Tutorial 12 - Solutions

Determining the Concentration of a Specific Ion Using Precipitation Titrations.

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 $AgNO_3$ solution, using sodium chromate solution ($Na_2CrO_{4\ (aq)}$) as an indicator. At the equivalence point, it was found that 53.4 mL of $AgNO_3$ solution had been added.

a) Calculate the moles of Ag^+ used.

moles = M x Litres

= 0.500 M x 0.0534 L

= 0.0267 moles of Ag^+

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

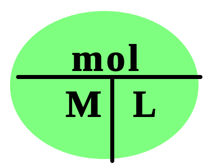
$$\mathbf{Ag^+_{(aq)} + Cl^-_{(aq)} \rightarrow AgCl_{(s)}}$$

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 moles of Cl^-



d) Calculate the $[Cl^-]$ in the sample of pool water.

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

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

so $[Cl^-]$ = 0.534 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 $Fe(NO_3)_{3\,(aq)}$ 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:

moles = M x L

= 0.200 M x 0.0156 L

= 0.00312 moles of SCN^-

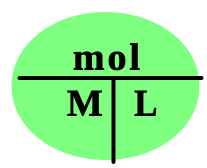
Balanced Net-Ionic Equation:

$$Ag^+_{(aq)} + SCN^-_{(aq)} \rightarrow AgSCN_{(s)}$$

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 moles of Ag^+

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

$$M = \frac{\text{moles}}{L} = \frac{0.00312 \text{ moles}}{0.025 \text{ L}} = \underline{0.1248 \text{ M}}$$

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

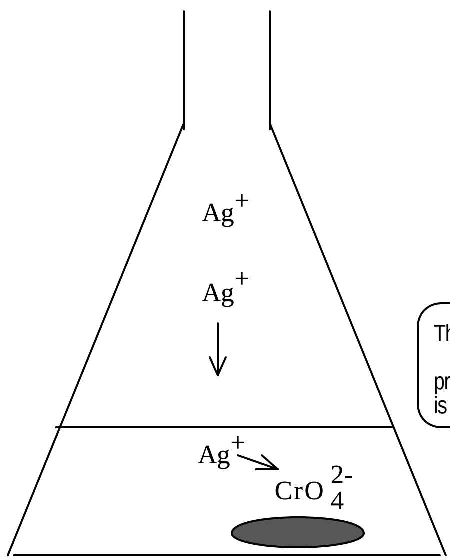
[Ag^+] = 0. 125 M

<u>Answer to **Question 3** on page 13 of Tutorial 12.</u>

3. Explain how the indicator Na_2CrO_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_2CrO_4. But recall from the last page that the colour of Ag_2CrO_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.



The Cl^- ions are now gone, so the only thing Ag^+ ions can precipitate with are the CrO_4^{2-} ions. The precipitate that forms is a <u>brick red</u> colour.

So, as soon as all the Cl^- is consumed, and a small amount of Ag_2CrO_4 forms, a faint brick red colour will be noticed. At this point, we would STOP the titration.

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

4. Explain how the indicator $Fe(NO_3)_3$ 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 $AgSCN_{(s)}$:

$$\underset{\textit{colourless}}{Ag^+_{(aq)}} + \underset{\textit{colourless}}{SCN^-_{(aq)}} \rightarrow \underset{\textit{white precipitate}}{AgSCN_{(s)}}$$

Once just enough SCN^- solution has been added to react with all the Ag^+ ions, (the ***equivalence point***), any excess SCN^- ions added will react with the indicator, Fe^{3+} ions and form a complex ion (a larger ion made up of smaller ones) called $FeSCN^{2+}$.
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:

$$\underset{\textit{pale rust}}{Fe^{3+}_{(aq)}} + \underset{\textit{colourless}}{SCN^-_{(aq)}} \rightarrow \underset{\textit{dark red}}{FeSCN^{2+}_{(aq)}}$$

A slight permanent red would appear at the *endpoint* of the titration. This would indicate that the *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

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