Year 12 Chemistry **Topic Test #4 (Acids & Bases) - 2013**

Name: **ANSWERS** Mark = _____ / 38

Part One: Multiple Choice Section **10 marks**

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1. **B** 2. **C** 3. **D** 4. **C** 5. **B** 6. **B** 7. **B** 8. **D** 9. **C** 10. **C**

Part Two: Short Answer Section **28 marks**

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11. Information for two acids is as follows:

0.001 mol L^{-1} of $HC\ell O_4$ has a pH = 3; 0.001 mol L^{-1} of $HC\ell O$ has a pH = 4

(a) Using calculations, equations and the above data for the two acids, explain why a solution of $NaC\ell O_4$ is neutral and a solution of $NaC\ell O$ is basic.

[$HC\ell O_4$] = 0.001 mol L^{-1} & [H^+] = 0.001 mol L^{-1} (i.e. pH = 3)
∴ $HC\ell O_4$ has ionised completely and is a strong acid. ✓

As a result, $C\ell O_4^-$ is the conjugate base of a strong acid, so is neutral.
∴ $NaC\ell O_4$ solution will be neutral (pH = 7). ✓

[$HC\ell O$] = 0.001 mol L^{-1} & [H^+] = 0.0001 mol L^{-1} (i.e. pH = 4)
∴ $HC\ell O$ has ionised partially and is a weak acid. ✓

As a result, $C\ell O^-$ is the conjugate base of a weak acid, so is slightly basic.
∴ $NaC\ell O$ solution will be basic (pH > 7), ✓
due to the following hydrolysis reaction:

$$C\ell O^- + H_2O \rightleftarrows HC\ell O + OH^-$$ ✓

(5 marks)

(b) Using the information above, which of the following chemicals could be used as a buffer?

$HC\ell O_4$ $HC\ell O$ $NaC\ell O_4$ $NaC\ell O$

$HC\ell O$ and $NaC\ell O$ ✓

(1 mark)

(c) Explain, with equation(s), what happens when $HC\ell$(aq) is added to this buffer.

H_3O^+ introduced to the buffer is neutralised by reaction with the basic component of the buffer ($C\ell O^-$): ✓

$$C\ell O^-(aq) + H_3O^+(aq) \rightarrow HC\ell O(aq) + H_2O(l)$$ ✓✓

(3 marks)

12. The pH of pure water at 50°C is 6.63.

(a) What is the value of K_W, the equilibrium constant for water, at 50°C? Show all working.

$[H^+] = 10^{-pH} = 10^{-6.63} = 2.344 \times 10^{-7}$ mol L^{-1} ✓

Since water is neutral:
$[OH^-] = [H^+] = 2.344 \times 10^{-7}$ mol L^{-1} ✓

$K_W = [H^+].[OH^-] = (2.344 \times 10^{-7}).(2.344 \times 10^{-7})$
$= \underline{5.50 \times 10^{-14}}$ ✓

(3 marks)

(b) At 50°C, what is the pH of 0.0256 mol L^{-1} NaOH(aq)?

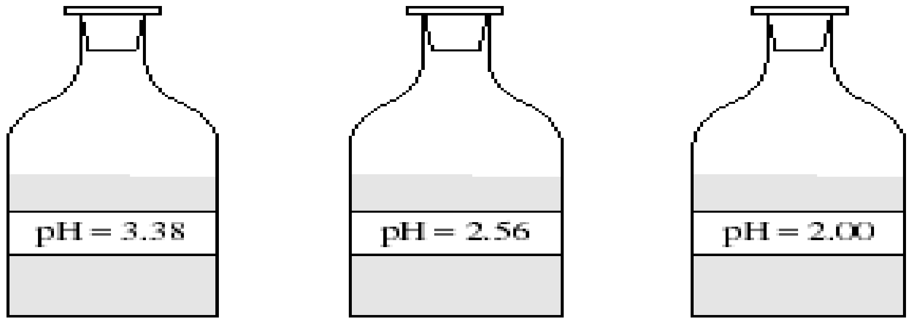
$[OH^-] = [NaOH] = 0.0256$ mol L^{-1} ✓

$[H^+] = K_W/[OH^-] = 5.50 \times 10^{-14}/0.0256 = 2.148 \times 10^{-12}$ mol L^{-1} ✓

$pH = -\log[H^+] = \underline{11.7}$ ✓

(3 marks)

13. The pH values of three acids with the same concentrations at 25°C are shown below. Citric acid is **<u>triprotic</u>**.



0.01 mol L^{-1} methanoic acid	0.01 mol L^{-1} citric acid	0.01 mol L^{-1} hydrobromic acid

(a) 30.0 mL of the three acids were neutralised with 0.02 mol L^{-1} NaOH(aq). If 15.0 mL of NaOH(aq) was needed to neutralise the given HBr(aq), then what volume of NaOH(aq) was needed to neutralise:

(i) methanoic acid **<u>15.0 mL</u>** ✓

(1 mark)

(ii) citric acid **<u>45.0 mL</u>** ✓✓

(2 marks)

(b) Consider these indicators and their end points:

phenolphthalein	(pH range 8.3 to 10)	phenol red	(pH range 7 to 8)
bromothymol blue	(pH range 6.0 to 7.6)	methyl red	(pH range 4.4 to 6.3)

Which indicator would be best for the titration between methanoic acid and sodium hydroxide? Using equations, explain your choice.

The equivalence point for the reaction between a weak acid and strong base would be in basic region ✓

due to the hydrolysis of the methanoate ion ($HCOO^-$):

$$\mathbf{HCOO^-(aq) + H_2O(l) \rightarrow HCOOH(aq) + OH^-(aq)}$$ ✓

The end point of the indicator must match (or be close to) the equivalence point of the reaction, thus phenolphthalein ✓

(3 marks)

(c) List the three acids in increasing order of electrical conductivity

methanoic acid > citric acid > hydrobromic acid ✓

(1 mark)

(d) What is the $[OH^-]$ present in citric acid at 25°C?

pOH = 14 – pH = 11.44 ✓

$[OH^-] = 10^{-pOH} = 10^{-11.44}$ = <u>3.63×10^{-12} mol L^{-1}</u> ✓

or **$[H^+] = 10^{-pH} = 10^{-2.56} = 2.754 \times 10^{-3}$ mol L^{-1}** ✓

$[OH^-] = K_W/[H^+] = 1 \times 10^{-14}/2.754 \times 10^{-3}$ = <u>3.63×10^{-12} mol L^{-1}</u> ✓

(2 marks)

14. When sodium dihydrogenphosphate solution is added to pond water, the pH of the water decreased from 7.83 to 5.56 but when sodium hydrogenphosphate was added the pH increased.

Using equation(s), explain these observations.

Sodium dihydrogenphosphate (NaH_2PO_4) dissociates to produce Na^+ and $H_2PO_4^-$.

Since pH < 7, $H_2PO_4^-$ ions must hydrolyse to produce an acidic solution: ✓

$$H_2PO_4^-(aq) + H_2O(l) \rightarrow H_3O^+(aq) + HPO_4^{2-}(aq)$$ ✓

Sodium hydrogenphosphate (Na_2HPO_4) dissociates to produce Na^+ and HPO_4^{2-}.

Since pH > 7, HPO_4^{2-} ions hydrolyses to produce a basic solution: ✓

$$HPO_4^{2-}(aq) + H_2O(l) \rightarrow OH^-(aq) + H_2PO_4^-(aq)$$ ✓

(4 marks)

End of Test