3.7 Aqueous Solutions 2005 Answers

QUESTION ONE: ACIDS AND BASES

Order of increasing pH: HCl, NH_4Cl, NaCl, NH_3, NaOH

HCl: strong acid, fully dissociates giving high concentration of $H_3O^+(aq)$, i.e. low pH.

$$HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

NH_4Cl: dissolves to give a solution of $NH_4^+(aq)$ and $Cl^-(aq)$ ions. The $NH_4^+(aq)$ ions are weakly acidic and partially dissociate in water to give a small increase in $[H_3O^+(aq)]$.

$$NH_4^+(aq) + H_2O(aq) \rightleftharpoons NH_3(aq) + H_3O^+(aq)$$

NaCl: dissolves to give $Na^+(aq)$ and $Cl^-(aq)$ ions, but neither of these have any reaction with water so solution is neutral.

NH_3: is a weak base and reacts with water to give a small increase in $OH^-(aq)$ ions, making the solution weakly alkaline.

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

NaOH: is a strong base and fully dissociates, giving a high concentration of $OH^-(aq)$ ions, i.e. high pH.

$$NaOH(s) \rightarrow Na^+(aq) + OH^- (aq)$$

A = Order of substances correct with no justification. OR Correct order but reversed pH, with justification of at least one weak and one strong species (may have arrow or equilibrium signs if equation written).
OR Places four species in order, with correct justification and equations for any two.

M = Places ALL the substances in the correct order, with valid discussion for at least three species and at least two equations involving one weak and one strong example.

E = Full discussion linking ALL substances with their relative pH and including equations for NH_3, NH_4Cl, and one other. ***Must use appropriate arrows or equilibrium signs, and CLEARLY state reasons for high or low concentration of H_3O^+ or OH^- ions.***

QUESTION TWO: PRECIPITATING SILVER CHLORIDE

a Solubility is the amount of substance that will dissolve in a given volume to form a saturated solution (at that temperature).

A = Definition clearly identifies that it is the maximum amount that will dissolve in the given volume OR amount dissolved in a saturated solution OR concentration that ions can reach before a precipitate forms.

b K_s is an equilibrium constant for the reaction $AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$
and $K_s = [Ag^+(aq)][Cl^-(aq)]$

More solid dissolves when the temperature is increased as the equilibrium shifts in the endothermic direction, which means it shifts to the right and increases the concentration of ions in solution. This increases K_s.

A = Answer recognises that increased K_s value means increased solubility or concentration of ions.
M = Answer correctly links increase in K_s to the shift in equilibrium for the equation they have written, resulting in increased solubility OR discusses shift in endothermic direction and links it to increased concentration of ions / solubility / shift in given equation.

c (i) $[Ag^+] = [Cl^-] =$ *solubility, s*
$K_s = s^2$
so $s = \sqrt{K_s}$
$= 1.25 \times 10^{-5}$ mol L^{-1}

(ii) $Ks = [Ag^+]^2[CrO_4^{2-}]$
so $s = \sqrt[3]{K_s/4}$
$= 6.88 \times 10^{-5}$ mol L^{-1}

. **A = Solubility of EITHER AgCl OR Ag_2CrO_4 calculated correctly.**
M = Correct process for calculation of solubility of BOTH AgCl and Ag_2CrO_4, plus ONE value of solubility correct.
E = Calculations correctly show that $Ag2CrO_4$ is more soluble than AgCl with correct units of mol L^{-1}.

d $K_s(Ag_2CrO_4) = [Ag^+]^2\,[CrO_4^{2-}]$
$= 1.30 \times 10^{-12}$

$$[Ag^+]^2 = \frac{1.30 \times 10^{-12}}{6.3 \times 10^{-2}}$$
$= 2.06 \times 10^{-10}$

$[Ag^+] = 1.44 \times 10^{-5}$ mol L^{-1}

A = Correct process for the calculation OR Calculation to get $[Ag^+]^2 = 2.06 \times 10^{-10}$.
M = $[Ag^+]$ correctly calculated.

QUESTION THREE: ETHANOIC ACID SOLUTIONS

a In a buffer solution the pH hardly changes when acid or base **is added**. In this case addition of **acid** results in the reaction:

$$CH_3COO^- + H_3O^+ \rightarrow CH_3COOH + H_2O$$

which means there is essentially no change in the concentration of H_3O^+.

