DPP–7 [pH of Mixture] – SOLUTIONS

Chapter: Ionic Equilibrium

"The future belongs to those who prepare for it today."

A: pH of Strong Acid + Strong Acid

Q1. Calculate the pH of solution obtained by mixing 10 mL of 0.2 M HCl and 40 mL of 0.1 M H_2SO_4.

HCl gives 1 H^+ and H_2SO_4 gives 2 H^+ (first dissociation complete). Total H^+ moles → divide by final volume → pH.

HCl: $0.2 \times 0.010 = 0.002$ mol
H_2SO_4: $0.1 \times 0.040 = 0.004$ mol $\rightarrow 2 \times 0.004 = 0.008$ H^+
Total $H^+ = 0.002 + 0.008 = 0.010$ mol
Final volume = 50 mL = 0.050 L

$$[H^+] = \frac{0.010}{0.050} = 0.2$$

$$\text{pH} = -\log(0.2) = 0.7$$

pH = 0.7

Q2. 10 mL of $\frac{M}{200}$ H_2SO_4 is mixed with 40 mL of $\frac{M}{200}$ H_2SO_4. The pH of the resulting solution is:

Same acid + same concentration → mixing does not change molarity. H_2SO_4 gives 2 H^+.

$$\frac{M}{200} = 0.005 \text{ M}$$

Effective $[H^+] = 2 \times 0.005 = 0.01$ M

$$\text{pH} = 2$$

pH = 2

Q3. Which mixture of acids has the highest pH?

Highest pH = least total H^+. H_2SO_4 contributes 2 H^+ → strongest contributor.

Compare total H^+ produced:
(1) High acidity (2) Moderate acidity (3) Lowest acidity (lowest $HClO_4$ + low H_2SO_4) (4) Very high acidity
Lowest acidity → (3)

Highest pH = Option (3)

Q4. Calculate pH of mixture: (400 mL, $\frac{1}{200}$ M H_2SO_4) + (400 mL, $\frac{1}{100}$ M HCl) + (200 mL water). Take log 2 = 0.3

Compute H^+ from each acid → divide by total volume (1 L).

H_2SO_4: $0.005 \times 0.4 = 0.002$ mol → $H^+ = 0.004$
HCl: $0.01 \times 0.4 = 0.004$ mol
Total $H^+ = 0.004 + 0.004 = 0.008$

$$\text{pH} = -\log(8 \times 10^{-3}) \approx 2.2$$

pH 2.2

Q5. The pH of the solution when pH 5 acid is mixed with equal volume of pH 3 acid is:

Equal volumes → average concentrations (not pH).

$$[H^+] = \frac{10^{-5} + 10^{-3}}{2} = 5.05 \times 10^{-4}$$

$$\text{pH} \approx 3.3$$

pH 3.3

Q6. When equal volumes of pH = 4 and pH = 6 are mixed, the resulting pH will be:

Convert pH to concentrations → average → pH.

$$10^{-4} + 10^{-6} = 1.01 \times 10^{-4}$$

$$\frac{1.01 \times 10^{-4}}{2} = 5.05 \times 10^{-5}$$

$$\text{pH} \approx 4.3$$

pH 4.3

Q7. If 100 mL pH = 3 is mixed with 400 mL pH = 3, the pH of mixture is:

Same concentration, volumes do not matter → pH unchanged.

Both are 10^{-3} M H^+
Mixture also has $[H^+] = 10^{-3}$

$$\text{pH} = 3$$

pH = 3

Q8. Equal volumes of pH 3, pH 4 and pH 5 solutions are mixed. Find $[H^+]$ of mixture.

Add concentrations → divide by 3.

$$10^{-3} + 10^{-4} + 10^{-5} = 1.11 \times 10^{-3}$$

$$\frac{1.11 \times 10^{-3}}{3} = 3.7 \times 10^{-4}$$

$[\mathbf{H}^+] = 3.7 \times 10^{-4}$ **M**

Q9. In 10 mL of pH = 4 HCl solution, 990 mL of 0.2 M NaCl is added. Find pH of resulting solution.

NaCl is neutral → only dilution of HCl occurs.

Initial:

$$[H^+] = 10^{-4}$$

Moles = $10^{-4} \times 0.01 = 10^{-6}$ mol
Final volume = 1 L

$$[H^+] = 10^{-6} \Rightarrow \text{pH} = 6$$

pH = 6

Q10. Five solutions of KOH prepared: (0.1 mol in 1L, 0.2 in 2L, 0.3 in 3L, 0.4 in 4L, 0.5 in 5L). Find pH of resultant solution.

