{\text{Na}}\overset{-1}{\text{Cl}}(\text{aq}) + \overset{+1}{\text{H}}\overset{-2}{\text{O}}(\text{l})$ Acid-base reaction \therefore not a redox reaction
- B. $\overset{-1}{5\text{I}}(\text{aq}) + \overset{+5}{\text{IO}_3^-}(\text{aq}) + \overset{+1}{6\text{H}}(\text{aq}) \rightarrow \overset{0}{3\text{I}_2}(\text{s}) + \overset{+1}{3\text{H}}\overset{-2}{\text{O}}(\text{l})$
- C. $\overset{+1}{2\text{Na}}\overset{-1}{\text{Cl}}(\text{s}) + \overset{+1}{\text{H}}\overset{+6}{\text{S}}\overset{-2}{\text{O}_4}(\text{aq}) \rightarrow \overset{+1}{2\text{H}}\overset{-1}{\text{Cl}}(\text{g}) + \overset{+1}{\text{Na}}\overset{+6}{\text{S}}\overset{-2}{\text{O}_4}(\text{s})$
- D. $\text{CaO}(\text{s}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{Ca}(\text{OH})_2(\text{aq})$
- E. $\overset{+1}{2\text{H}_2}\overset{-1}{\text{O}_2}(\text{aq}) \rightarrow \overset{+1}{2\text{H}_2}\overset{-1}{\text{O}}(\text{l}) + \overset{0}{\text{O}_2}(\text{g})$

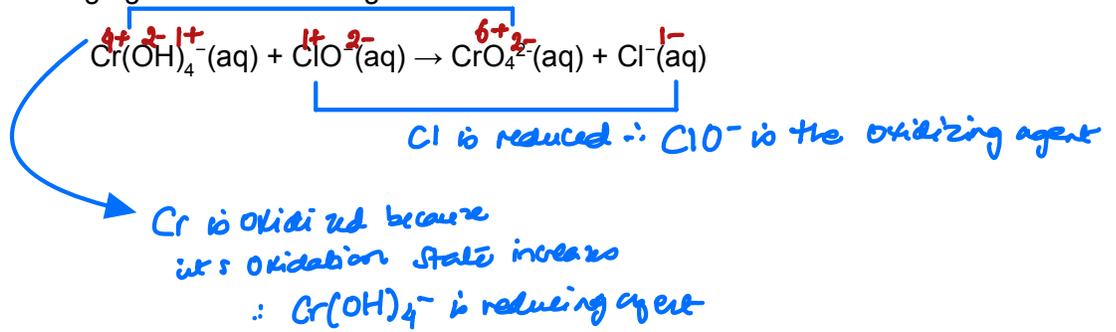
55. Balance the oxidation-reduction reaction in acidic solution. What is the sum of the coefficients?

33



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

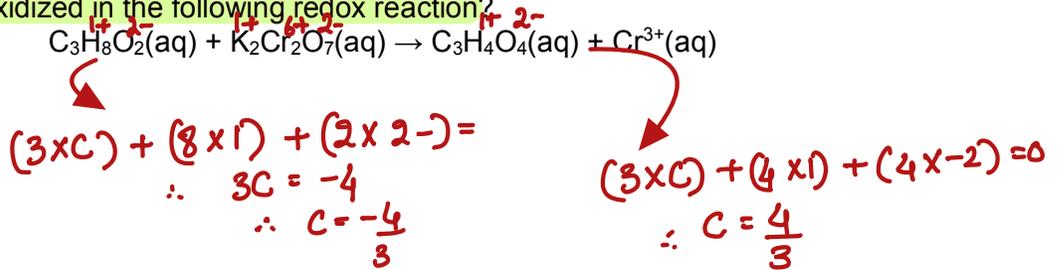
ClO^-



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

A

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



58. 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?

