St. Andrew's Junior College
H1 Chemistry 2022
Physical Chemistry Tutorial
Theories of Acids and Bases

1. Identify the conjugate acid-base pair in the following reaction.

$$HCl + NaHCO_3 \rightarrow H_2CO_3 + NaCl$$

2. Calculate the pH of solutions with the following H_3O^+ concentrations in mol dm^{-3}

(a) 6.80×10^{-3} **(b)** 3.20×10^{-5}

3. Calculate the pH of solutions with the following OH^- concentrations in mol dm^{-3}

(a) 3.70×10^{-10} **(b)** 6.40×10^{-5}

4. Calculate the H_3O^+ and OH^- concentrations in solutions with the following pH values:

(a) 9.21 **(b)** 13.7

5. Calculate the pH and pOH of the following solutions:

(a) 25 cm^3 of 0.200 mol dm^{-3} of hydrochloric acid made up to 500 cm^3 of solution.

(b) 1.00 g of calcium hydroxide dissolved in water and made up to 250 cm^3 of solution

6. Explain the following terms:

(i) pH **(ii)** K_a **(iii)** K_b **(iv)** K_w

7. Calculate the *degree of dissociation* of the weak monobasic acids listed below:

(a) a solution of 0.0100 mol dm^{-3} CH_3COOH has a pH of 3.38.

(b) a solution of 0.200 mol dm^{-3} HCN has a pH of 5.05

8. Calculate the *percentage dissociation* of the weak monoacidic bases listed below:

(a) a solution of 0.0100 mol dm^{-3} CH_3NH_2 has $[OH^-]$ of 4.78×10^{-7} mol dm^{-3}.

(b) a solution of 0.0500 mol dm^{-3} $C_6H_5NH_2$ has $[OH^-]$ of 9.65×10^{-5} mol dm^{-3}

9. **(a)** State if each of the following salts is acidic, basic or neutral.

A – acidic salt ; **B** – basic salt ; **N** – neutral salt

	Salt	**A, B or N**
(i)	KBr	
(ii)	C_6H_5ONa	
(iii)	NaCN	

	Salt	**A, B or N**
(iv)	$(C_6H_5NH_3)_2SO_4$	
(v)	$(CH_3)_2NH_2Br$	
(vi)	CH_3COOK	

(b) **(i)** For $C_6H_5O^-Na^+$, write an equation to illustrate its acidity/basicity.

(ii) For CH_3NH_3Cl, write an equation to illustrate its acidity/basicity.

10. Indicate if the solution containing the specified species is a buffer solution.
If it is, state if it is acidic or basic buffer.
X: not buffer **A**: acidic buffer **B**: basic buffer

	Solution	**X** or **A** or **B**
(a)	CH_3COOH / CH_3COONa	
(b)	CH_3COOH / NaC*l*	
(c)	HC*l* / NaC*l*	
(d)	HNO_3 / $NaNO_3$	
(e)	HCN / KCN	
(f)	NH_3 / NH_4NO_3	
(g)	NaOH / Na_2SO_4	

ANSWERS

2 (a) 2.17 (b) 4.49

3 (a) 4.57 (b) 9.81

4 (a) $[H_3O^+] = 6.17 \times 10^{-10}$ mol dm^{-3}; $[OH^-] = 1.62 \times 10^{-5}$ mol dm^{-3}
(b) $[H_3O^+] = 2.00 \times 10^{-14}$ mol dm^{-3}; $[OH^-] = 0.500$ mol dm^{-3}

5 (a) pH = 2; pOH = 12.
(b) pOH = 0.967; pH = 13.0

7 (a) $\alpha = 0.04$
(b) $\alpha = 4.46 \times 10^{-5}$

8 $\alpha = 4.78 \times 10^{-3}$ %
$\alpha = 0.193$ %

9 (a) (i) N (ii) B (iii) B (iv) A (v) A (vi) B

10 (a) A (b) X (c) X (d) X (e) A (f) B (g) X

Self – Check Answers:

1. Identify the Bronsted acid and its conjugate base and the Bronsted base and its conjugate acid in the following:

HCl	+	NH_3	→	NH_4^+	+	Cl^-
acid		base		conjugate acid		conjugate base

2 Applying pH = - log $[H^+]$ or pH = - log $[H_3O^+]$

(a) 2.17 (b) 4.49

3 Applying pOH = - log $[OH^-]$

Followed by pH = 14 – pH

(a) 4.57 (b) 9.81

4 (a) $[H_3O^+] = 10^{-9.21} = 6.17 \times 10^{-10}$ mol dm^{-3}

$[OH^-] = 10^{-14} / [H_3O^+] = 1.62 \times 10^{-5}$ mol dm^{-3}

(b) $[H_3O^+] = 10^{-13.7} = 2.00 \times 10^{-14}$ mol dm^{-3}

$[OH^-] = 10^{-14} / [H_3O^+] = 0.500$ mol dm^{-3}

5 (a) $[HCl] = [H^+] = [(25 / 1000) \times 0.2] / (500/1000) = 0.01$ mol dm^{-3}

pH = - log $[H^+]$ = 2

pOH = 14 – 2 = 12.

(b) mole of calcium hydroxide = 1.00 / [40.1 + 2(17)] = 0.0135 mol

Mol of OH^- ions = 0.0270

$[OH^-] = 0.0270 / 250 \times 1000 = 0.108$ mol dm^{-3}

pOH = - log $[OH^-]$ = 0.967

pH = 14 – 0.967 = 13.0

6 (i) pH = −lg $[H^+]$

(ii) For a weak acid HA,

$$HA(aq) + H_2O(l) \rightleftharpoons A^-(aq) + H_3O^+(aq)$$

$$K_a = \frac{[H^+(aq)][A^-(aq)]}{[HA(aq)]} \text{ mol dm}^{-3}.$$

(iii) For a weak base B,

$$B(aq) + H_2O(l) \rightleftharpoons BH^+(aq) + OH^-(aq)$$

$$K_b = \frac{[OH^-(aq)][BH^+(aq)]}{[B(aq)]} \text{ mol dm}^{-3}.$$

(iv) $H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$

$K_w = [H^+][OH^-]$ mol^2 dm^{-6}.

7

degree of dissociation, $\alpha = \frac{[H^+]}{[HA]}$

(c) $\alpha = 10^{-3.38} / 0.01 = 0.04$

(d) $\alpha = 10^{-5.05} / 0.2 = 4.46 \times 10^{-5}$

8 $\alpha = \frac{[OH^-]}{[B]} \times 100\ \%$

(a) $\alpha = 4.78 \times 10^{-7} / 0.01 \times 100\% = 4.78 \times 10^{-3}\ \%$

$\alpha = 9.65 \times 10^{-5} / 0.05 \times 100\% = 0.193\ \%$

9 (a) (i) N (ii) B (iii) B (iv) A (v) A (vi) B

(b) (i) $C_6H_5O^- + H_2O \rightleftharpoons C_6H_5OH + OH^-$

(ii) $CH_3N^+H_3 + H_2O \rightleftharpoons CH_3NH_2 + H_3O^+$

10 (a) A (b) X (c) X (d) X (e) A (f) B (g) X

[redacted]

Theories of acids and bases

1. **[2013/P3/Q2]**

(a) Explain what is meant by the terms *Brønsted-Lowry base* and *conjugate acid-base pair*.

Brønsted-Lowry base is a proton acceptor.

In an acid-base reaction, a H^+ is transferred from an acid to a base. For example:

$$HA \rightleftharpoons H^+ + A^-$$

HA and A^- is a conjugate acid-base pair, which differs by one H^+.

Examiner's comments:

Most candidates gave a correct definition of a Bronsted-Lowry base as a proton acceptor. Some defined a Lewis base (electron-pair donor) and a minority defined a Lewis acid. Far fewer were able to offer an adequate description of a conjugate acid-base pair. Several wrote about strong acids being associated with weak bases, or vice versa, without focussing on the fact that the pair differ by the presence or absence of a proton.

(b) The ammonia molecule can react in various ways: as an acid, base, oxidising agent or reducing agent.

Study the following reactions and decide in which way ammonia is reacting in each case. Explain your answers fully.

(i) $NH_3 + H_2O \xrightleftharpoons{\text{room temperature}} NH_4OH$

(ii) $NH_3 + HNO_2 \xrightarrow{\text{warm}} N_2 + 2H_2O$

(iii) $NH_3 + NaH \xrightarrow{33\ ^{o}C} NaNH_2 + H_2$

(iv) $2NH_3 + NaOCl \xrightarrow{\text{room temperature}} N_2H_2 + NaCl + H_2O$

(i) Ammonia is reacting as a base. NH_3 accepts a H^+ from H_2O to form NH_4^+.