When **OH^-** is added it reacts as follows:

$$CH_3COOH + OH^- \rightarrow CH_3COO^- + H_2O$$

and there is therefore almost no change in concentration of OH^- in the solution.

A = Answer recognises that CH_3COOH reacts with added base, and CH_3COO^- reacts with added acid.
OR correctly justified buffering action in one direction with equation (added acid or added base only).
M = Explanation of how this mixture acts as a buffer includes balanced equations for both reactions shown OR discussion relating to shifting the equilibrium on a single equation,

$$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$$

b $Ka = 1.76 \times 10^{-5}$

$$= \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]}$$

$[H_3O^+]$ = inv log (–5.00)

$= 1 \times 10^{-5}$

$[CH_3COOH] = 0.0500 \text{ mol L}^{-1}$

$$[CH_3COO^-] = \frac{K_a\,[CH\ 3CO\ COOH]}{[H_3O^+]}$$

$= 0.0880 \text{ mol L}^{-1}$

A = Correct process for calculation.
M = Correct answer for concentration of CH_3COO^-.

QUESTION FOUR: ANALYSIS OF A WEAK ACID

a $HPab + H_2O \rightleftharpoons H_3O^+ + Pab^-$

b $$K_a = \frac{[H_3O^+][Pab^-]}{[HPab]}$$

A = Correct equilibrium equation OR equation for *K*a.

c $[H_3O^+] = 6.03 \times 10^{-4} \text{ mol L}^{-1}$

A = Correct answer.

d **i** $n(OH^-) = 0.0120 \text{ L} \times 0.050 \text{ mol L}^{-1}$
$= 6.00 \times 10^{-4}$ mol

$$[HPab] = \frac{n}{V}$$
$$= \frac{6.00 \times 10^{-4} \text{ mol}}{0.0200 \text{ mol L}^{-1}}$$
$$= 3.00 \times 10^{-2} \text{ mol L}^{-1}$$
$$= 0.0300 \text{ mol L}^{-1}$$

ii $$Ka = \frac{[H_3O^+][Pab^-]}{[HPab]}$$

$[H_3O^+]=[Pab^-]$ and $[HPab]= 0.0300\text{mol L}^{-1}$

$$K_a = \frac{(6.03 \times 10^{-4})^2}{0.0300}$$
$$= 1.21 \times 10^{-5}$$

$$pK_a = – \log (1.21 \times 10^{-5})$$
$$= 4.92$$

A = Correct concentration of HPab OR K_a is correctly calculated from the given pK_a.
M = Correct concentration of HPab AND correct process for part (ii), with clearly stated and valid expression for K_a or pK_a.
E = Clearly shows the correct method using the correct data to determine the value of pK_a(HPab). It clearly states what species are involved, and identifies that $[H_3O^+] = [Pab^-]$.

e The equivalence point would be pH greater than 7 as at this point it is a solution of NaPab and the Pab^- ion is weakly basic. It reacts with water to produce an excess of OH^- ions.

$$Pab^- + H_2O \rightleftharpoons HPab + OH^-$$

A = Recognises pH greater than 7 due to solution being basic OR OH^- greater than H^+
OR because titration is weak acid with strong base OR halfway between initial and ‘final’ pH.
M = Links pH greater than 7 to reaction of Pab^- with H_2O to produce excess of OH^- ions.
MUST include equation for reaction of Pab^- with H_2O OR Correctly calculated pH of weakly basic solution.

f Sketch of graph with valid shape, showing:
- initial pH at 3.22,
- half-equivalence point at pH = 4.92
- equivalence point at pH greater than 7 (in range 8–10) and volume of 12 mL.

A = General shape of curve correct, with at least TWO of the following points correct:
- **starting at pH about 3.2**
- **midpoint of buffer region at about 4.9**
- **equivalence point at 12 mL which is given on *x* axis**

M = ALL correct as per stated evidence.

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Judgement Statement

Achievement

SIX opportunities answered at Achievement level or higher.
6 × A

Achievement with Merit

EIGHT opportunities answered with at least FIVE at Merit level or higher.
5 × M *plus* 3 × A

Achievement with Excellence

TEN opportunities answered with at least TWO at Excellence level and FIVE at Merit level.
2 × E *plus* 5 × M *plus* 3 × A