0.87 g



	$[\text{Ca}^{2+}]$	$[\text{ClO}_3^-]$
Initial	0	0
Change	+S	+2S
Equilibrium	S	2S

$$\textcircled{2} K_{sp} = [\text{Ca}^{2+}][\text{ClO}_3^-]^2 = [S][2S]^2 = 4S^3$$

$$\therefore 7.1 \times 10^{-7} = 4S^3$$

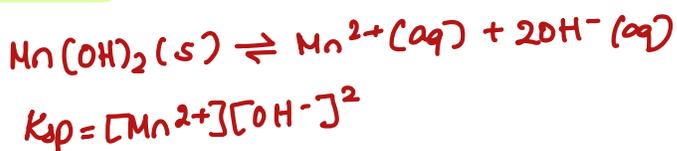
$$\therefore S = \sqrt[3]{\frac{7.1 \times 10^{-7}}{4}} = 5.619954323 \times 10^{-3} \text{ M}$$

$$\textcircled{3} 750 \text{ mL} \times \frac{5.6 \times 10^{-3} \text{ mol Ca}(\text{ClO}_3)_2}{1000 \text{ mL}} \times \frac{206.98 \text{ g Ca}(\text{ClO}_3)_2}{1 \text{ mol Ca}(\text{ClO}_3)_2} = 0.8724136093 \approx 0.87 \text{ g}$$

59. Which one of these is the solubility product constant for $\text{Mn}(\text{OH})_2$?

A

- A. $K_{sp} = [\text{Mn}^{2+}][\text{OH}^-]^2$
- B. $K_{sp} = [\text{Mn}^{2+}][2 \text{OH}^-]^2$
- C. $K_{sp} = [\text{Mn}^{2+}]^2[\text{OH}^-]^2$
- D. $K_{sp} = [\text{Mn}^{2+}][\text{OH}^-]$
- E. $K_{sp} = [\text{Mn}^{2+}]^2[2 \text{OH}^-]^2$



60. The table lists five compounds and their K_{sp} value. Which is least soluble? \rightarrow lowest K_{sp}

A

ZnS	2×10^{-25}
TlBr	3.4×10^{-6}
AgCl	1.8×10^{-10}
FeS	6×10^{-19}
CuI	1.1×10^{-12}

All salts produce 2 ions

All of the ions have the same dissociation stoichiometry which means that they produce 2 ions \therefore we can compare K_{sp} directly related to solubility

- A. ZnS
- B. TlBr
- C. AgCl
- D. FeS
- E. CuI

61. Which statement is true about redox reactions?

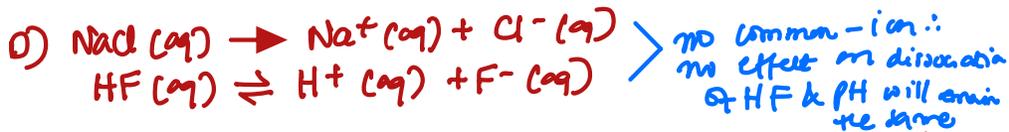
D

- ~~A.~~ A half-reaction can occur by itself
- ~~B.~~ A redox reaction in base can include excess H^+ after it has been balanced
- ~~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

- A) two half reactions, oxidation & reduction need to occur simultaneously
- B) OH^- should be present as an excess
- C) incorrect, has to be one oxidation & one reduction reaction
- D) True

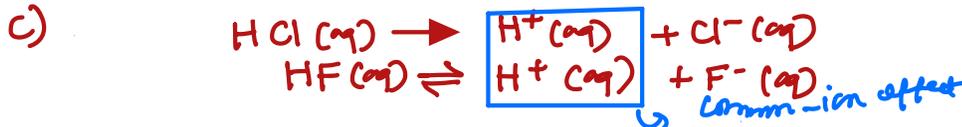
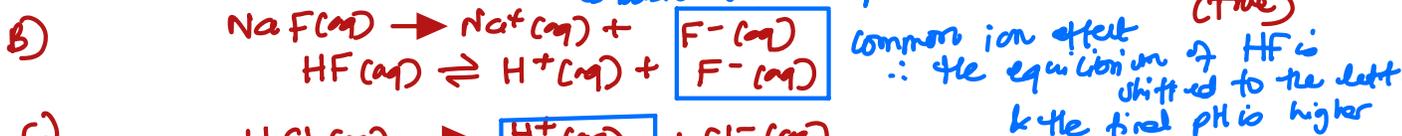
62. Which of these statements is false?

