CHEM1011 WORKSHEET B

SAMPLE EXAM QUESTIONS

ACIDS, BASES AND pH

Question 1.
(a) Complete the following table of conjugate acid/base pairs.

Conjugate acids have one more 'H' than the conjugate base!

Acid	Base
HF	*F^-*
H_3PO_4	$H_2PO_4^-$
H_2O	*OH^-*
HNO_2	NO_2^-
CO_3^{2-}	— *(there isn't one)*

(b) Calculate the pH of a 0.15 M solution of vitamin B_5 (pantothenic acid) given $K_a = 3.95 \times 10^{-5}$.

$K_a = \frac{x^2}{[HA]}$ *where* $x = [H^+]$

Check: the least accurate value determines the number of significant figures!

rearrange: $x = (3.95 \times 10^{-5} \times 0.15)^{1/2}$

so: $pH = -\log [H^+] = 2.61$ *(answer to 2 significant figures)*

(c) A buffer is made up to contain $[NH_4^+] = 0.2500$ M and $[NH_3] = 0.1234$ M.
Given that pK_a $(NH_4^+) = 9.24$, calculate the pH of the buffer.

Use the Henderson-Hasselbalch equation

$$pH = pK_a + \log \frac{[A^-]}{[HA]}$$

$= 9.24 + \log (0.1234/0.2500) = 8.93$ *(answer to 2 significant figures)*

Check: if more of acid than base, then pH is lower than pK$_a$!

Question 2.

(a) Calculate the pH of an aqueous buffer solution that contains 0.06 M Na_2HPO_4 and 0.12 M KH_2PO_4, given that $pK_a(2)$ for phosphoric acid is 7.21.

Check: if more of acid than base, then pH is <u>lower</u> than pK_a!

$$pH = pK_a + \log\frac{[A^-]}{[HA]}$$

$= 7.21 + \log(0.06/0.12) = 6.91$

(b) Calculate the pH and pH change when 0.8 g of NaOH is added to 400 mL to the above solution.

NaOH neutralises the acid according to:

$$H_2PO_4^- + OH^- \rightarrow HPO_4^{2-} + H_2O$$

Need to calculate new conc's of buffer components.

Initial moles:

$n(H_2PO_4^-) = c \times V = 0.12\ mol\ L^{-1} \times (400mL/1000mL/L) = 48\ mmol$

$n(HPO_4^{2-}) = {}^1/_2 \times n(H_2PO_4^-) = 24\ mmol$ *(Note: ½ as much $H_2PO_4^-$ as HPO_4^{2-})*

Change in moles (due to NaOH):

$n(NaOH) = m(NaOH)/M(NaOH) = 0.8\ g/40\ g\ mol^{-1} = 20\ mmol$

Final moles

$n(H_2PO_4^-) = 48 - 20 = 28\ mmol$

$n(HPO_4^{2-}) = 24 + 20 = 44\ mmol$

So final conc's: *(Note: 1 mmol = 1/1000 mol so the 1000's cancel)*

$c(H_2PO_4^-) = n/V = 28/1000\ mol / (400mL/1000mL/L) = 0.0700\ molL^{-1}$

$c(HPO_4^{2-}) = n/V = 44/1000\ mol / (400mL/1000mL/L) = 0.1100\ molL^{-1}$

Therefore:

$$pH = pK_a + \log\frac{[A^-]}{[HA]}$$

$= 7.21 + \log(0.1100/0.0700) = 7.41$

Check: if more base than acid, then pH is <u>higher</u> than pK_a!

And |pH change| = 7.41 – 6.91 = 0.50 pH units

(c) Briefly explain why the pH change is so small.

The added base reacts with $H_2PO_4^-$ (the <u>acid</u> in the buffer) and, therefore, the concentration of $H^+(aq)$ is <u>almost</u> unchanged.

Question 3.

(a) Fill in the missing entries in the table below for an aqueous solution at 25 °C.

$[H^+]$ /mol L^{-1}	pH	$[OH^-]$ /mol L^{-1}	pOH
$= 10^{-pH}$ $= 3.5 \times 10^{-3}$	$= 14 - pOH$ $= 2.5$	$= 10^{-pOH}$ $= 3.2 \times 10^{-12}$	11.5

or $= K_w/[H^+] = 10^{-14}/3.5 \times 10^{-3}$

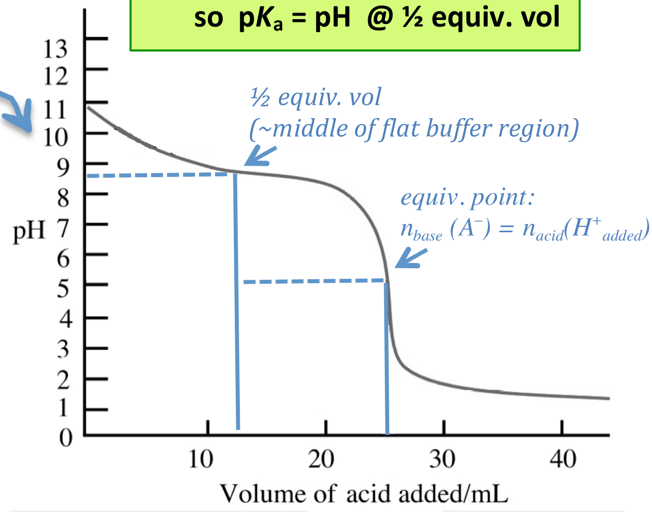
$= 2.9 \times 10^{-12}$ *(so precise answers will depend on fig's carried & route taken)*

(b) The following questions refer to the titration curve shown below right.

(i) Estimate the pK_a of the substance being titrated.

at pK_a = pH, $[A^-]$ = [HA]
so pK_a = pH @ ½ equiv. vol

pK_a = pH @ ½ equiv. vol ~ 8.7
(any number a bit less than 9 would do)



(ii) What class of titration does this curve typify?
Circle your answer in the list below.

strong base/strong acid

strong base/weak acid

weak base/strong acid

weak base/weak acid

(c) pK_w = 11.0 for superheated H_2O at 250 °C. Calculate the pH of neutral water at this temperature.

At neutral pH: $[H^+]$ = $[OH^-]$

$$H_2O(l) \overset{K_w}{==} H^+(aq) + OH^-(aq)$$

In neutral H_2O: $[H^+] = [OH^-]$

so: $K_w = [H^+][OH^-] = [H^+]^2$

and $pK_w = -\log([H^+]^2) = -2\log[H^+] = 2 \times pH$

Hence: $pH = ½\, pK_w = ½ \times 11.0 = 5.5$ *for neutral water at 250 °C*