ACID BASE PRACTICE QUESTIONS 3 - ANSWERS

1. Solutions of lithium carbonate, sodium chloride and ammonium sulfate have their pH tested.
   Which of the following is the correct classification from the test?

	$Li_2CO_3(aq)$	$NaCl(aq)$	$(NH_4)_2SO_4(aq)$
(a)	acidic	neutral	basic
(b)	basic	neutral	acidic
(c)	neutral	acidic	basic
(d)	neutral	acidic	neutral

2. In which of the following examples is the underlined substance acting as a base?

(a) $\underline{NH_3}$ (aq) + H_2O (l) $\quad NH_4^+$ (aq) + OH^- (aq)

(b) $\underline{H_2O}$ (l) + O^{2-} (aq) $\quad 2OH^-$ (aq)

(c) $\underline{2K}$(s) + $2H_2O$ (l) $\quad 2K^+$ (aq) + $2OH^-$ (aq) + H_2 (g)

(d) $\underline{H_2O}$ (l) + CO_3^{2-} (aq) $\quad OH^-$ (aq) + HCO_3^- (aq)

3. For the titration between dilute ethanoic acid (in a burette) and standardised sodium hydroxide in a conical flask, which of the following procedures is incorrect?

(a) Prior to adding the acid to the burette, rinse the burette with distilled water and then a small portion of the acid solution.

(b) Pipette out 20.00 mL aliquots of the sodium hydroxide solution into three separate conical flasks which have each been rinsed with distilled water.

(c) Rinse the pipette with the standardised sodium hydroxide solution before transferring the first aliquot to the conical flask.

(d) Add a few drops of methyl orange indicator to the acid in the burette.

4. A student requires a solution of pH = 4 for an experiment.
Which of the following procedures will yield such a solution?

(a) Add 50 mL of a solution of pH = 2 to 50 mL of a solution of pH = 6.

(b) Add 95 mL of distilled water to 5 mL of 0.1 mol L^{-1} HCl(aq).

(c) Add 50 mL of 0.1 mol L^{-1} NaOH to 25 mL of 0.1 mol L^{-1} HCl(aq).

(d) Add 1.0 mL of 0.01 mol L^{-1} HCl(aq) to 99 mL of distilled water.

5. A chemist added 20.0 mL of 0.0010 mol L^{-1} hydrochloric acid to 100.0 mL of 0.100 mol L^{-1} potassium chloride solution. Which one of the following is the correct pH of the resulting solution?

(a) 2.6

(b) 3.0

(c) 3.8

(d) 5.2

6. Equal volumes of 0.1 mol L^{-1} solutions of ethanoic acid and sodium ethanoate are mixed. Which of the following statements about the resulting solution is true?

(a) The concentration of both hydrogen ions and ethanoate ions is high.

(b) The concentration of both hydrogen ions and ethanoate ions is low.

(c) The concentration of hydrogen ions is higher than the concentration of ethanoate ions.

(d) The concentration of hydrogen ions is lower than the concentration of ethanoate ions.

7. Which of the following statements best describes 10 mol L^{-1} ammonia solution?

(a) A concentrated solution of a weak base.

(b) A dilute solution of a strong base.

(c) A concentrated solution of a strong base.

(d) A dilute solution of a strong electrolyte.

8. Which of the following examples represents an acid-base reaction?

(a) $NH_4^+(aq) + OH^-(aq) \rightarrow NH_3(aq) + H_2O(l)$
(b) $2NO_3^-(aq) + 2H^+(aq) + 3H_2O_2(aq) \rightarrow 2NO(g) + 3O_2(g) + 4H_2O(l)$
(c) $2K(s) + 2H_2O(l) \quad 2K^+(aq) + 2OH^-(aq) + H_2(g)$
(d) $Ca^{2+}(aq) + CO_3^{2-}(aq) \quad CaCO_3(s)$

9. For the titration between dilute ethanoic acid (in a burette) and standardised sodium hydroxide in a conical flask, which of the following procedures is incorrect?

(a) Prior to adding the acid to the burette, rinse the burette with distilled water and then a small portion of the acid solution.

(b) Pipette out 20.00 mL aliquots of the sodium hydroxide solution into three separate conical flasks which have each been rinsed with the sodium hydroxide solution.

(c) Rinse the pipette with the standardised sodium hydroxide solution before transferring the first aliquot to the conical flask.

(d) Add a few drops of phenolphthalein to each of the conical flasks containing the sodium hydroxide aliquots.

10. Complete the following table.

Solute (0.1 mol L^{-1})	pH of solution (< 7, = 7 or > 7)	Explanation
Potassium chloride		
Ammonium chloride		
Sodium carbonate		

[6 marks]

KCl	pH = 7	salts of strong acids/bases or K^+ and Cl^- ions have no tendency to react with water.
NH_4Cl	pH < 7	$NH_4^+(aq) + H_2O(l) \leftrightarrow NH_3(aq) + H_3O^+(aq)$ Hydrolysis of $NH_4^+(aq)$ produces $H_3O^+(aq)$ which reduces pH below 7.
Na_2CO_3	pH > 7	$CO_3^{2-}(aq) + H_2O(l) \rightarrow HCO_3^-(aq) + OH^-(aq)$ Hydrolysis of CO_3^{2-} produces $OH^-(aq)$ which increases pH above 7.