D



from the equation below NaOH (strong base) reacts with weak acid HF and shifts its equilibrium to the right

- A. Addition of NaOH will increase dissociation of HF, and the pH of the final solution will be higher
- B. Addition of NaF will decrease dissociation of HF, and the pH of the final solution will be higher
- C. Addition of HCl will decrease dissociation of HF, and the pH of the final solution will be lower
- D. Addition of NaCl will decrease dissociation of HF, and the pH of the final solution will be lower



63. For the buffer solutions below, in which case would the buffer capacity not be exhausted either by the addition of 0.5 moles of HCl or by the addition of 0.5 moles of NaOH?

B

- A. 0.80 M HF and 0.20 M NaF
- B. 0.80 M HF and 0.90 M NaF
- C. 0.10 M HF and 0.20 M NaF
- D. 0.10 M HF and 0.60 M NaF

strategy: look for the choice that has both the conc. of HF and NaF greater than 0.5 moles so neither the weak acid (HF) is completely consumed by NaOH or the conjugate base salt NaF is consumed by HCl

64. To prepare a buffer solution with pH = 4.70, how many moles of solid NaN_3 should be added to a 1.0-L solution that is 0.40 M in HN_3 ? (K_a for $\text{HN}_3 = 2.6 \times 10^{-5}$).

0

- A. 0.12 moles
- B. 0.31 moles
- C. 0.40 moles
- D. 0.52 moles
- E. none of the above

$$\text{pH} = \text{p}K_a + \log \frac{[\text{NaN}_3]}{[\text{HN}_3]}$$

$$4.70 = -\log(2.6 \times 10^{-5}) + \log \frac{[\text{NaN}_3]}{[0.40]}$$

$$4.70 = 4.585026652 + \log \frac{[\text{NaN}_3]}{[0.40]}$$

$$0.114973348 = \log \frac{[\text{NaN}_3]}{[0.40]} \rightarrow \text{take the inverse of log on both sides}$$

$$1.308086808 = \frac{[\text{NaN}_3]}{[0.40]}$$

$$\therefore [\text{NaN}_3] = 0.521234723 \frac{\text{mol}}{\text{L}} \times 1.0\text{L} = 0.52 \text{ moles}$$

reaction of cation or anion of a salt to give an acidic or basic solution

65. Which of the following equilibria best represents the hydrolysis reaction that occurs in an aqueous solution of NH_4Cl ?

B

- A. $\text{Cl}^-(\text{aq}) + \text{H}_3\text{O}^+(\text{aq}) \rightleftharpoons \text{HCl}(\text{aq}) + \text{H}_2\text{O}(\text{l})$
- B. $\text{NH}_4^+(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_3(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$
- C. $\text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightleftharpoons \text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l})$
- D. $\text{Cl}^-(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{HCl}(\text{aq}) + \text{OH}^-(\text{aq})$
- E. $\text{NH}_4^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightleftharpoons \text{NH}_4\text{Cl}(\text{s})$

conjugate acid of NH_3 \therefore weak acid $\leftarrow \text{NH}_4\text{Cl} \rightarrow$ conjugate base of strong acid \therefore neutral pH



Use this information for questions 67-69:

A 24.8 mL solution of 0.399 M aqueous hydrofluoric acid, HF, is titrated with a 0.235 M aqueous potassium hydroxide, KOH, solution. (K_a for HF = 7.2×10^{-4})

