CHEMISTRY (CODE - 043)

CHAPTER TEST - STEP-BY-STEP SOLUTIONS AND MARKING SCHEME

Class XI (2026-27) | Chapter 6: Equilibrium

Max. Marks: 35

SECTION A (12 × 1 = 12 marks)

Each correct answer carries 1 mark.

Q1. Ans: C. $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ [1]

$K_p = K_c(RT)^{\Delta n}$. They are equal only when $\Delta n = 0$. For $H_2 + I_2 \rightleftharpoons 2HI$, $\Delta n = 2 - 2 = 0$. The others have $\Delta n = +1, -2$ and $+1$.

Q2. Ans: B. 1.5 [1]

At equilibrium (plateau): [X] = 0.40 M and [Y] = 0.60 M. $K_c = [Y]/[X] = 0.60 / 0.40 =$ **1.5**.

Q3. Ans: A. high pressure and low temperature [1]

The forward reaction is exothermic, so low temperature favours NH_3. The number of moles of gas decreases (4 → 2), so high pressure favours the forward reaction.

Q4. Ans: D. forward, at equilibrium, reverse [1]

X: $Q_c < K_c$ → net reaction goes forward. Y: $Q_c = K_c$ → at equilibrium. Z: $Q_c > K_c$ → net reaction goes in the reverse direction.

Q5. Ans: B. BF_3 [1]

BF_3 is electron deficient and accepts a lone pair (from NH_3) → Lewis acid. NH_3, OH^- and H_2O donate lone pairs (Lewis bases).

Q6. Ans: D. 12 [1]

$[OH^-] = 2 \times 0.005 = 0.01\ M = 10^{-2}\ M$ → pOH = 2 → pH = 14 − 2 = **12**.

Q7. Ans: C. S [1]

$[OH^-] = 10^{-4}$ M → pOH = 4 → pH = 14 − 4 = 10, which is solution **S**.

Q8. Ans: B. $4S^3$ [1]

$MX_2 \rightleftharpoons M^{2+} + 2X^-$: $[M^{2+}] = S$, $[X^-] = 2S$. $K_{sp} = S \times (2S)^2 =$ **$4S^3$**.

Q9. Ans: C. $CH_3COOH + CH_3COONa$ [1]

An acidic buffer is a weak acid with its salt formed with a strong base (the conjugate base), so that it resists the change in pH.

Q10. Ans: A. Both A and R are true, and R is the correct explanation of A. [1]

A catalyst lowers E_a of both directions equally, so the rates of both reactions increase equally; the equilibrium composition and K_c do not change.

Q11. Ans: D. A is false but R is true. [1]

For 10^{-8} M HCl the H^+ from water cannot be ignored: $[H^+] = 10^{-8} + x$, giving $[H^+] \approx 1.05 \times 10^{-7}$ M, so pH ≈ 6.98 (an acid cannot give a pH above 7). A is false; R is true.

Q12. Ans: B. Both A and R are true, and R is not the correct explanation of A. [1]

For an exothermic reaction K_c decreases with a rise in temperature (the equilibrium shifts to the left). The increase in rate with temperature is true but does not explain the change in K_c.

SECTION B (4 × 2 = 8 marks)

Q13. [2]

(i) P = H_2 (jumps up when added, then decreases); Q = I_2 (decreases as it reacts with the extra H_2); R = HI (increases). The reactant added was **H_2**. [1]

(ii) Le Chatelier's principle: a change in any of the factors that determine the equilibrium conditions causes the system to change in such a way as to reduce or counteract the effect of the change. [½] K_c does **not** change (it depends only on temperature); the equilibrium shifts forward and is re-established with the same K_c. [½]

Q14. [2]

(i) Larger K_a means a stronger acid: **HCN < CH_3COOH < HCOOH < HF**. [½] The weakest acid (HCN) has the strongest conjugate base (CN^-). [½]

(ii) $K_a \times K_b = K_w \rightarrow K_b = 10^{-14} / (5 \times 10^{-10}) =$ **2×10^{-5}** [1]

Q15. [2]

(i) $K_c = [NO_2]^2 / [N_2O_4]$; $K_p = (p_{NO2})^2 / p_{N2O4}$ [1]

(ii) $\Delta n = 2 - 1 = +1$. $K_p = K_c(RT)^{\Delta n} = 0.05 \times (0.08 \times 500) = 0.05 \times 40 =$ **2 bar** [1]

Q16. [2]

(i) P (CH_3COOH) = acid (donates H^+); Q (H_2O) = base (accepts H^+); R (H_3O^+) = conjugate acid; S (CH_3COO^-) = conjugate base. (Pairs: P/S and R/Q.) [1]

(ii) S is a **relatively strong** (not negligible) base: acetic acid is a weak acid and a weak acid has a strong conjugate base (it readily accepts H^+). [1]

