Raffles Institution
Year 6 H2 Chemistry 2026
Tutorial 15 – Acid-Base Equilibria

Suggested Answers to Self–Check Questions

1 (a) (i) A Brønsted–Lowry acid is a proton donor while a Brønsted–Lowry base is a proton acceptor.
In this case, an acid–base reaction involves the transfer of a proton from the acid to the base.

(ii) Acid 1 and base 1 constitute one conjugate acid–base pair.
Acid 2 and base 2 constitute another conjugate acid–base pair.

I. $CH_3COOH + H_2SO_4 \rightleftharpoons CH_3COOH_2^+ + HSO_4^-$
base 1 acid 2 acid 1 base 2

II. $CH_3NH_2(aq) + H_2O(l) \rightleftharpoons CH_3NH_3^+(aq) + OH^-(aq)$

III. $HNO_2(aq) + CN^-(aq) \rightleftharpoons HCN(aq) + NO_2^-(aq)$
acid 1 base 2 acid 2 base 1

(b) (i) A Lewis acid is an electron–pair acceptor while a Lewis base is an electron–pair donor.
In this case, an acid–base reaction involves the formation of a dative covalent bond between the acid and the base.

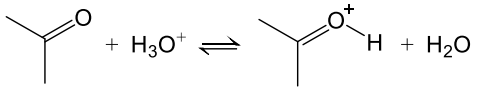
(ii) I. $(CH_3)_3N(g) + BF_3(g) \rightleftharpoons (CH_3)_3NBF_3(s)$
base acid

II. $Al(OH)_3(s) + OH^-(aq) \rightleftharpoons [Al(OH)_4]^-(aq)$
acid base

(c) (i) $HBr(g) + H_2O(l) \rightarrow H_3O^+(aq) + Br^-(aq)$
Brønsted–Lowry acid base
Lewis acid base

(ii) $HIO + NH_2^- \rightleftharpoons NH_3 + IO^-$
Brønsted–Lowry acid base
Lewis acid base

(iii)



Brønsted–Lowry base acid
Lewis base acid

2	**(a)** 0.010 mol dm^{-3} H_2SO_4 *(strong dibasic acid)* $H_2SO_4(aq) \longrightarrow 2H^+(aq) + SO_4^{2-}(aq)$ $[H^+] = 2 \times 0.010 = 0.0200$ mol dm^{-3} pH = –lg (0.0200) = <u>1.70</u>	**(b)** 0.40 g dm^{-3} NaOH *(strong base)* $NaOH(aq) \longrightarrow Na^+(aq) + OH^-(aq)$ $[OH^-] = \frac{0.40}{23.0+16.0+1.0} = 0.0100$ mol dm^{-3} pOH = –lg (0.0100) = 2.00 pH = 14 – 2.00 = <u>12.0</u>

(c) 14.00 cm^3 of 0.10 mol dm^{-3} H_2SO_4 is added to 20.00 cm^3 of 0.10 mol dm^{-3} NaOH *(strong acid partially neutralised by strong base)*

$n_{H^+} = 2 \times \frac{14.00}{1000} \times 0.10 = 0.0028$ mol

$n_{OH^-} = \frac{20.00}{1000} \times 0.10 = 0.002$ mol

n_{H^+} in resultant solution $= 0.0028 - 0.002 = 0.0008$ mol

$\left[H^+\right]$ in resultant solution $= \frac{0.0008}{(14.00 + 20.00)} \times 1000$

$= 0.02353$ mol dm^{-3}

pH of resultant solution = –lg (0.02353) = <u>1.63</u>

(d) 0.30 mol dm^{-3} CH_3CH_2COOH *(weak acid)*

$CH_3CH_2COOH(aq) \rightleftharpoons H^+(aq) + CH_3CH_2COO^-(aq)$

$$K_a = \frac{[H^+][CH_3CH_2COO^-]}{[CH_3CH_2COOH]}$$

$K_a = 10^{-pK_a} = 10^{-4.89} = 1.288 \times 10^{-5}$ mol dm^{-3}

Since CH_3CH_2COOH is a weak acid with a small K_a, $[CH_3CH_2COOH]_{eqm} \approx [CH_3CH_2COOH]_{initial}$ = 0.30 mol dm^{-3}

$\left[H^+\right] = \sqrt{K_a\left[CH_3CH_2COOH\right]} = \sqrt{1.288 \times 10^{-5} \times 0.30}$