At equivalence point the moles of the acid = moles of base

66. What volume of potassium hydroxide is required to reach the equivalence point?

E

- A. 14.6 mL
- B. 24.8 mL
- C. 28.6 mL
- D. 32.4 mL
- E. 42.1 mL



$$\begin{aligned} \therefore 24.8 \text{ mL} \times \frac{0.399 \text{ mol HF}}{1000 \text{ mL}} \times \frac{1 \text{ mol KOH}}{1 \text{ mol HF}} \times \frac{1000 \text{ mL}}{0.235 \text{ mol KOH}} \\ = 42.10723404 \text{ mL} \\ \sim 42.1 \text{ mL} \end{aligned}$$

67. What is the pH of the above solution at the equivalence point?

A

- A. above 7.00
- B. below 7.00
- C. exactly 7.00



Basic salt \therefore pH is above 7

68. What is the pH of the solution after the addition of 21.1 mL of potassium hydroxide? Keep 3 decimal places.

3.145



Initial	0.0049585	0.0049585	0	—
Change	-0.0049585	-0.0049585	+0.0049585	—
Final	0.0049367	0	0.0049585	—

\downarrow weak acid
 \downarrow conjugate basic salt
 buffer

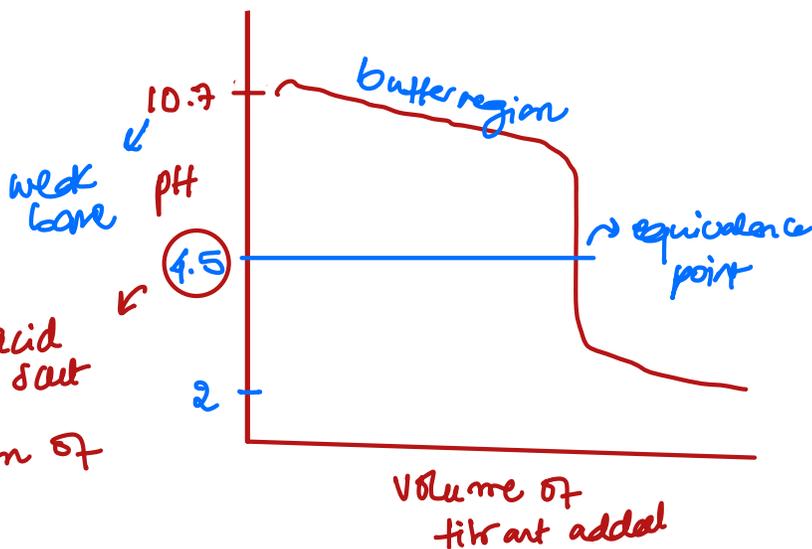
Henderson-Hasselbalch equation

$$\begin{aligned} \therefore \text{pH} &= \text{pK}_a + \log \frac{[\text{F}^-]}{[\text{HF}]} \\ &= -\log(7.2 \times 10^{-4}) + \log \frac{[0.0049585]}{[0.0049367]} \\ &= 3.144581085 \approx 3.145 \end{aligned}$$

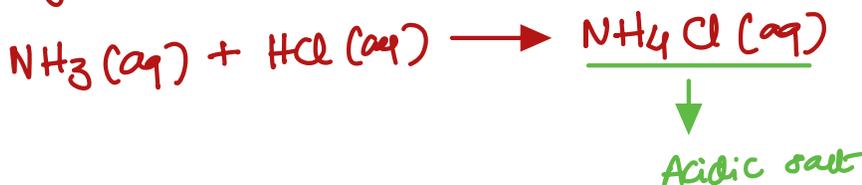
69. An initial pH of 10.7 and an equivalence point at pH = 4.5 corresponds to a titration curve for a

D

- A. Strong acid to which strong base is added
- B. Strong base to which strong acid is added
- C. Weak acid to which strong acid is added
- D. Weak base to which strong acid is added
- E. Weak base to which strong base is added