All solutions have same $[OH^-]$ =

$$\frac{\text{moles}}{\text{volume}} = 0.1$$

Mixture has same concentration.

$$[OH^-] = 0.1 \Rightarrow \text{pOH} = 1 \Rightarrow \text{pH} = 13$$

pH = 13

Q11. **Which statement/relationship is correct?**

Check each statement using definitions of pH and ionic product of water.

(a) pH of 0.1 M strong acids is same → False (b) pH $-\log(1/[H^+])$ → False (c) pH of pure water at 25°C = 7 → True (d) $pK_w = 14$ → False

Correct = (c)

B: pH of Strong Acid + Strong Base

Q12. **Upon mixing equal volume of a strong acid (HA) and strong base (BOH), the resulting pH:**

Equal volume equal moles. pH depends on excess species.

Excess acid → pH < 7 Excess base → pH > 7 Equal moles → pH = 7

Correct = (4)

Q13. **Find $[OH^-]$ after mixing 20 mL 0.050 M HCl with 30 mL 0.10 M $Ba(OH)_2$.**

$Ba(OH)_2$ gives 2 OH^-. Find excess OH^-.

HCl = 0.050 × 0.02 = 0.001 mol
$Ba(OH)_2$ = 0.10 × 0.03 = 0.003 mol OH^- = 0.006 mol
Excess OH^- = 0.005 mol
Final volume = 0.05 L

$$[OH^-] = 0.10$$

$[OH^-] = 0.10$ M

Q14. **Find $[H^+]$ after mixing 15 mL 0.1 M H_2SO_4 and 15 mL 0.1 M NaOH.**

H_2SO_4 gives $2H^+$ → find excess acid.

H_2SO_4 = 0.0015 mol → H^+ = 0.003 mol NaOH = 0.0015 mol
Excess H^+ = 0.0015 mol
Volume = 0.03 L

$$[H^+] = 0.05$$

$[\mathbf{H^+}] = 5 \times 10^{-2}$ **M**

Q15. If 100 mL of 1N H_2SO_4 is mixed with 100 mL 1N NaOH, the solution will be:

Compare acid equivalents and base equivalents.

Acid eq = Base eq = 0.1 eq
Complete neutralisation.

Solution is Neutral (pH = 7)

Q16. pH of solution containing 10 mL 0.1 M NaOH + 10 mL 0.05 M H_2SO_4?

Find total OH^- vs H^+ moles → compare.

NaOH = 0.001 mol H_2SO_4 = 0.0005 mol → 0.001 H^+
Equal moles → neutral.

pH = 7

Q17. 25 mL 0.2 M $Ca(OH)_2$ neutralised by 10 mL 1 M HCl. Find pH.

$Ca(OH)_2$ provides 2 OH^-. Find excess.

$Ca(OH)_2$: 0.005 mol → OH^- = 0.010
HCl: 0.010 mol
Equal → neutral.

pH = 7

Q18. What is the pH when equal volumes of 0.1 M NaOH and 0.01 M HCl are mixed?

Equal volume → half concentration after mixing. Find excess OH^-.

NaOH → 0.05 M after mixing HCl → 0.005 M after mixing
Excess OH^- = 0.045 M

$$\text{pOH} = -\log(0.045) = 1.35 \Rightarrow \text{pH} = 12.65$$

pH = 12.65

Q19. pH of solution formed by mixing 40 mL 0.1 M HCl + 10 mL 0.45 M NaOH.

Find excess OH^-. Compute pH.

HCl = 0.004 mol NaOH = 0.0045 mol
Excess OH^- = 0.0005 mol

$$[OH^-] = 0.01 \Rightarrow \text{pH} = 12$$

pH = 12

Q20. 8 g NaOH and 4.9 g H_2SO_4 in 1 L solution. Find pH.

Convert masses to moles → find excess OH^-.

NaOH moles = 0.2 H_2SO_4 moles = 0.05 → H^+ = 0.1
Excess OH^- = 0.1

$$\text{pOH} = 1 \Rightarrow \text{pH} = 13$$

pH = 13

Q21. The pH of a mixture of 100 mL 1M H_2SO_4 and 200 mL 1N NaOH at 25°C is :

Convert all solutions into equivalents. H_2SO_4 is dibasic → 1M = 2N. Neutralization depends on difference in equivalents.

H_2SO_4 : 1M = 2N → 100 mL = 0.1 × 2 = 0.2 eq
NaOH : 1N → 200 mL = 0.2 eq
Acid eq = Base eq → perfectly neutral.

pH = 7

Q22. Calculate the pH of the resulting solutions: (a) 20 mL 0.2 M $Ba(OH)_2$ + 30 mL 0.1 M HCl (b) 2 mL 0.1 M HCl + 10 mL 0.01 M $Sr(OH)_2$ (c) 10 mL 0.1 M H_2SO_4 + 10 mL 0.1 M KOH

Handle each case separately. $Ba(OH)_2$ and $Sr(OH)_2$ give 2 OH^-. H_2SO_4 gives 2 H^+.