11. Iron is often found in the earth's crust as a hydrated iron oxide ($Fe_2O_3.xH_2O$).

A 0.668 gram sample of this ore was dissolved in excess sulfuric acid and all the iron was converted into an iron(II)sulfate solution. This solution was titrated against a 0.050 mol L^{-1} acidified potassium permanganate solution and exactly 25.00 mL of the purple solution was needed to complete the titration. Find the value of x in the hydrated formula above. [9 marks]

$5Fe^{2+} + MnO_4^- + 8H^+ \rightarrow 5Fe^{3+} + Mn^{2+} + 4H_2O$ (1)

$n(MnO_4^-) = n(KMnO_4) = c \times V = 0.05 \times 0.025 = 0.00125$ mol MnO_4^- (1)

from the equation, $n(Fe^{2+}) = 5 \times n(MnO_4^-) = 5 \times).00125 = 0.00625$ mol Fe^{2+} (1)

$n(Fe) = n(Fe^{2+}) = 0.00625$ mol Fe

$n(Fe_2O_3) = ½\, n(Fe) = 0.003125$ mol Fe_2O_3 (1)

$m(Fe_2O_3) = n \times M = 0.003125 \times 159.7 = 0.4990625$ g Fe_2O_3 (1)

$m(H_2O) = m(\text{sample}) - m(Fe_2O_3) = 0.6680 - 0.4990625 = 0.1689$ g H_2O (1)

$n(H_2O) = m/M = 0.1689 / 18.016 = 0.009377$ mol H_2O (1)

Hence, $n(H_2O) / n(Fe_2O_3) = 0.009377 / 0.003125 = 3 : 1$ (1)

<u>Hence, x = 3</u> (1)

12. A group of students was given the task of determining the percentage of ethanoic acid in a sample of vinegar. To begin with, they placed a sample of sodium carbonate in an oven at 110ºC for twenty four hours. They took a 9.700 gram sample of this sodium carbonate and using distilled water, dissolved it and transferred it into a 500 mL volumetric flask and made it up to the mark. 20.00 mL aliquots of this standard solution were titrated against a HCl solution and the following results were obtained:

Trial	1	2	3	4
Volume of HCl (mL)	20.20	19.90	19.85	19.80

A solution of sodium hydroxide was prepared and standardised against this HCl solution. 25.00 mL aliquots of the HCl solution required an average of 17.50 mL of the sodium hydroxide solution to complete the titration.

25.00 mL of the vinegar solution was transferred to a 250 mL volumetric flask and made up to the mark with distilled water. 20.00 mL aliquots of the dilute vinegar solution were titrated against the standardised sodium hydroxide solution. The following results were obtained:

Trial	1	2	3	4	5
Initial reading (mL)	0.15	3.75	7.15	10.15	13.15
Final reading (mL)	3.75	7.15	10.15	13.15	16.15

If the density of the pure vinegar was 1.02 g cm^{-3}, determine the percentage by mass of ethanoic acid in the pure vinegar. [12 marks]

The reaction is : $2H^+(aq) + CO_3^{2-}(aq) \rightarrow H_2O(l) + CO_2(g)$

$n(CO_3^{2-}) = n(Na_2CO_3) = m/M = 9.7 / 105.99 = 0.091518$ mol CO_3^{2-}

$c(CO_3^{2-}) = c(Na_2CO_3) = n / V = 0.091518 / 0.5 = 0.183036$ mol L^{-1}

$n(CO_3^{2-})$ in 20.0 mL $= c \times V = 0.183036 \times 0.02 = 0.003661$ mol CO_3^{2-} (2)

from the equation, $n(HCl) = 2 \times n(CO_3^{2-}) = 2 \times 0.0036601 = 0.007321$ mol HCl

$c(HCl) = n / V = 0.0073214 / 0.01985 = 0.36884$ mol L^{-1} HCl (2)

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

$n(H^+) = n(HCl) = c \times V = 0.36884 \times 0.025 = 0.009221$ mol H^+

from the equation, $n(OH^-) = n(H^+) = 0.009221$ mol OH^-

$c(NaOH) = c(OH^-) = n / V = 0.009221 / 0.0175 = 0.5269\ mol\ L^{-1}\ NaOH$ (2)

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

$n(OH^-) = n(NaOH) = c \times V = 0.5269 \times 0.003 = 0.0015807\ mol\ OH^-$

from the equation : $n(CH_3COOH) = n(OH^-) = 0.0015807\ mol\ CH_3COOH$ (2)