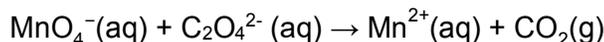


Example: the titration of NH_3 with HCl

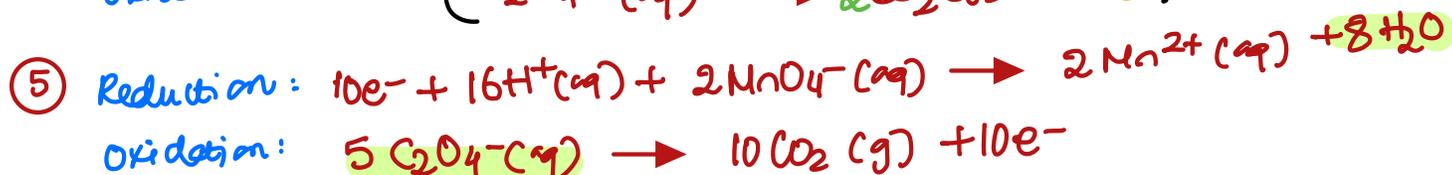
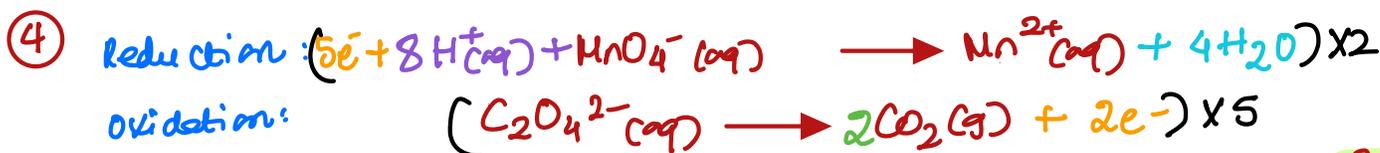
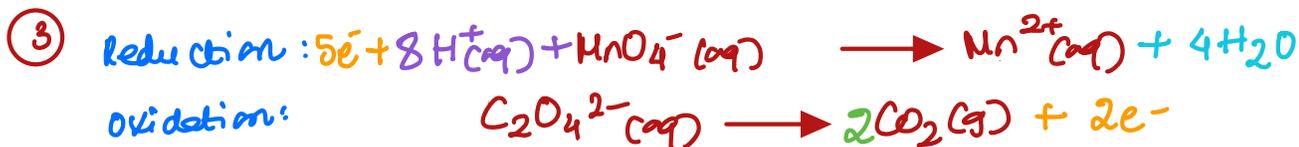
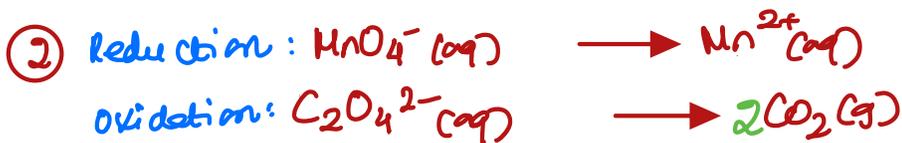
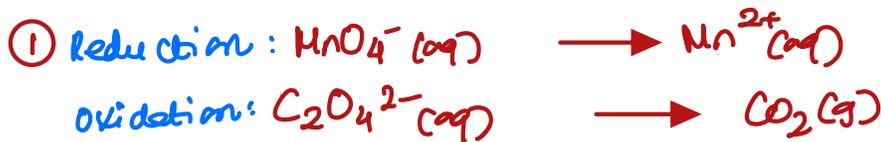


70. Consider the following redox reaction. What are the coefficients for $C_2O_4^{2-}$ and H_2O in the balanced reaction under acidic conditions?

A



- A. $C_2O_4^{2-} = 5, H_2O = 8$
- B. $C_2O_4^{2-} = 1, H_2O = 1$
- C. $C_2O_4^{2-} = 5, H_2O = 1$
- D. $C_2O_4^{2-} = 1, H_2O = 4$
- E. $C_2O_4^{2-} = 3, H_2O = 2$



71. Which of the copper(II) salts will dissolve to the greatest extent in water?

B

- A. $Cu(OH)_2$ $K_{sp} = 4.8 \times 10^{-20}$
- B. CuC_2O_4 $K_{sp} = 4.4 \times 10^{-10}$
- C. $Cu_3(PO_4)_2$ $K_{sp} = 1.4 \times 10^{-37}$
- D. CuS $K_{sp} = 8.0 \times 10^{-37}$
- E. $Cu_3(AsO_4)_2$ $K_{sp} = 8.0 \times 10^{-36}$