$= 1.966 \times 10^{-3}$ mol dm^{-3}

pH = –lg (1.966 x 10^{-3}) = <u>2.71</u>

(e) 2.00 mol dm^{-3} $CH_3CH_2NH_2$ *(weak base)*

$CH_3CH_2NH_2(aq) + H_2O(l) \rightleftharpoons CH_3CH_2NH_3^+(aq) + OH^-(aq)$

$$K_b = \frac{[OH^-][CH_3CH_2NH_3^+]}{[CH_3CH_2NH_2]}$$

Since $CH_3CH_2NH_2$ is a weak base with a small K_b, $[CH_3CH_2NH_2]_{eqm} \approx [CH_3CH_2NH_2]_{initial}$ = 2.00 mol dm^{-3}

$\left[OH^-\right] = \sqrt{K_b\left[CH_3CH_2NH_2\right]} = \sqrt{5.1 \times 10^{-4} \times 2.00}$

$= 0.03194$ mol dm^{-3}

pOH = –lg (0.03194) = 1.496

pH = 14 – 1.496 = <u>12.5</u>

(f) 0.015 mol dm^{-3} $C_6H_5COO^-K^+$ *(basic salt containing $C_6H_5COO^-$ which is the <u>conjugate base</u> of the weak acid C_6H_5COOH)*

$C_6H_5COO^-(aq) + H_2O(l) \rightleftharpoons C_6H_5COOH(aq) + OH^-(aq)$

$$K_b = \frac{[OH^-][C_6H_5COOH]}{[C_6H_5COO^-]}$$

Since $C_6H_5COO^-$ is a weak base with a small K_b, $[C_6H_5COO^-]_{eqm} \approx [C_6H_5COO^-]_{initial}$ = 0.015 mol dm^{-3}

$[OH^-] = \sqrt{K_b[C_6H_5COO^-]}$

$\left[OH^-\right] = \sqrt{\frac{K_w}{K_a}\left[C_6H_5COO^-\right]} = \sqrt{\frac{1.0 \times 10^{-14}}{6.5 \times 10^{-5}} \times 0.015}$

$= 1.519 \times 10^{-6}$ mol dm^{-3}

pOH = –lg (1.519 x 10^{-6}) = 5.818

pH = 14 – 5.818 = <u>8.18</u>

(g) 0.25 mol dm^{-3} methylammonium nitrate $CH_3NH_3NO_3$ *(acidic salt containing $CH_3NH_3^+$ which is the <u>conjugate acid</u> of the weak base CH_3NH_2)*

$CH_3NH_3^+(aq) + H_2O(l) \rightleftharpoons CH_3NH_2(aq) + H_3O^+(aq)$

$$K_a = \frac{[H_3O^+][CH_3NH_2]}{[CH_3NH_3^+]}$$

Since $CH_3NH_3^+$ is a weak acid with a small K_a, $[CH_3NH_3^+]_{eqm} \approx [CH_3NH_3^+]_{initial}$ = 0.25 mol dm^{-3}

$\left[H_3O^+\right] = \sqrt{K_a\left[CH_3NH_3^+\right]}$

$= \sqrt{\frac{K_w}{K_b}\left[CH_3NH_3^+\right]}$

$= \sqrt{\frac{1.0 \times 10^{-14}}{4.4 \times 10^{-4}} \times 0.25}$

$= 2.384 \times 10^{-6}$ mol dm^{-3}

pH = –lg (2.384 x 10^{-6}) = <u>5.62</u>

(h) A solution containing 0.50 mol dm^{-3} CH_3CH_2COOH and 0.35 mol dm^{-3} $CH_3CH_2COO^-K^+$ *(buffer)*

$CH_3CH_2COOH(aq) \rightleftharpoons CH_3CH_2COO^-(aq) + H^+(aq)$

$$pH = pK_a + \lg\frac{\left[CH_3CH_2COO^-\right]}{\left[CH_3CH_2COOH\right]}$$

$$= -\lg\left(1.3 \times 10^{-5}\right) + \lg\left(\frac{0.35}{0.50}\right) = \underline{4.73}$$

(i) A solution containing 0.060 mol dm^{-3} of NH_3 and 0.080 mol dm^{-3} of NH_4Cl *(buffer)*

$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

$$pOH = pK_b + \lg\frac{\left[NH_4^+\right]}{\left[NH_3\right]}$$

$$= -\lg\left(1.74\times10^{-5}\right) + \lg\left(\frac{0.080}{0.060}\right) = 4.884$$

pH = 14 – 4.884 = <u>9.12</u>

3 **(a)** $HX(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + X^-(aq)$

At equilibrium, $[H_3O^+] = [X^-] = 10^{-3.25} = 5.623 \times 10^{-4}$ mol dm^{-3}
$[HX] = 1.00 \times 10^{-3} - 5.623 \times 10^{-4} = 4.377 \times 10^{-4}$ mol dm^{-3}

$$K_a = \frac{\left[H_3O^+\right]\left[X^-\right]}{\left[HX\right]} = \frac{\left(5.623\times10^{-4}\right)^2}{4.377\times10^{-4}} = \underline{7.23 \times 10^{-4} \text{ mol dm}^{-3}}$$

