

# PRACTICE EXERCISE SOLUTIONS

## Acids, Bases and Acid-Base Equilibria

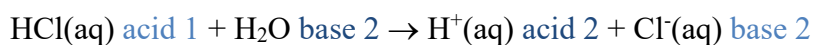
1. Classify the following as Brønsted acid, base or both:

a) HI- Brønsted acid                      b)  $\text{SO}_3^{2-}$  - Brønsted base

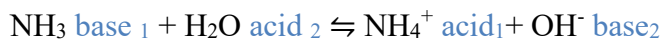
c)  $\text{HCO}_3^-$  - Brønsted acid/base              d)  $\text{NO}_3^-$  - Brønsted base

2. Identify the conjugate acid/base pairs when:

(a) HCl is added to water



(b) Ammonia is added to water



3. Write the formula for the conjugate base for the following compounds,

(a)  $\text{HSO}_4^- / \text{SO}_4^{2-}$                                       (b)  $\text{H}_3\text{PO}_4 / \text{H}_2\text{PO}_4^-$

(c)  $\text{H}_2\text{CO}_3 / \text{HCO}_3^-$                                       (d)  $\text{H}_2\text{O} / \text{OH}^-$

4. Write the formula for the conjugate acid for the following compound,

(a)  $\text{PO}_4^{3-} / \text{HPO}_4^{2-}$                                       (b)  $\text{H}_2\text{O} / \text{H}_3\text{O}^+$

(c)  $\text{CH}_3\text{CH}_2\text{OH} / \text{CH}_3\text{CH}_2\text{OH}_2^+$                       (d)  $\text{SO}_3^{2-} / \text{HSO}_3^-$

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5. Calculate the pH of: (a) 100 mLs of H<sub>2</sub>O



$$K_w = 1.0 \times 10^{-14} = [\text{H}_3\text{O}^+][\text{OH}^-]$$

$$[\text{H}_3\text{O}^+] = \sqrt{1 \times 10^{-14}} = 1 \times 10^{-7} \text{M}$$

$$\text{pH} = -\log(1 \times 10^{-7}) = -(-7) = 7$$

- (b) 100 mL 0.01M HCl + 500 mL H<sub>2</sub>O

$$50 \% \text{ dilution results in new conc of HCl} = \frac{100 \text{ mL} \times 0.01 \text{ M}}{100 \text{ mL} + 500 \text{ mL}} = 0.001667 \text{ M}$$

$$\text{pH} = -\log(0.001667 \text{ M}) = -(-2.778) = \mathbf{2.78}$$

- (c) 100 mL of a 0.05 M NaOH solution

$$[\text{OH}^-] = 5 \times 10^{-2} \text{ M}$$

$$\text{pOH} = -\log(5 \times 10^{-2} \text{ M}) = -(-1.30) = 1.3$$

$$\text{pH} = \text{p}K_w - \text{pOH} = 14 - 1.3 = \mathbf{12.7}$$

- (d) 50 mL of 0.01 M NaOH + 25 mL 0.01 M NaOH

$$\begin{aligned} \text{Total mmols of NaOH} &= (50 \text{ mL} \times 0.05 \text{ M NaOH}) + (25 \text{ mL} \times 0.01 \text{ M NaOH}) \\ &= 5.0 \text{ mmols} \end{aligned}$$

$$\text{Conc of final solution} = \frac{5.0 \text{ mmols}}{50 \text{ mL} + 25 \text{ mL}} = 0.0667 \text{ M}$$

$$[\text{OH}^-] = 0.0667 \text{ M}$$

$$\text{pOH} = -\log(0.0667 \text{ M}) = -(-1.18) = 1.18$$

$$\text{pH} = \text{p}K_w - \text{pOH} = 14 - 1.18 = \mathbf{12.82}$$

- (e) 50 mL of 0.5 M HCl + 50 mL of 0.5 M HCl

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$$\text{Total mmols of HCl} = (50 \text{ mL} \times 0.5 \text{ M HCl} + 50 \text{ mL} \times 0.5 \text{ M HCl})$$

$$= 50 \text{ mmols}$$

$$\text{Conc of final solution} = \frac{50 \text{ mmols}}{50 \text{ mL} + 50 \text{ mL}} = 0.5\text{M}$$

$$[\text{H}^+] = 0.5\text{M}$$

$$\text{pH} = -\log(0.5 \text{ M}) = -(-0.30) = 0.30$$

6. Calculate the  $[\text{H}_3\text{O}^+]$  concentration of a solution that has a pH of :

(a) 10

$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

$$10 = -\log[\text{H}_3\text{O}^+]$$

$$[\text{H}_3\text{O}^+] = \text{antilog } 10, 1/x = \underline{\mathbf{1 \times 10^{-10} \text{M} \text{ or } 10^{-\text{pH}} = 10^{-10} \text{M}}}$$

(b) for pH 2,  $[\text{H}_3\text{O}^+] = 10^{-2}\text{M}$

7. Calculate the  $[\text{OH}^-]$  concentration of a solution that has a pOH of :

(a) 12  $[\text{OH}^-] = 10^{-12}\text{M}$

(c) 3  $[\text{OH}^-] = 10^{-3}\text{M}$

8. In an acid-base titration 25.00mL of a 0.1020M HCl solution was titrated with:

a) 0 mL NaOH

b) 15 mL NaOH

c) 25 mL NaOH

d) 30 mL NaOH

e) 35 mL NaOH

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Calculate the pH of the final solution after each addition of titrant NaOH (0.08050M)