(ii) Ammonia is reacting as a reducing agent. NH_3 decreases the oxidation number of N from +3 in HNO_2 to 0 in N_2.

(iii) Ammonia is reacting as an acid. NH_3 loses a H^+ to H^- to form NH_2^-.

(iv) Ammonia is reacting as a reducing agent. NH_3 decreases the oxidation number of Cl from +1 in OCl^- to –1 in Cl^-.

Examiner's comments:

Some candidates correctly assigned the various roles of ammonia. Some candidates did not explain their reasons sufficiently. A number of candidates seemed to have confused oxidation with reduction, or acid with base.

K_w

2 . N2013 / I / 11

The dissociation constant, K_w, for the ionisation of water, $H_2O \rightleftharpoons H^+ + OH^-$, at different temperatures is given below.

Temperature / °C	**K_w / mol^2 dm^{-6}**
0	1.15×10^{-15}
25	1.00×10^{-14}
50	5.50×10^{-14}

What can be deduced from this information?

A Only at 25°C are $[H^+]$ and $[OH^-]$ equal.

B The equilibrium lies furthest to the right at 0°C.

C The forward reaction is exothermic.

D The pH of water decreases as temperature increases.

Ans: D

A: Incorrect.

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

$[H^+] = [OH^-]$ for pure water at any temperature

B: Incorrect. $K = \frac{Kf}{Kb}\frac{Kf}{Kb}$. Comparing K_w at 0^oC and 50^oC, K_w at 0^oC is smaller. This implies that at lower temperature, the backward reaction is favoured. Hence, the equilibrium position lies to the left at 0^oC

C: Incorrect. $K = \frac{Kf}{Kb}\frac{Kf}{Kb}$. Comparing K_w at 0^oC and 50^oC, K_w at 50^oC is larger. This implies that at higher temperature, the forward reaction is favoured. Hence, the forward reaction is endothermic.

D: Correct. $K_w = [H^+][OH^-]$ As temperature increases, K_w increases, implying $[H^+]$ increases. Thus, pH decreases.

Calculations of [H^+] and pH for strong acids and strong bases

3. [2012 P1 Q10]

Stomach juices have a pH of 1.0.

Aspirin is a monobasic (monoprotic) acid represented by H*A* ($K_a = 1 \times 10^{-4}$ mol dm^{-3}) which dissociates into ions H^+ and A^-.

What are the relative concentrations of H^+, A^- and H*A* when aspirin from a tablet enters the stomach?

A $[HA] > [H^+] = [A^-]$

B $[H^+] = [A^-] > [HA]$

C $[H^+] > [A^-] > [HA]$

D $[H^+] > [HA] > [A^-]$

Ans: D

Stomach juice, pH =1

$[H^+] = 0.1$ mol dm^{-3}

Since HA is a weak acid, it dissociates partially.

$$HA \rightleftharpoons H^+ + A^-$$

high $[H^+]$ will suppress the dissociation of aspirin thus $[HA] > [A^-]$

4 **2018/P2 Q7**

Hydrogen chloride dissolves in water to give hydrochloric acid, which is a strong acid.

Hydrogen fluoride dissolves in water to give hydrofluoric acid, which is a weak acid.

(a) **(i)** Explain the meaning of the terms *acid*, *strong acid* and *weak acid*, in terms of the Bronsted-Lowry theoy of acids.

A strong acid dissociates fully / completely in water and is a proton donor.

A weak acid dissociates partially in water and is a proton donor.

(ii) Suggest a reason why hydrochloric acid is a stronger acid than hydrofluoric acid.

Chlorine has a larger atomic size than fluorine. There is less effective orbital overlap between hydrogen and chlorine. The bond strength in HC*l* is weaker and hence the HC*l* bond is more easily broken to release H^+.

(b) **(i)** Define the term pH.

pH = -lg $[H^+]$

(ii) Calculate the pH of a 1.5×10^{-4} mol dm^{-3} solution of hydrochloric acid.

[ANS: 3.82]

pH = -lg (1.5×10^{-4}) = 3.82

(iii) Barium hydroxide is a strong base.

Calculate the pH of a 0.075 mol dm^{-3} solution of barium hydroxide.