Method 1: Even all the copper compounds do not have the same dissociation stoichiometry, the K_{sp} values for choices A, C, D, E are way lower compared to choice B

Method 2: Calculate the solubility from K_{sp} for each compound & compare the values



	$[Cu^{2+}]$	$[OH^-]$
Initial	0	0
Change	+S	+2S
Equilibrium	S	2S

$$\begin{aligned} \therefore K_{sp} &= [Cu^{2+}][OH^-]^2 \\ 4.8 \times 10^{-20} &= [S][2S]^2 \\ 4.8 \times 10^{-20} &= 4S^3 \\ \therefore S &= \sqrt[3]{\frac{4.8 \times 10^{-20}}{4}} \\ &= 2.29 \times 10^{-7} M \end{aligned}$$



	$[\text{Cu}^{2+}]$	$[\text{C}_2\text{O}_4^{2-}]$
Initial	0	0
Change	+s	+s
Equilibrium	s	s

$$\therefore K_{sp} = [\text{Cu}^{2+}][\text{C}_2\text{O}_4^{2-}]$$

$$4.4 \times 10^{-10} = [s][s]$$

$$4.4 \times 10^{-10} = s^2$$

$$\therefore s = \sqrt{4.4 \times 10^{-10}}$$

$$= 2.10 \times 10^{-5} \text{ M}$$

↓ highest solubility



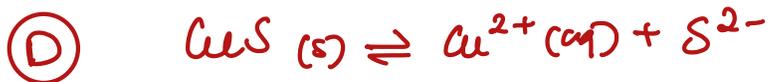
	$[\text{Cu}^{2+}]$	$[\text{PO}_4^{3-}]$
Initial	0	0
Change	+3s	+2s
Equilibrium	3s	2s

$$K_{sp} = [\text{Cu}^{2+}]^3 [\text{PO}_4^{3-}]^2$$

$$1.4 \times 10^{-37} = [3s]^3 [2s]^2$$

$$1.4 \times 10^{-37} = 108 s^5$$

$$\therefore s = 1.67 \times 10^{-8} \text{ M}$$



	$[\text{Cu}^{2+}]$	$[\text{S}^{2-}]$
Initial	0	0
Change	+s	+s
Equilibrium	s	s

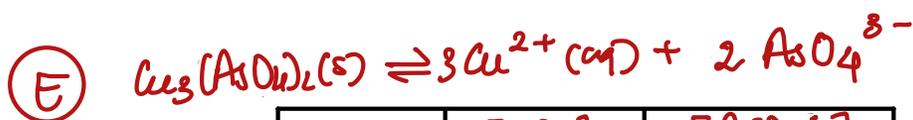
$$K_{sp} = [\text{Cu}^{2+}][\text{S}^{2-}]$$

$$8.0 \times 10^{-37} = [s][s]$$

$$8.0 \times 10^{-37} = s^2$$

$$\therefore s = \sqrt{8.0 \times 10^{-37}}$$

$$= 8.94 \times 10^{-19} \text{ M}$$



	$[\text{Cu}^{2+}]$	$[\text{AsO}_4^{3-}]$
Initial	0	0
Change	+3s	+2s
Equilibrium	3s	2s

$$K_{sp} = [\text{Cu}^{2+}]^3 [\text{AsO}_4^{3-}]^2$$

$$8.0 \times 10^{-36} = [3s]^3 [2s]^2$$

$$8.0 \times 10^{-36} = 108 s^5$$

$$\therefore s = \sqrt[5]{\frac{8.0 \times 10^{-36}}{108}}$$

$$= 3.75 \times 10^{-8} \text{ M}$$

Formula Sheet

Length

1 kilometer = 0.62137 mile
1 inch = 2.54 centimeters (exactly)
1 Ångstrom = 1×10^{-10} meter