OR

$AlCl_3$ – Lewis acid (electron-deficient, accepts a lone pair); F^- – Lewis base (donates a lone pair); Mg^{2+} – Lewis acid (accepts lone pairs); H_2O – Lewis base (donates a lone pair from oxygen). [½ each]

SECTION C (2 × 3 = 6 marks)

Q17. [3]

(i) $K_c = [CO_2][H_2] / ([CO][H_2O])$ [1]

(ii) At equilibrium: $[CO] = [H_2O] = 0.10 - x$; $[CO_2] = [H_2] = x$. [½]

$K_c = x^2 / (0.10 - x)^2 = 4 \rightarrow x / (0.10 - x) = 2 \rightarrow x = 0.20 - 2x \rightarrow 3x = 0.20 \rightarrow$ **x = 0.067 M** [1]

(iii) $[CO_2] = [H_2] =$ **0.067 M**; $[CO] = [H_2O] = 0.10 - 0.067 =$ **0.033 M**. [½]

Q18. [3]

(i) $K_a = c\alpha^2 \rightarrow \alpha = \sqrt{(K_a/c)} = \sqrt{(10^{-5}/0.1)} = \sqrt{(10^{-4})} =$ **0.01** (1%). [½]

$[H^+] = c\alpha = 0.1 \times 0.01 = 10^{-3}$ M → **pH = 3** [½]

(ii) $HA \rightleftharpoons H^+ + A^-$: $[A^-] \approx 0.1$ M (from NaA), $[HA] \approx 0.1$ M. [½]

$[H^+] = K_a[HA]/[A^-] = 10^{-5} \times 0.1/0.1 =$ **10^{-5} M → pH = 5** [½]

(iii) Common ion effect. The added A^- ions shift the equilibrium $HA \rightleftharpoons H^+ + A^-$ to the left (to consume the added ion), so the ionisation of HA is suppressed and $[H^+]$ falls (pH rises from 3 to 5). [1]

SECTION D (1 × 4 = 4 marks)

Q19. [4]

(i) Solutions which resist a change in pH on dilution or on the addition of small amounts of acid or alkali are called buffer solutions. [1]

(ii) pH = pK_a + log([salt]/[acid]) = 4.75 + log(0.20/0.10) = 4.75 + log 2 = 4.75 + 0.30 = **5.05** [1]

(iii) The added H^+ combines with the acetate ions (the conjugate base): CH_3COO^- + H^+ → CH_3COOH. Since the large reservoir of acetate takes up the H^+, the ratio [salt]/[acid] changes only slightly and the pH remains almost constant. [1]

(iv) NH_4OH (ammonia solution) and NH_4Cl (pK_a of NH_4^+ = 9.25). [½] On dilution the ratio [salt]/[acid] (or [conjugate acid]/[base]) in the logarithmic term remains unchanged, so the pH does not change. [½]

SECTION E (1 × 5 = 5 marks)

Q20. [5]

(i) $AB_2 \rightleftharpoons A^{2+} + 2B^-$; $[A^{2+}] = S$, $[B^-] = 2S$. $K_{sp} = [A^{2+}][B^-]^2 = S(2S)^2 = \mathbf{4S^3}$ [1]

(ii) AB: $K_{sp} = S^2 \rightarrow S = \sqrt{(4 \times 10^{-10})} = \mathbf{2 \times 10^{-5}}$ **M** [1]

AB_2: $4S^3 = 4 \times 10^{-12} \rightarrow S^3 = 10^{-12} \rightarrow S = \mathbf{1 \times 10^{-4}}$ **M** [1]

(iii) AB_2 is more soluble ($10^{-4} > 2 \times 10^{-5}$ M), although its K_{sp} is smaller. [½] No – K_{sp} values can be compared directly only for salts of the same formula type (same ion ratio), because the K_{sp}–S relationship differs. [½]

(iv) $[A^+] = 0.1$ M, so $S = [B^-] = K_{sp}/[A^+] = 4 \times 10^{-10} / 0.1 = \mathbf{4 \times 10^{-9}}$ **M** (the 2×10^{-5} M of A^+ from AB is negligible). [½] **Common ion effect** (solubility decreases). [½]

OR

(i) NaCl – neutral (salt of strong acid and strong base). NH_4Cl – acidic (strong acid + weak base). CH_3COONa – basic (weak acid + strong base). KF – basic (weak acid HF + strong base KOH). [½ each = 2]

(ii) $CH_3COONa \rightarrow CH_3COO^- + Na^+$; $CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$. Acetate is the conjugate base of a weak acid and accepts H^+ from water, producing OH^- ions → $[OH^-] > [H^+]$, pH > 7. [1]

(iii) pH = 7 + ½(4.8 − 4.6) = 7 + 0.1 = **7.1** [1]

(iv) $K_a \times K_b = K_w \rightarrow K_a = 10^{-14} / (2 \times 10^{-5}) = \mathbf{5 \times 10^{-10}}$ [1]

***** END OF SOLUTIONS *****