[ANS: 13.2]

$[OH^-]$ = 2 x 0.075 = 0.15

pOH = -lg 0.15 = 0.824

pH = 14 – 0.824 = 13.2

5. **J90/P2/2**

Apple juice has a pH of 3.5.

(a) **(i)** Define pH.

pH = - log $[H^+]$

(ii) Calculate the molar concentration of hydrogen ions in apple juice.

[ANS: 3.16 x 10^{-4} mol dm^{-3}]

$[H^+] = 10^{-3.5} = 3.16 \times 10^{-4}$ mol dm^{-3}

Apple juice can be titrated with standard alkali.

A 25.0 cm^3 sample of apple juice was exactly neutralised by 27.5 cm^3 of 0.10 mol dm^{-3} sodium hydroxide using phenolphthalein as indicator.

(b) Assuming that apple juice contains a single acid which is monobasic, calculate the molar concentration of the acid in the juice.

[ANS: 0.110 mol dm^{-3}]

Amt of acid = amt of NaOH = 0.0275 x 0.01 = 0.00275 mol

[acid] = 0.00275 / 0.025 = 0.110 mol dm^{-3}

(c) **(i)** What do you understand by the terms strong and weak when applied to acids?

A strong acid fully dissociates in water whilst a weak acid partially dissociates in water.

(ii) What can you deduce about the strength of the acid in apple juice using the two results you have obtained in (a)(ii) and (b)?

$[H^+] << [HA]$. Hence, apple juice is a weak acid.

(iii) Calculate the degree of dissociation of the acid in apple juice.

[ANS: 2.87 x 10^{-3}]

$3.16 \times 10^{-4} / 0.11 = 2.87 \times 10^{-3}$

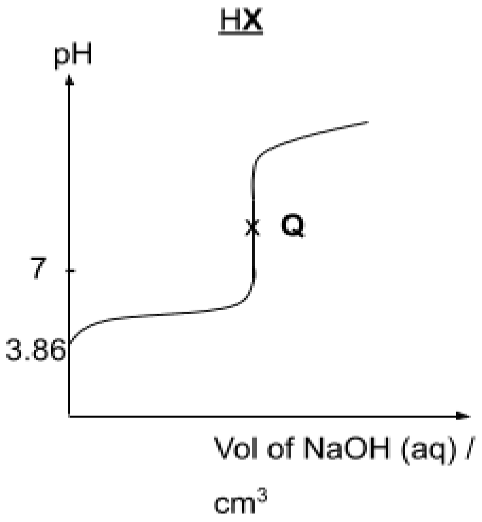
(d) Suggest why phenolphthalein is a suitable indicator for this titration.

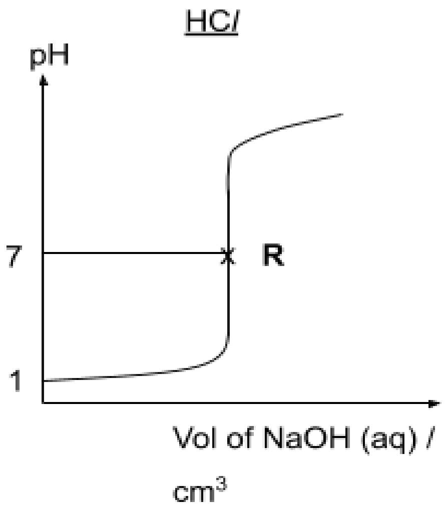
Titration of weak acid with strong base will result in basic salt formed.

Equivalence point > 7 and its vertical section of the titration curve will lie between pH 7 - 10. Hence, the working range of the indicator coincides with the equivalence point/rapid pH change in the titration curve.

Titration Curve

6 A way to determine the acid strength of HX and HC*l* is to titrate with aqueous NaOH. 25.0 cm^3 of 0.100 mol dm^{-3} HX and 25.0 cm^3 of 0.100 mol dm^{-3} HC*l* was pipetted into two different conical flasks. They were titrated separately with 25.0 cm^3 of 0.100 mol dm^{-3} aqueous NaOH in the burette. A pH meter was used to measure the pH change in the conical flask when aqueous NaOH was added. The titration curves of pH against volume of aqueous NaOH added was plotted below for HX and HC*l*.





(a) Given that the initial pH of H**X** and HC*l* is 3.86 and 1 respectively, calculate the initial concentration of H$^+$ ions in H**X** and HC*l*.