Energy

1 joule = $1 \text{ kg}\cdot\text{m}^2/\text{s}^2$
1 calorie = 4.184 joules
1 Calorie = 1 kilocalorie = 1000 calories
1 L·atm = 101.325 joules

Pressure

1 pascal = $1 \text{ N}/\text{m}^2 = 1 \text{ kg}/\text{m}\cdot\text{s}^2$
1 atmosphere = 101.325 kilopascals = 760 mm Hg = 760 torr = 14.70 lb/in²
1 bar = 1×10^5 Pa (exactly)

Temperature

0 K = -273.15°C
K = $^\circ\text{C} + 273.15$
 $^\circ\text{C} = (5/9)(^\circ\text{F} - 32)$

Mass

1 kg = 2.205 lbs

Volume

1 mL = $1 \text{ cm}^3 = 1 \text{ cc}$

Constants

$c = 2.998 \times 10^8 \text{ m}/\text{sec}$
 $h = 6.626 \times 10^{-34} \text{ J}\cdot\text{sec}^{-1}$
 $R = 0.08206 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K} = 8.314 \text{ J}/\text{mol}\cdot\text{K}$
Specific heat of water = 4.184 J/g·K
Mass of an electron: $9.109 \times 10^{-31} \text{ kg}$
Mass of a proton: $1.673 \times 10^{-27} \text{ kg}$
 $RH = 2.18 \times 10^{-18} \text{ J}$
Specific heat of water = 4.184 J/g·K
STP = 273.15 K and 1 atm
Avogadro's number: 6.022×10^{23}

Equations

d (density) = m/V

$P_1V_1 = P_2V_2$

$V_1/T_1 = V_2/T_2$

$P_1V_1/n_1T_1 = P_2V_2/n_2T_2$

$PV = nRT$

$(P + a(n^2/V^2))\cdot(V - nb) = nRT$

molar mass (M) = mRT/PV

density (d) = MP/RT

$x_A = n_A/n_{\text{tot}} = P_A/P_{\text{tot}} = V_A/V_{\text{tot}}$

$P_{\text{tot}} = P_A + P_B + \dots$

$n_{\text{tot}} = n_A + n_B + \dots$

$$\mu_{rms} = \sqrt{\frac{3RT}{M}}$$

$$\frac{\text{Rate of effusion A}}{\text{Rate of effusion B}} = \sqrt{\frac{MW_B}{MW_A}}$$

$$Q = C \times \Delta T = c_{\text{specific}} \times m \times \Delta T$$

$$Q = n \times \Delta H \text{ (kJ/mol)} = m \times \Delta H \text{ (kJ/g)}$$

$$w = -P\Delta V$$

$$\Delta E = q + w$$

$$\Delta H^\circ = \sum n\Delta H_f^\circ(\text{products}) - \sum n\Delta H_f^\circ(\text{reactants})$$

$$\Delta H^\circ = \sum n\Delta H^\circ(\text{bonds broken}) - \sum n\Delta H^\circ(\text{bonds formed})$$

$$E = hv$$

$$c = \lambda\nu$$

$$\lambda = h/mv$$

$$\Delta E = -2.18 \times 10^{-18} J \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$$

$$\ln\left(\frac{P_2}{P_1}\right) = \frac{\Delta H_{vap}}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$$

$$C_g = kP_g$$

$$P_{\text{solution}} = P_{\text{solvent}} X_{\text{solvent}}$$

$$P_{\text{solution}} = \sum P_j = \sum P_j X_j$$

$$\Delta T_b = K_b m_i$$

$$\Delta T_f = K_f m_i$$

$$\pi = MRT_i$$

Thermodynamic and Electrochemistry

$$S = k_b \times \ln(W)$$

$$k_b = 1.381 \times 10^{-23} \text{ J/K}$$

$$\Delta S = q_{\text{rev}}/T$$

$$\Delta S_{\text{surr}} = q_{\text{surr}}/T = -q_{\text{rev}}/T$$

$$\Delta S_{\text{univ}} = \Delta S_{\text{sys}} + \Delta S_{\text{surr}}$$

$$\Delta S^\circ_{\text{rxn}} = \sum \nu S^\circ_{\text{products}} - \sum \nu S^\circ_{\text{reactants}}$$

