ACID/BASE PAPER 1A (HL)

1. A
2. B
3. B
4. B
5. A
6. C
7. A
8. C
9. D
10. B
11. A
12. B
13. B
14. A
15. C
16. B
17. B
18. A
19. C
20. C
21. A
22. B
23. B
24. C
25. A
26. B
27. A
28. C
29. B
30. B

ACID/BASE PAPER 1B (HL)

1. (a) $\text{pH} = \text{p}K_a + \log([\text{A}^-]/[\text{HA}])$; $4.50 = 4.76 + \log(\text{Ratio})$ [1]. Ratio = $10^{-0.26} = 0.55$ [1].

(b) $[\text{CH}_3\text{COO}^-] = 0.55 * 0.200 = 0.11$ mol dm^{-3} [1]. Moles needed = $0.11 * 0.500$ dm^3 = 0.055 mol [1]. Mass = $0.055 * 82.03 = 4.51$ g [1].

(c) $\text{H}^+ + \text{CH}_3\text{COO}^- \text{->} \text{CH}_3\text{COOH}$ [1]. $\text{OH}^- + \text{CH}_3\text{COOH} \text{->} \text{CH}_3\text{COO}^- + \text{H}_2\text{O}$ [1].

(d) pH remains unchanged (the ratio of A- to HA is constant) [1]. Buffer capacity decreases (less moles of buffer components present to neutralize) [1].

2. (a) Tangent drawn and half-equivalence identified at 12.5 cm^3 [1]. pH at V=12.5cm^3 is 3.8. Therefore pK_a = 3.8 [1].

(b) $K_a = 10^{-3.8} = 1.58 \times 10^{-4}$ [1].

(c) At equivalence, the solution contains NaA (a weak base salt) [1]. $\text{A}^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HA}(aq) + \text{OH}^-(aq)$ producing a basic pH > 7 [1].

3. (a) $\text{CH}_3\text{NH}_2(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{CH}_3\text{NH}_3^+(aq) + \text{OH}^-(aq)$ [1].

(b) pOH = 14.00 - 11.52 = 2.48. $[\text{OH}^-] = 10^{-2.48} = 3.31 \times 10^{-3}$ mol dm^{-3} [1].

(c) Assume $[\text{OH}^-] \approx [\text{CH}_3\text{NH}_3^+]$ [1]. And assume initial concentration ≈ equilibrium concentration. [1]. $K_b = (3.31\times10^{-3})^2 / 0.050 = 2.19 \times 10^{-4}$ [1].

ACID/BASE PAPER 2 (HL)

1. (a) Moles $\text{NH}_3 = 0.020 \times 0.150 = 0.0030$ mol [1]. Volume HCl = 0.0030 / 0.100 = 0.030 dm^3 (30.0 cm^3) [1].

(b) $K_b = 10^{-4.75} = 1.78\times10^{-5}$ [1]. $[\text{OH}^-] = \sqrt{(1.78\times10^{-5} \times 0.150)} = 1.63\times10^{-3}$ M [1]. pOH = 2.79, pH = 11.21 [1].

(c) Moles HCl added = 0.015 × 0.100 = 0.0015 mol [1]. Moles NH_3 remaining = 0.0030 - 0.0015 = 0.0015 mol [1]. Moles NH_4^+ formed = 0.0015 mol (at half-equivalence!)[1]. pOH = pK_b = 4.75. pH = 9.25 [1].

2. (a) Acidic [1]. $NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$ [1].

(b) Salt hydrolyses to produce OH^-. $K_b = 10^{-14} / 10^{-4.76} = 5.75\times10^{-10}$ [1]. $[OH^-] = \sqrt{(5.75\times10^{-10} \times 0.20)} = 1.07\times10^{-5}$ [1]. pOH = 4.97 [1]. pH = 9.03 [1].

3. (a) $[HA] = [A^-]$, so pH = $pK_a = -\log(1.34\times10^{-5}) = 4.87$ [1].

(b) HA reacting with added OH-: HA becomes 0.045 mol, A- becomes 0.055 mol [1]. pH = 4.87 + log(0.055/0.045) [1]. pH = 4.87 + 0.087 = 4.96 [1].

(c) 0.050 moles (equal to the base component) [1].

4. (a) End point occurs when pK_{in} = pH [1]. pH = $-\log(2.5\times10^{-5}) = 4.60$ [1].

(b) If a system at equilibrium is subjected to change, it will shift to counteract [1]. High $[H^+]$ at pH 2 shifts the equilibrium to the LHS, increasing [HIn] (yellow form) [1].

5. (a) An electron-pair donor [1].

(b) NH_3 is the Lewis base (donates lone pair) [1]. BF_3 is the Lewis acid (accepts electron pair into empty orbital) [1].

(c) Nucleophiles seek positive charge and donate pairs, acting as Lewis bases [1]. Electrophiles seek negative charge and accept pairs, acting as Lewis acids [1].

6. (a) Barium hydroxide + Sulfuric acid -> Barium sulfate + Water [1].

(b) Initial high conductivity due to mobile Ba^{2+} and OH^- [1]. As acid is added, precipitate $BaSO_4$ forms and H^+ neutralizes OH^-, removing mobile ions [1]. Conductivity drops to near zero at the equivalence point [1].

7. (a) $[H^+] = \sqrt{(2.92\times10^{-14})} = 1.71\times10^{-7}$ mol dm^{-3} [1]. pH = $-\log(1.71\times10^{-7}) = 6.77$ [1].

(b) It is still neutral because $[H^+] = [OH^-]$ [1]. Neutrality depends on equality of ions, not pH being exactly 7.0 [1].