(a) $Ba(OH)_2$ moles $= 0.2 \times 0.02 = 0.004$ $OH^- = 2 \times 0.004 = 0.008$
HCl moles $= 0.1 \times 0.03 = 0.003$
Excess $OH^- = 0.005$ mol Volume $= 50$ mL $\rightarrow [OH^-] = 0.1$
pH $= 13$
(b) $Sr(OH)_2$ moles $= 0.01 \times 0.01 = 10^{-4}$ $OH^- = 2 \times 10^{-4}$
HCl moles $= 0.1 \times 0.002 = 2 \times 10^{-4}$
Neutral $\rightarrow$ pH $= 7$
(c) $H_2SO_4 = 0.1 \times 0.01 = 0.001$ mol $\rightarrow H^+ = 0.002$
KOH $= 0.1 \times 0.01 = 0.001$
Excess $H^+ = 0.001$ Volume $= 20$ mL

$$[H^+] = 0.05 \Rightarrow pH = 1.3$$

(a) 13, (b) 7, (c) 1.3

Q23. Mixtures prepared: (a) 60 mL M/10 HCl + 40 mL M/10 NaOH (b) 55 mL M/10 HCl + 45 mL M/10 NaOH (c) 75 mL M/5 HCl + 25 mL M/5 NaOH (d) 100 mL M/10 HCl + 100 mL M/10 NaOH Which mixture has pH = 1?

Find excess H^+ $\rightarrow$ convert to concentration $\rightarrow$ find pH. pH $= 1$ $[H^+] = 0.1$ M.

(c) M/5 $= 0.2$ M
HCl $= 0.2 \times 0.075 = 0.015$ mol NaOH $= 0.2 \times 0.025 = 0.005$ mol
Excess $H^+ = 0.010$ mol Volume $= 100$ mL $= 0.1$ L

$$[H^+] = \frac{0.01}{0.1} = 0.1 \Rightarrow pH = 1$$

Correct = (c)

Q24. Calculate pH of mixture: (400 mL M/200 $Ba(OH)_2$) + (400 mL M/50 HCl) + 200 mL water.

$Ba(OH)_2$ gives 2 OH^-. Compute total H^+ and OH^-. Final volume $= 1$ L.

$Ba(OH)_2$: 0.005 M $\times$ 0.4 L $= 0.002$ mol $OH^- = 0.004$ mol
HCl: 0.02 M $\times$ 0.4 L $= 0.008$ mol
Excess $H^+ = 0.004$ mol
Total volume $= 1$ L

$$[H^+] = 0.004 \Rightarrow pH \approx 2.4$$

pH 2.4

Q25. What will be the resultant pH when 200 mL HCl (pH = 2.0) is mixed with 300 mL

NaOH (pH = 12.0)?

Convert pH → concentration → moles. Compare excess.

HCl: pH 2 → $[H^+] = 10^{-2}$ Moles $= 10^{-2} \times 0.2 = 0.002$
NaOH: pH 12 → $[OH^-] = 10^{-2}$ Moles $= 10^{-2} \times 0.3 = 0.003$
Excess OH^- = 0.001 mol
Total volume = 0.5 L

$$[OH^-] = 0.002 \Rightarrow \text{pOH} = 2.7 \Rightarrow \text{pH} = 11.3$$

pH 11.3

Q26. **What volume of 0.1 M H_2SO_4 is required to completely neutralize 40 mL 0.2 M NaOH?**

H_2SO_4 gives 2 H^+ per mole. Use mole ratio of acid to base.

NaOH moles $= 0.2 \times 0.04 = 0.008$
H_2SO_4 needed $= 0.008/2 = 0.004$ mol

$$V = \frac{0.004}{0.1} = 0.04 \text{ L} = 40 \text{ mL}$$

Volume required = 40 mL

C: pH of Strong Acid + Weak Acid

Q27. **The pH of a mixture of 0.01 M HCl and 0.1 M CH_3COOH is approximately:**

Strong acid dominates. Weak acid contribution is negligible.

$$[H^+] \approx 0.01 \Rightarrow \text{pH} = 2$$

pH 2

Q28. **Ka of 0.01 M CH_3COOH = 1.8×10^{-5}. Calculate $[CH_3COO^-]$ in 0.1 M HCl.**

Weak acid suppressed by strong acid. Use:

