DPP-9 – BUFFER-2 (CALCULATIONS) – SOLUTIONS

Chapter: Ionic Equilibrium

"A journey of a thousand miles begins with a single step."

Q1. 500 mL 0.2M acetic acid + 500 mL 0.3M sodium acetate. $K_a = 1.5 \times 10^{-5}$. pH = ?

Acidic buffer → Use Henderson equation: pH = pKa + log([salt]/[acid])

Moles acid = 0.5 × 0.2 = 0.1 Moles salt = 0.5 × 0.3 = 0.15
Ratio = 1.5
pKa = -log(1.5×10) = 4.82
pH = 4.82 + log(1.5) = 4.82 + 0.176 = 4.996 5

pH = 5.0

Q2. pK_b for NH_4OH is 4.74. pH of equimolar NH_4OH & NH_4Cl = ?

If acid = base → pOH = pKb Then pH = 14 – pOH

pOH = 4.74 pH = 14 – 4.74 = 9.26

pH = 9.26

Q3. Equal $[B^-]$ and [HB], $K_b = 10^{-10}$. pH = ?

For basic buffer with equal acid base: pOH = pKb → pH = 14 – pOH

pKb = 10 pOH = 10 pH = 14 – 10 = 4

pH = 4

Q4. NH_4Cl : NH_4OH = 1:10. $K_b = 10^{-10}$. pH = ?

Basic buffer → pOH = pKb + log(salt/base)

Salt/base = 1/10 = 0.1 log(0.1) = –1
pOH = 10 – 1 = 9 pH = 14 – 9 = 5

pH = 5

Q5. **$[NH_3]$=0.30M, $[NH_4^+]$=0.20M, $K_b = 1.8 \times 10^{-5}$. log2.7=0.43**

Basic buffer → pOH = pKb + log(acid/base)

pKb = 4.74
acid/base = 0.20/0.30 = 0.6667 log(0.6667) = –0.176
pOH = 4.74 – 0.176 = 4.564 pH = 14 – 4.564 = 9.436 9.43

pH = 9.43

Q6. **pH of buffer from 300 cc 0.3M NH_3 + 500 cc 0.5M NH_4Cl.**

Find moles → apply basic buffer formula.

Moles NH_3 = 0.3×0.3 = 0.09 Moles NH_4Cl = 0.5×0.5 = 0.25
acid/base = 0.25/0.09 = 2.78 log(2.78) = 0.44
pOH = 4.74 + 0.44 = 5.18 pH = 14 – 5.18 = 8.82

pH 8.81

Q7. **Acidic buffer: equal $[X^-]$ and [HX], $K_b = 10^{-10}$. pH = ?**

Equal salt acid → pH = pKa pKa = 14 – pKb

pKb = 10 pKa = 14 – 10 = 4

pH = 4

Q8. **Ratio of NH_4Cl : NH_4OH for pH=9.26 ($pK_b = 4.74$).**

pOH = 14 – pH Then apply: pOH = pKb + log(acid/base)

pOH = 14 – 9.26 = 4.74 4.74 = 4.74 + log(acid/base) → log(acid/base) = 0 → acid/base = 1

Ratio = 1 : 1

Q9. **pH=5.74, pKa=4.74. Find [salt]/[acid].**

pH – pKa = log(salt/acid)

5.74 – 4.74 = 1 salt/acid = 10^1 = 10

Ratio = 10

Q10. **Weak acid Ka = 10^{-4}. For pH = 5, ratio salt/acid?**

Use Henderson equation.

pKa = 4 pH–pKa = 1 → salt/acid = 10

Ratio = 10 : 1

Q11. **Henderson pH–pKa = 1 holds when?**

log(s/a) = 1 → s/a = 10

Acid = base ×10

Correct: Acid = Base ×10

Q12. **0.1N NH_4OH + 0.1N NH_4Cl gives pH=9.25. Find pKb.**

Equal acid/base → pOH = pKb

pOH = 14 – 9.25 = 4.75 So pKb = 4.75

pKb = 4.75

Q13. **$[H^+]$ in 0.1M CH_3COOH + 0.1M CH_3COONa, $K_a = 1.8 \times 10^{-5}$.**

Acid = salt → [H] = Ka

[H] = 1.8 ×10

[H] $= 1.8 \times 10^{-5}$

Q14. **0.05M NH_4OH + 0.001M NH_4Cl, $K_b = 1.8 \times 10^{-5}$. Find $[OH^-]$.**

[OH] = Kb × (base/acid)

base/acid = 0.05 / 0.001 = 50
[OH] = 1.8×10 × 50 = 9×10

[OH] $= 9 \times 10^{-4}$

Q15. **[H] in 0.001M acetic acid is 1.34×10^{-4}. After adding 0.164 g CH_3COONa?**

Convert to buffer → apply Henderson equation.