$$\Delta H^\circ_{\text{rxn}} = \sum \nu H^\circ_{\text{products}} - \sum \nu H^\circ_{\text{reactants}}$$

$$\Delta G^\circ_{\text{rxn}} = \sum \nu G^\circ_{\text{products}} - \sum \nu G^\circ_{\text{reactants}}$$

$$\Delta G = \Delta H - T\Delta S$$

$$\Delta G = \Delta G^\circ + RT \cdot \ln Q$$

$$R = 8.314 \text{ J/mol}\cdot\text{K}$$

$$\Delta G^\circ = -RT \cdot \ln K$$

$$\Delta G = -nFE_{\text{cell}}$$

$$F = 96485 \text{ J/(V}\cdot\text{mol e}^-)$$

$$E^\circ_{\text{cell}} = RT/nF \ln K$$

$$E^\circ_{\text{cell}} = (0.0257/n) \ln K = (0.0592/n) \log K$$

$$E_{\text{cell}} = E^\circ_{\text{cell}} - (RT/nF) \ln Q$$

$$E_{\text{cell}} = E^\circ_{\text{cell}} - (0.0257/n) \ln Q$$

$$\text{Electrolysis: } Q \text{ (total charge)} = I \times t = n \times F$$

Integrated Rate Laws & half-life

$$\ln \frac{[A]}{[A]_0} = -kt$$

$$\frac{1}{[A]} = kt + \frac{1}{[A]_0}$$

$$[A] = -kt + [A]_0$$

$$t_{1/2} = \frac{[A]_0}{2k}$$

$$t_{1/2} = \frac{\ln 2}{k} = \frac{0.693}{k}$$

$$t_{1/2} = \frac{1}{k[A]_0}$$

$$\ln \frac{k_2}{k_1} = -\frac{E_a}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$$

Equilibrium and Acid / Base

$$K_p = K_c \times (RT)^{\Delta n}$$

$$\text{pH} = -\log[\text{H}_3\text{O}^+]$$

$$K_w = 1.0 \times 10^{-14} \text{ at } 25^\circ\text{C}$$

$$K_w = [\text{H}_3\text{O}^+] \times [\text{OH}^-]$$

$$K_w = K_a \times K_b$$

$$\text{p}K_a = -\log[K_a]$$

$$\text{Buffer: pH} = \text{p}K_a + \log \frac{[\text{A}^-]}{[\text{HA}]}$$

$$\ln \frac{K_2}{K_1} = \frac{\Delta H_{rxn}^\circ}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$$

Periodic Table of the Elements

1																		2																															
1 H 1.01	2																2 He 4.00																																
3 Li 6.94	4 Be 9.01															5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18																												
11 Na 22.99	12 Mg 24.31															13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95																												
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.63	33 As 74.92	34 Se 78.97	35 Br 79.90	36 Kr 83.80																																
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.95	43 Tc [97]	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29																																
57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm [145]	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.05	71 Lu 174.97																																			
87 Fr [223]	88 Ra [226]															104 Rf [267]	105 Db [268]	106 Sg [269]	107 Bh [270]	108 Hs [269]	109 Mt [277]	110 Ds [281]	111 Rg [282]	112 Cn [285]	113 Nh [286]	114 Fl [290]	115 Mc [290]	116 Lv [293]	117 Ts [294]	118 Og [294]																			
13																		14		15		16		17																									
132.91																		137.33		178.49		180.95		183.84		186.21		190.23		192.22		195.08		196.97		200.59		204.38		207.2		208.98		[209]		[210]		[222]	
89																		90		91		92		93		94		95		96		97		98		99		100		101		102		103					
[227]																		232.04		231.04		238.03		[237]		[244]		[243]		[247]		[247]		[251]		[252]		[257]		[258]		[259]		[262]					