Answer to Question 1 on page 12 of Tutorial 12.

1. In order to find the concentration of chloride ion in a sample of pool water, a 50.0 mL sample of the pool water was titrated with 0.500 M AgNO₃ solution, using sodium chromate solution (Na₂CrO₄ (aq)) as an indicator. At the equivalence point, it was found that 53.4 mL of AgNO₃ solution had been added.

a) Calculate the moles of Ag⁺ used.

$$\text{moles} = M \times \text{Litres}$$

$$= 0.500 \text{ M} \times 0.0534 \text{ L}$$

$$= 0.0267 \text{ moles of Ag}^+$$

b) Write the balanced net-ionic equation for the titration.

$$\text{Ag}^+(aq) + \text{Cl}^-(aq) \rightarrow \text{AgCl} (s)$$

c) Determine the moles of Cl⁻ ions in the sample.

From the ratio of coefficients in the balanced net-ionic equation:

$$\text{moles of Cl}^- = \text{moles of Ag}^+ \times \frac{1 \text{ mol Cl}^-}{1 \text{ mol Ag}^+}$$

$$= 0.0267 \text{ moles of Cl}^-$$

d) Calculate the [Cl⁻] in the sample of pool water.

$$M = \frac{\text{moles}}{L} = \frac{0.0267 \text{ moles}}{0.05 \text{ L}} = 0.534 \text{ M}$$

Since the lowest number of significant digits in the question is 3, the answer would have three significant digits.

so $[\text{Cl}^-] = 0.534 \text{ M}$
Answer to Question 2 on page 13 of Tutorial 12.

2. A solution containing silver ions (Ag⁺) is titrated with 0.200 M KSCN solution to find the [Ag⁺] in the sample. The indicator Fe(NO₃)₃ (aq) is used to signal when the equivalence point is reached. It is found that 15.6 mL of 0.200 M KSCN is needed to titrate a 25.0 mL sample of Ag⁺ solution. Determine the [Ag⁺] in the sample. Show all steps in a clear concise manner. (Use question 1 as a guide.)

**moles of SCN⁻ used:**

\[
\text{moles} = M \times L = 0.200 \text{ M} \times 0.0156 \text{ L} = 0.00312 \text{ moles of SCN}⁻
\]

**Balanced Net-Ionic Equation:**

\[
\text{Ag}⁺(aq) + \text{SCN}⁻(aq) \rightarrow \text{AgSCN}(s)
\]

**moles of Ag⁺ in sample:**

Since the ratio of coefficients is 1:1,

\[
\text{moles of Ag}⁺ = \text{moles of SCN}⁻ \times \frac{1 \text{ mol Ag}⁺}{1 \text{ mol SCN}⁻} = 0.00312 \text{ moles of Ag}⁺
\]

**Molar concentration of Ag⁺ ([Ag⁺])**

\[
M = \frac{\text{moles}}{L} = \frac{0.00312 \text{ moles}}{0.025 \text{ L}} = 0.1248 \text{ M}
\]

The answer would be rounded to 3 significant digits, as this was the lowest number of significant digits in the data.

\[
[\text{Ag}⁺] = 0.125 \text{ M}
\]
Answer to **Question 3** on page 13 of Tutorial 12.

3. Explain how the indicator Na$_2$CrO$_4$ works in titrations for chloride (Cl$^-$) ion concentration using Ag$^+$ as a standard solution.

The Ag$^+$ ions will keep bonding with the Cl$^-$ ions forming the white precipitate AgCl *as long as there are Cl$^-$ ions present.*

As soon as all the Cl$^-$ ions are used up, the Ag$^+$ will then start precipitating with the CrO$_4^{2-}$ ions, forming the precipitate Ag$_2$CrO$_4$. But recall from the last page that the colour of Ag$_2$CrO$_4$ is **brick red**. Thus, as you can see, *as soon as all the Cl$^-$ ions are used up, the next drop of Ag$^+$ solution will turn the solution red.*

So, as soon as all the Cl$^-$ is consumed, and a small amount of Ag$_2$CrO$_4$ forms, a faint brick red colour will be noticed. At this point, we would STOP the titration.

*************************************************************************
Answer to Question 4 on Page 13 of Tutorial 12.

4. Explain how the indicator Fe(NO₃)₃ works in titrations for silver (Ag⁺) ion concentration using SCN⁻ as a standard solution.

The main reaction for the titration is a precipitation of Ag⁺ and SCN⁻ ions to form a precipitate of AgSCN (s):

\[
\text{Ag}^+(aq) + \text{SCN}^-(aq) \rightarrow \text{AgSCN(s)}
\]

\text{colourless} \quad \text{colourless} \quad \text{white precipitate}

Once just enough SCN⁻ solution has been added to react with all the Ag⁺ ions, (the \textit{equivalence point}), any excess SCN⁻ ions added will react with the indicator, Fe³⁺ ions and form a complex ion (a larger ion made up of smaller ones) called FeSCN²⁺.
This ion, called the ferrothiocyanate ion, is NOT a precipitate, BUT is IS a very intense red colour. You may recall seeing it when you did Experiment 19-A on equilibrium. The reaction is:

\[
\text{Fe}^{3+}(aq) + \text{SCN}^-(aq) \rightarrow \text{FeSCN}^{2+}(aq)
\]

\text{pale rust} \quad \text{colourless} \quad \text{dark red}

A slight permanent red would appear at the \textit{endpoint} of the titration. This would indicate that the \textit{equivalence point} (the point where there is just enough SCN⁻ to react with all the Ag⁺ in the sample) has been reached.

This is the end of Tutorial 12 - Solutions