$c(CH_3COOH)$ in 20.0 mL $= n / V = 0.0015807 / 0.02 = 0.079037\ mol\ L^{-1}\ CH_3COOH$

$n(CH_3COOH)$ in 250 mL vol. flask $= c \times V = 0.079038 \times 0.25 = 0.019759\ mol\ CH_3COOH$ (1)

$n(CH_3COOH)$ in 25 mL pure vinegar $= n(CH_3COOH)$ in 250 mL flask $= 0.019759\ mol\ CH_3COOH$

$m(CH_3COOH) = n \times M = 0.019759 \times 60.052 = 1.18658\ g\ CH_3COOH$ (1)

$m(CH_3COOH\ soln.) = r \times V = 1.02 \times 25 = 25.5\ g.$ (1)

$\% (CH_3COOH) = m(pure) / m(impure) \times 100 = [1.18658 / 25.5] \times 100 = 4.65325$

Hence the solution is 4.65 % CH_3COOH by mass (1)

1. Human blood uses effective chemical reactions which enable it to resist changes in pH above or below its normal value of 7.4. If 10 mL of 1.00 mol L^{-1} HCl(aq) is added to 1.00 L of blood, the pH drops by only 0.2 units. This happens mainly due to equilibrium reactions involving carbonic acid (a weak acid) and hydrogencarbonate ion.

The three related equilibrium reactions which help to regulate blood pH are as follows:

(i) Gaseous carbon dioxide in the lungs is in equilibrium with dissolved carbon dioxide in the blood:

$CO_2(g)$ $\quad$ $CO_2(aq)$……………………..….......................Equation (1)

(ii) The dissolved carbon dioxide reacts with water in the blood forming carbonic acid:

$CO_2(aq)$ + $H_2O(l)$ $\quad$ $CO_3(aq)$………........................…...Equation (2)

(iii) The weak acid carbonic acid is also in equilibrium with hydrogencarbonate ion and hydronium ion:

$H_2CO_3(aq)$ + $H_2O(l)$ $\quad$ $HCO_3^-(aq)$ + $H_3O^+(aq)$……...Equation (3)

Using the above reactions, and your knowledge of Le Chatelier's Principle and/or acid-base theory:

(a) Explain how an increase in $[H_3O^+(aq)]$ affects all three reactions. (6)

(b) Discuss the **position** of equilibrium in Equation (3). (4)

(c) Explain how hydrogencarbonate ion helps the pH of the blood to remain at 7.4. (4)

(d) **Compare** the effect of adding 10 mL of 1.00 mol L^{-1} HCl(aq) to a litre of blood (assume pH = 7) with the effect of adding the same amount of the acid to a litre of pure water (pH = 7). (6)

1(a): In equation ...(3) as given, if $[H_3O^+]$ is increased, LCP predicts that $\leftarrow$ (reverse reaction) is favoured.

This will increase $[H_2CO_3(aq)]$.

In equation ....(2) as given, if $[H_2CO_3(aq)]$ is increased, LCP predicts that $\leftarrow$ will be favoured.

This will increase $[CO_2(aq)]$.

In equation ...(1) as given, if $[CO_2(aq)]$ is increased, LCP predicts that $\leftarrow$ will be favoured.

This will increase $[CO_2(g)]$ in the lungs.

1(b): In equation ...(3), since $H_2CO_3(aq)$ is a weak acid, the position of equilibrium will be to the left i.e. there will be a high concentration of dissolved molecules, but only a low concentration of ions in the solution.

1(c): The pH of the blood is 7.4. Hence, $[H_3O^+(aq)] = 4.0 \times 10^{-8}$ mol L^{-1} (Slightly below 10^{-7} mol L^{-1})

If $[H_3O^+(aq)]$ is increased, Le Chatelier's Principle predicts that this will consume $H_3O^+(aq)$, thus producing a higher concentration of neutral H_2CO_3 molecules. This will increase the pH back to around its original value.

If $[H_3O^+(aq)]$ is decreased, Le Chatelier's Principle predicts that the weak acid $H_2CO_3(aq)$ will ionise to produce more $HCO_3^-(aq)$ and $H_3O^+(aq)$. This will lower the pH back to around its original value.

1(d): If 10 mL of 1.00 mol L^{-1} HCl is added to 1.0 L of water:

$n(H^+(aq)) = n(HCl) = c.V = (1.00 \text{ mol } L^{-1})(0.010 \text{ L}) = 0.010$ mol

$[H^+(aq)] = n/V = (0.010 \text{ mol}) / (1.01 \text{ L}) = 9.90 \times 10^{-3}$ mol L^{-1}

Hence, pH $= - \log (9.90 \times 10^{-3}) = 2.00$

Clearly, if 10 mL of 1.00 mol L^{-1} HCl(aq) is added to 1.00 L of pure water, the pH drops from 7 to 2 (a drop of 5 units) whereas for a similar volume of blood, the pH drops by only 0.2 units!!

Thus, blood has an effective chemical means of minimizing changes to its pH!