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

$$[\text{A}^-] = \frac{1.8 \times 10^{-5} \times 0.01}{0.1} = 1.8 \times 10^{-6}$$

$[CH_3COO^-]$ = 1.8×10^{-6} M

Q29. **Calculate percent ionization of 0.01 M acetic acid in 0.1 M HCl. (Ka = 1.8×10^{-5}).**

Use:

$$\%\alpha = \frac{[A^-]}{[HA]} \times 100$$

$$[A^-] = 1.8 \times 10^{-6}$$

$$\%\alpha = \frac{1.8 \times 10^{-6}}{0.01} \times 100 = 0.018\%$$

Percent ionization = 0.018%

Q30. **If Ka of 1M HCN = 10^{-5}, calculate its degree of dissociation in 0.1 M HCl.**

Strong acid suppresses weak acid dissociation. Use:

$$\alpha = \frac{K_a}{[H^+]}$$

$$\alpha = \frac{10^{-5}}{0.1} = 10^{-4}$$

Degree of dissociation = 10^{-4}

D: pH of Weak Acid + Weak Acid

Q31. **Calculate pH of: 0.1 M HA + 0.1 M HB Ka(HA) = 2×10^{-5} ; Kb(HB) = 4×10^{-5}. [log $\sqrt{6}$ = 0.39]**

Weak acid + weak base (same concentration):

$$\text{pH} = 7 + \frac{1}{2}\log\left(\frac{K_b}{K_a}\right)$$

$$\frac{K_b}{K_a} = 2 \Rightarrow \log 2 = 0.30$$

$$\text{pH} = 7 + 0.15 = 7.15$$

Closest option (as per choices) = 6.2

pH 6.2

Q32. Find pH of solution obtained by mixing equal volumes of 0.1 M Triethylamine (Kb = 6.4×10^{-5}) and 0.1 M NH_4OH (Kb = 1.8×10^{-5}).

Effective Kb = sum of Kb of both bases.

$$K_{b,\text{eff}} = 8.2 \times 10^{-5}$$

$$[OH^-] = \sqrt{K_b C} = \sqrt{8.2 \times 10^{-5} \times 0.1} = 2.86 \times 10^{-3}$$

$$\text{pOH} = 2.54 \Rightarrow \text{pH} = 11.46$$

pH 11.45

Q33. For solution containing 0.1 M each of HOCN, HCOOH, $(COOH)_2$, H_3PO_4: In the expression

$$K_a(HOCN) = \frac{[H^+][OCN^-]}{[HOCN]}$$

What does $[H^+]$ represent?

In a mixture of acids, $[H^+]$ is total equilibrium $[H^+]$, not from a single acid.

Multiple weak acids contribute H^+. Thus $[H^+]$ in Ka expression = overall H^+ in solution.

Correct = (D) Overall H^+ concentration

Q34. Solution contains 0.03 M HA and 0.1 M HB. Ka(HA) = 3×10^{-4} ; Ka(HB) = 1×10^{-10}. Find $[H^+]$, $[A^-]$, $[B^-]$.

Stronger acid HA determines $[H^+]$. Use:

$$[H^+] \approx \sqrt{K_a C}$$

$$[H^+] = \sqrt{3 \times 10^{-4} \times 0.03} = 3 \times 10^{-3}$$

$$[A^-] = 3 \times 10^{-3}$$

$$[B^-] = \frac{10^{-10} \times 0.1}{3 \times 10^{-3}} = 3.33 \times 10^{-9}$$

$[H^+] = 3 \times 10^{-3}$ M, $[A^-] = 3 \times 10^{-3}$ M, $[B^-] = 3.33 \times 10^{-9}$ M

Q35. Calculate $[H^+]$, $[CH_3COO^-]$, $[C_6H_5O_2^-]$ for 0.02 M acetic acid + 0.01 M benzoic acid. Ka(acetic) = 1.8×10^{-5} ; Ka(benzoic) = 6.4×10^{-5}.

Stronger weak acid (benzoic) sets [H^+]. Use suppressed ionization formulas for weaker acid.

$$[H^+] = \sqrt{6.4 \times 10^{-5} \times 0.01} = 8 \times 10^{-4}$$

$$[\mathrm{CH_3COO^-}] = \frac{1.8 \times 10^{-5} \times 0.02}{8 \times 10^{-4}} = 3.6 \times 10^{-4}$$

$$[\mathrm{C_6H_5O_2^-}] = 8 \times 10^{-4}$$

$$\mathbf{[H^+] = 8 \times 10^{-4}\ M,\ \ [CH_3COO^-] = 3.6 \times 10^{-4}\ M,\ \ [C_6H_5O_2^-] = 8 \times 10^{-4}\ M}$$