(b) $CH_3COO^-Na^+(aq) \rightarrow CH_3COO^-(aq) + Na^+(aq)$
$CH_3COO^-(aq) + H_2O(l) \rightleftharpoons CH_3COOH(aq) + OH^-(aq)$

At equilibrium, $[CH_3COOH] = [OH^-] = 10^{-(14 - 8.37)} = 2.344 \times 10^{-6}$ mol dm^{-3}
$[CH_3COO^-] = 1.00 \times 10^{-2} - 2.344 \times 10^{-6} = 9.998 \times 10^{-3}$ mol dm^{-3}

$$K_b = \frac{[CH_3COOH][OH^-]}{[CH_3COO^-]} = \frac{(2.344 \times 10^{-6})^2}{9.998 \times 10^{-3}} = \underline{5.497 \times 10^{-10} \text{ mol dm}^{-3}}$$

pK_b of CH_3COO^- = –lg (5.497×10^{-10}) = 9.260
pK_a of CH_3COOH = 14 – 9.260 = <u>4.74</u>

(c) $(NH_4)_2SO_4(aq) \rightarrow 2NH_4^+(aq) + SO_4^{2-}(aq)$
$NH_4^+(aq) + H_2O(l) \rightleftharpoons NH_3(aq) + H_3O^+(aq)$

Initial $[NH_4^+] = (2)(0.050) = 0.100$ mol dm^{-3}

At equilibrium, $[NH_3] = [H_3O^+] = 10^{-5.12} = 7.586 \times 10^{-6}$ mol dm^{-3}
$[NH_4^+] = 0.100 - 7.586 \times 10^{-6} = 0.100$ mol dm^{-3}

$$K_a = \frac{[NH_3][H_3O^+]}{[NH_4^+]} = \frac{(7.586 \times 10^{-6})^2}{0.100} = 5.755 \times 10^{-10} \text{ mol dm}^{-3}$$

pK_a of NH_4^+ = –lg (5.755×10^{-10}) = 9.240
pK_b of NH_3 = 14 – 9.240 = <u>4.76</u>

4 (Ans: **B**)
In the absence of water, HC*l*(g) remains as an undissociated simple molecule, and is not dissociated into H^+ and C*l*$^-$ ions, thus it does not have acidic properties (options A, C and D are incorrect) and does not act as mobile charge carriers (option B is correct).

5 (Ans: **D**)
K_a only changes with temperature. Since the temperature is constant, K_a remains constant as volume changes.

6 (Ans: **D**)
Overall Equation:
$2MnO_4^- + 5SO_2 + 2H_2O \rightarrow 2Mn^{2+} + 5SO_4^{2-} + 4H^+$
Since H^+ is produced, $[H^+]$ increases, causing pH to decrease gradually.

7 (Ans: **B**)
At 25 °C,
pH = – lg $[H^+]$ = – lg (10^{-7}) = <u>7</u> and
pK_w = <u>14</u>

At equilibrium,
$[H^+] = [OH^-] = 10^{-7}$ mol dm^{-3}
$[H_2O]$ = 1000/18 = 55.56 mol dm^{-3} (See note below)

$$K_a = \frac{[H^+][OH^-]}{[H_2O]} = \frac{(10^{-7})^2}{55.56} = 1.800 \times 10^{-16} \text{ mol dm}^{-3}$$

pK_a of H_2O = –lg (1.800×10^{-16}) = <u>15.7</u>

Hence, pH < pK_w < pK_a

Note: How to find $[H_2O]$?
Consider 1 dm^3 of water,
Assume density of water = 1 g cm^{-3}
$\Rightarrow$ mass of 1 dm^3 of H_2O = 1000 g

$$n_{H_2O} = \frac{1000}{2 \times 1.0 + 16.0} = 55.56 \text{ mol}$$

$\therefore [H_2O] = 55.56$ mol dm^{-3}

8 (Ans: **A**)
Pyruvic acid is a weak acid, hence upon titration with NaOH (strong base), at equivalence point, pH > 7 due to anion hydrolysis.

$CH_3COCOO^- + H_2O \rightleftharpoons CH_3COCOOH + OH^-$