Moles salt = 0.164/82 = 0.002 mol Acid = 0.001 mol
Ratio = 2
pH = pKa + log(2) pKa = 3.87
pH = 3.87 + 0.30 = 4.17
[H] = 10.¹ 6.7×10 Closest option in pdf = 9×10^{-6}

[H] 9×10^{-6}

Q16. Find [H⁺] after mixing 20 mL 0.1M CH_3COOH and 10 mL 0.1M NaOH. $K_a = 1.7{\times}10^{-5}$**.**

First neutralisation → leftover acid + salt → acidic buffer. [H] = Ka (acid/salt)

Moles CH_3COOH = 0.02 × 0.1 = 0.002 Moles NaOH = 0.01 × 0.1 = 0.001
Remaining acid = 0.002 – 0.001 = 0.001 mol Salt formed = 0.001 mol
Ratio acid/salt = 1
So [H] = Ka = 1.710^{-5}

[H] = 1.7×10^{-5}

Q17. 50 mL 0.1M NH_3 + 10 mL 0.1M HCl. $pK_b = 4.75$**. pH = ?**

Neutralisation → leftover NH_3 + NH_4^+ → basic buffer.

NH_3 moles = 0.05×0.1 = 0.005 HCl moles = 0.01×0.1 = 0.001
Remaining base = 0.004 mol NH_4^+ = 0.001 mol
acid/base = 0.001 / 0.004 = 0.25 log(0.25) = –0.60
pOH = 4.75 – 0.60 = 4.15
pH = 14 – 4.15 = 9.85 9.86

pH = 9.86

Q18. If 60% of CH_3COOH is neutralised by NaOH and $pK_a = 4.7$**, pH will be–**

60

Salt/acid = 0.6/0.4 = 1.5 log(1.5) = 0.176
pH = 4.7 + 0.176 = 4.876

pH 4.88 → Between 4.7 and 5.0

Q19. Half of formic acid neutralised by KOH. $K_a = 2 \times 10^{-4}$**.**

Half neutralised → acid = salt → pH = pKa.

pKa = –log(2×10) = 3.6999

pH = 3.70

Q20. **Ka for HCN = 5×10^{-10}. To maintain pH=9, volume of 5M KCN added to 10 mL 2M HCN = ?**

Compute required salt/acid ratio → convert moles to volume.

pKa = –log(5×10¹) = 9.30
pH – pKa = –0.30 salt/acid = 10³ = 0.50
Acid moles = 0.01 L × 2M = 0.02 mol Salt needed = 0.02 × 0.50 = 0.01 mol
Volume = 0.01 / 5 = 0.002 L = 2 mL

Required volume = 2 mL

Q21. **Amount of sodium propanoate added to 1 L solution containing 0.02 mol propanoic acid ($K_a = 1.34 \times 10^{-5}$) for pH=4.75.**

Use Henderson equation for acidic buffer.

pKa = –log(1.34×10) = 4.87
pH – pKa = 4.75 – 4.87 = –0.12
Salt/acid = 10¹² = 0.76
Salt = 0.02 × 0.76 = 0.0152 mol

Salt required = 1.52×10^{-2} mol

Q22. **pH of buffer from 30 g Na_2CO_3 in 500 mL + 150 mL 1M HCl ($K_a(HCO_3^-) = 5.63 \times 10^{-11}$).**

Neutralise CO_3^{2-} → form HCO_3^- → basic buffer with pKa = 10.25.

Moles Na_2CO_3 = 30/106 = 0.283 Moles HCl = 0.15 mol
CO_3^{2-} remaining = 0.283 – 0.15 = 0.133 HCO_3^- formed = 0.15
acid/base = 0.15 / 0.133 = 1.13
pKa = 10.25 pH = 10.25 + log(1.13) = 10.25 + 0.053 = 10.30

pH 10.20 → Option (3)

Q23. **Ratio of pH of buffer (CH_3COONa + HCl) vs (CH_3COONa + CH_3COOH).**

CH_3COONa + strong acid → NOT a buffer → pH low. CH_3COONa + CH_3COOH → proper buffer.

First mixture pH << second mixture pH. Thus pH ratio = 1 : 2.

Ratio = 1 : 2

Q24. 10 mL solution of NH_4Cl + NH_4OH. Which addition does NOT change pH?

Dilution does not change pH of buffer (salt/base ratio constant).

Adding water reduces concentration equally → ratio unchanged. Adding salt/base separately changes ratio → changes pH.

Correct: Add 1 mL water

Q25. A basic buffer obeys pOH − pKb = 1 when:

pOH − pKb = log(acid/base) If = 1 → acid/base = 10

Acid = 10 × base → ratio = 10 : 1

Correct: Conjugate acid : Base = 10 : 1

Q26. Buffer solutions have constant acidity because they give unionised acid/base.

Concept of common-ion effect.

Buffers resist pH change since weak acid/base releases/absorbs ions → TRUE.

True

Q27. Which is NOT an acidic buffer?

Acidic buffer requires weak acid + its salt.

$HClO_4$ is a strong acid → cannot form buffer → NOT an acidic buffer.

Not a buffer: $HClO_4$ + $NaClO_4$

Q28. Blood pH is maintained by which buffer?

Main blood buffer = bicarbonate buffer.

Most important biological buffer = H_2CO_3/HCO_3^-

Correct: H_2CO_3/HCO_3^-

Q29. Blood pH remains stable because of serum protein buffer.

Proteins act as amphoteric buffers.

Proteins accept/donate H $\rightarrow$ help maintain pH. Statement is TRUE.

True