**a) 0 mL NaOH**

Only solution that is considered is 0.1025M HCl

$$\text{Therefore pH} = -\log(0.1025\text{M}) = -0.99$$

**b) 15 mL NaOH**

$$\text{Amount of HCl initially} = 0.1020 \text{ mol dm}^{-3} \times 0.025 \text{ dm}^3 = 0.00255 \text{ mols}$$

$$\text{Amount of NaOH added} = 0.08050 \text{ mol dm}^{-3} \times 0.015 \text{ dm}^3 = 0.001208 \text{ mols}$$

$$\text{Amount of HCl remaining} = 0.00255 - 0.001208 \text{ mols} = 0.001342 \text{ mols}$$

$$\text{Total volume} = 0.025 \text{ dm}^3 + 0.015 \text{ dm}^3 = 0.040 \text{ dm}^3$$

$$\text{Conc of HCl} = \frac{0.001342 \text{ mol}}{0.040 \text{ dm}^3} = 0.03355 \text{ M}$$

$$\text{pH} = -\log(0.03355) = -(-1.47) = 1.47$$

**c). 25 mL NaOH**

$$\text{Amount of HCl initially} = 0.1020 \text{ mol dm}^{-3} \times 0.025 \text{ dm}^3 = 0.00255 \text{ mols}$$

$$\text{Amount of NaOH added} = 0.08050 \text{ mol dm}^{-3} \times 0.025 \text{ dm}^3 = 0.002013 \text{ mols}$$

$$\text{Amount of HCl remaining} = 0.00255 - 0.002013 \text{ mols} = 0.000537 \text{ mols}$$

$$\text{Total volume} = 0.025 \text{ dm}^3 + 0.025 \text{ dm}^3 = 0.050 \text{ dm}^3$$

$$\text{Conc of HCl} = \frac{0.000537 \text{ mol}}{0.050 \text{ dm}^3} = 0.01074 \text{ M}$$

$$\text{pH} = -\log(0.01074) = -(-1.97) = 1.97$$

**d). 30 mL NaOH**

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$$\text{Amount of HCl initially} = 0.1020 \text{ mol dm}^{-3} \times 0.025 \text{ dm}^3 = 0.00255 \text{ mols}$$

$$\text{Amount of NaOH added} = 0.08050 \text{ mol dm}^{-3} \times 0.030 \text{ dm}^3 = 0.002415 \text{ mols}$$

$$\text{Amount of HCl remaining} = 0.00255 - 0.002415 \text{ mols} = 0.000135 \text{ mols}$$

$$\text{Total volume} = 0.025 \text{ dm}^3 + 0.025 \text{ dm}^3 = 0.050 \text{ dm}^3$$

$$\text{Conc of HCl} = \frac{0.000135 \text{ mol}}{0.055 \text{ dm}^3} = 0.02455 \text{ M}$$

$$\text{pH} = -\log(0.02455) = -(-1.61) = 1.61$$

e). **35 mL NaOH**

$$\text{Amount of HCl initially} = 0.1020 \text{ mol dm}^{-3} \times 0.025 \text{ dm}^3 = 0.00255 \text{ mols}$$

$$\text{Amount of NaOH added} = 0.08050 \text{ mol dm}^{-3} \times 0.035 \text{ dm}^3 = 0.002818 \text{ mols}$$

$$\text{Amount of NaOH remaining} = 0.002818 - 0.00255 \text{ mols} = 0.000268 \text{ mols}$$

$$\text{Total volume} = 0.025 \text{ dm}^3 + 0.035 \text{ dm}^3 = 0.060 \text{ dm}^3$$

$$\text{Conc of NaOH} = \frac{0.000268 \text{ mol}}{0.060 \text{ dm}^3} = 0.004467 \text{ M}$$

$$\text{pOH} = -\log(0.004467) = -(-2.35) = 2.35$$

$$\text{pH} = 14 - 2.35 = 11.65$$

9 The pH of a 0.0600 M weak monoprotic acid (HA) is 3.44.

(a) What is the percentage of the acid that is ionized in this solution?

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Ans. pH= 3.44, therefore  $[H^+] = 2.188 \times 10^{-4}M$

$$\begin{aligned} \% \text{ Ionisation} &= \frac{[H^+]}{[HA]} \times 100 \\ &= \frac{3.631 \times 10^{-4} M}{0.0600 M} \\ &= \underline{0.61 \%} \end{aligned}$$

(b) Calculate the  $K_a$  of the acid.



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



Initial	0.0600 M	0	0
Change	$-3.631 \times 10^{-4}M$	$3.631 \times 10^{-4}M$	$3.631 \times 10^{-4}M$
Equilibrium	$0.0600 - 3.631 \times 10^{-4}M$	$3.631 \times 10^{-4}M$	$3.631 \times 10^{-4}M$

Assume:  $0.0600 - 3.631 \times 10^{-4}M \approx 0.0600 M$

$$K_a = \frac{[3.631 \times 10^{-4} M][3.631 \times 10^{-4} M]}{[0.0600 M]}$$

$$= 2.20 \times 10^{-6}$$

10. What is the pH of a 0.40 M ammonia solution?  $K_b = 1.85 \times 10^{-5}$ ?



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$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$



Initial	0.40 M	0	0
Change	-x	+ x	+x
Equilibrium	0.40 M - x	x	x

$$1.8 \times 10^{-5} = \frac{(x)(x)}{0.40-x} \text{ applying the approximation } 0.40 - x \approx 0.4$$

$$= \frac{x^2}{0.40} \Rightarrow x = \sqrt{(0.4)(1.8 \times 10^{-5})} = 2.7 \times 10^{-3} \text{ M}$$

$$\text{pOH} = -\log(2.7 \times 10^{-3} \text{ M}) = \mathbf{2.57}$$

$$\text{pH} = 14 - \text{pOH}$$

$$= 14 - 2.57 = \mathbf{11.43}$$