[ANS: 1.38 x 10^{-4}; ANS: 0.1.00 mol dm^{-3}]

Initial [H$^+$] in H**X** = 10$^{-3.86}$ = 1.38 x 10^{-4} mol dm^{-3}

Initial [H$^+$] in HC*l* = 10^{-1} = 0.100 mol dm^{-3}

(b) Hence, deduce the acid strength of H**X** and HC*l*.

[H**X**] = [HC*l*] = 0.100 mol dm^{-3}

[H$^+$] in H**X** << [H**X**]

H**X** partially dissociates and hence it is a weak acid.

[H$^+$] in HC*l* = [HC*l*]

HC*l* dissociates fully hence it is a strong acid.

(c) State a suitable indicator for the titration between H**X** and NaOH.

Phenolphthalein

Buffer

7. N 95 / I / 4

"Acidity regulators" are food additives that have a buffering action on the pH of foodstuffs. Mixtures of citric acid and its sodium salt are often used for this purpose.

$C_5H_7O_4CO_2H \rightleftharpoons C_5H_7O_4CO_2^- + H^+$ $K_a = 7.4 \times 10^{-4}$ mol dm^{-3}

(a) With the aid of equations, explain how the citric acid / sodium citrate buffer system regulates the acidity on the addition of

(i) H^+ ions, (ii) OH^- ions

$C_5H_7O_4CO_2H + OH^- \square C_5H_7O_4CO_2^- + H_2O$

The **small amount of OH^- ions** are removed by the **large amount of** $C_5H_7O_4CO_2H$ in the buffer. Therefore pH remains almost unchanged.

$H^+ + C_5H_7O_4CO_2^- \longrightarrow C_5H_7O_4CO_2H$

The **small amount of H^+ ions** are removed by the **large amount** of $C_5H_7O_4CO_2^-$ in the buffer. Therefore pH remains almost unchanged.

(b) Write an expression for the acid dissociation constant, K_a, of citric acid. State the units.

$K_a = [C_5H_7O_4CO_2^-][H^+] / [C_5H_7O_4CO_2H]$

Units: mol dm^{-3}

(c) Define the term K_w and explain why, at 25°C, water has a pH of 7.

$K_w = [H^+][OH^-]$

$[H^+][OH^-] = 1.00 \times 10^{-14}$

Since $[H^+] = [OH^-]$,

$[H^+]^2 = 1.00 \times 10^{-14}$

$[H^+] = \sqrt{1.00 \times 10^{-14}} = 1.00 \times 10^{-7}$ mol dm^{-3}

$pH = -\lg(1.00 \times 10^{-7}) = 7$

8. J91/P3/4(a) & (b)

The water of Lake Nakuru in the Kenyan rift valley contains dissolved sodium carbonate and sodium hydrogencarbonate. The following equilibrium exists:

$$HCO_3^-\,(aq) \rightleftharpoons H^+\,(aq) + CO_3^{2-}\,(aq)$$

(a) Explain how this solution acts as a buffer on the addition of either acid or alkali.

Where there is a small amount of acid H^+ added, the large reservoir of CO_3^{2-} will react with the addition H^+. Hence, pH is maintained.

$$CO_3^{2-} + H^+ \longrightarrow HCO_3^-$$

When there is small amount of alkali OH^- added, the large reservoir of HCO_3^- will react with the additional OH^-. Hence, pH is maintained.

$$HCO_3^- + OH^- \longrightarrow CO_3^{2-} + H_2O$$

(b) **(i)** Write an expression for the acid dissociation constant, K_a, of sodium hydrogen carbonate.

(ii) The pH of Lake Nakuru is 10.3 and the ratio $\dfrac{[CO_3^{2-}(aq)]}{[HCO_3^-(aq)]}$ is 0.958.

Calculate the equilibrium constant for the above reaction using your answer to **(b)(i)**.

[ANS: 4.8 x 10^{-11} mol dm^{-3}]

$$HCO_3^-(aq) \rightleftharpoons H^+(aq) + CO_3^{2-}(aq)$$

$$K_a = \frac{[CO_3^{2-}]\,[H^+]}{[HCO_3^-]} = \frac{[CO_3^{2-}]}{[HCO_3^-]} \times [H^+]$$

$$= 0.958 \times 10^{-10.3} = \mathbf{4.8 \times 10^{-11}\ mol\ dm^{-3}}$$