The **topics** and **types of questions** examined in this Achievement Standard. Use this sheet to plan and organise your study so that you cover everything that is required.

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## 3.7 AQUEOUS SYSTEMS — AS 90700

Describe properties of aqueous systems

## 3.7 1. Solubility

- relative equilibrium concentrations of dissolved species
- variability in solubility of a sparingly soluble salt due to the formation of a complex ion, the addition of a common ion, or the reaction of a basic anion with added acid.
- calculations involving $K_s$ and solubility limited to:
  - AB, $A_2B$ and $AB_2$ type solids where neither of the ions A or B react further with water
  - calculating the concentration of one ion given the other
  - calculating the solubility in water and in solutions already containing one of the ions A or B (common ion)
  - predicting precipitation or dissolution

(a) Describe what is meant by the term 'solubility'.

(b) The solubility product, $K_s$, of AgCl has a value of $1.56 \times 10^{-10}$ at 25°C and this value increases to $2.15 \times 10^{-8}$ at 100°C.

Explain why $K_s$ is higher at 100°C. Include reference to the relevant equilibrium equation in your answer.

The chloride ion concentration in sea water can be determined by titrating a sample with aqueous silver nitrate ($AgNO_3$) using potassium chromate ($K_2CrO_4$) as the indicator.

As the silver nitrate is added, a precipitate of silver chloride, (AgCl) forms. When most of the AgCl has precipitated, the $Ag^+(aq)$ concentration becomes high enough for a red precipitate of $Ag_2CrO_4$ to form.

(c) Show that the solubility of $Ag_2CrO_4$ in pure water at 25°C is higher than that of AgCl.

$K_s(AgCl) = 1.56 \times 10^{-10}$ $\qquad$ $K_s(Ag_2CrO_4) = 1.30 \times 10^{-12}$

(d) If the concentration of chromate ions is $6.30 \times 10^{-3}$ mol $L^{-1}$ at the point when the $Ag_2CrO_4$ starts to precipitate, calculate the concentration of $Ag^+$ ions in the solution.

## 3.7 2. Acids and Bases

- correlation between acid or base strength, $K_a$ and pH
- calculations involving $K_a$, $K_w$ and pH limited to:
  - solutions of bases, monoprotic acids and buffers
  - those in which the extent of reaction is small that the equilibrium concentration of a dissolved weak acid can be approximated by the initial concentration
  - pH at a particular point in a titration

When bromine is added to water, it forms hypobromous acid (HOBr), a weak acid.

(a) (i) Write an equation to show the equilibrium system that is formed with hypobromous acid and water.

(ii) Write the $K_a$ expression for hypobromous acid.

(b) Calculate the pH of a 0.0525 mol $L^{-1}$ hypobromous acid solution.

$pK_a(HOBr) = 8.62$

## 3.7 3. Buffer Solutions

- the nature of buffer solutions

Two solutions, A and B, were made as described below.

Solution A: 50 mL of aqueous 1.00 mol $L^{-1}$ ammonium chloride was added to 50 mL of aqueous 1.00 mol $L^{-1}$ ammonia.

Solution B: 25 mL of aqueous 0.010 mol $L^{-1}$ hydrochloric acid was added to 50 mL of aqueous 0.010 mol $L^{-1}$ ammonia.

(a) (i) Write the $K_a$ expression for $NH_4^+$.

(ii) Show, by calculation, that the pH of each of the two solutions is 9.24.

$pK_a(NH_4^+) = 9.24$

(b) (i) Discuss the abilities of solutions A and B to act as buffers.

(ii) Compare how the pH of each solution would be affected when 1.00 mol $L^{-1}$ sodium hydroxide is added drop-wise to each solution. Calculations are not required, but you should include appropriate equations in your answer.

## 3.7 4. Titration Curves

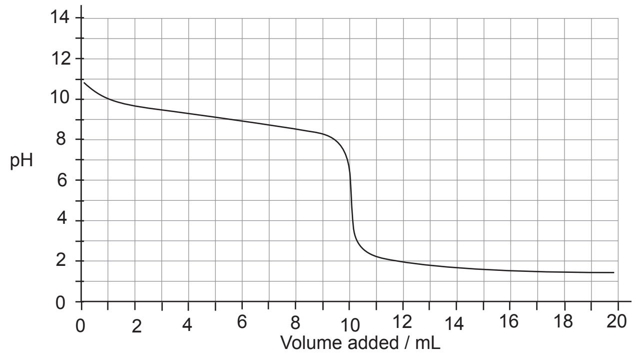
- features of titration curves include buffer region, equivalence point and selection of indicator
- titration of weak acids with weak bases are <u>excluded</u>

The graph below shows the change in pH when 40.0 mL of 0.0500 mol $L^{-1}$ aqueous $NH_3$ is titrated with 0.200 mol $L^{-1}$ aqueous HCl.

The equation for the reaction occurring during the titration is:

$$NH_3(aq) + HCl(aq) \rightarrow NH_4Cl(aq)$$

Titration Curve




(a) Use the curve to determine $pK_a(NH_4^+)$ and hence calculate $K_a(NH_4^+)$.

(b) Explain why the pH at the equivalence point for this titration is less than 7. (Include an equation to support your answer.)

(c) A $NH_4^+$ / $NH_3$ buffer solution is prepared with a pH of 9.6. Use the graph to describe how this buffer solution could be made from 0.0500 mol $L^{-1}$ $NH_3$ solution and 0.200 mol $L^{-1}$ HCl solutions.

A second titration is carried out – this time 40.0 mL of 0.0500 mol $L^{-1}$ $NH_4Cl$ solution is titrated against 0.200 mol $L^{-1}$ NaOH solution.

(d) Write an equation for the titration reaction.

(e) (i) Show that $[NH_3]$ at the equivalence point is 0.0400 mol $L^{-1}$.

(ii) Using $K_a(NH_4^+)$ determined in part (a) above, determine the pH at the equivalence point of the second titration.