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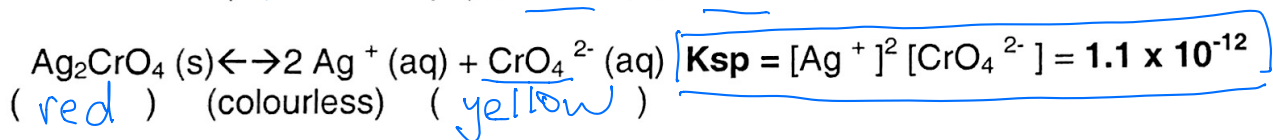
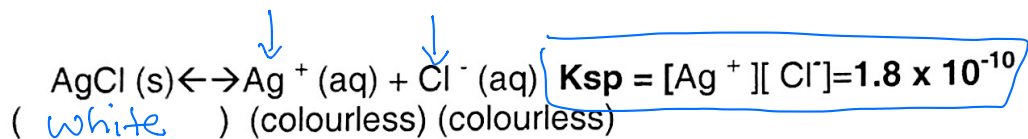
Chemistry 12
Solubility Lesson # 8
Chloride Titrations

Recall from Unit V + Chem 11:

A Titration is a process where a measured amount of a substance is reacted with an unknown quantity of another substance until the desired "equivalence point" is reached (coincides with a colour change)
The purpose of carrying out a TITRATION is to determine either [] or [] of an unknown substance.

For SILVER-CHLORIDE TITRATIONS the chromate ion is used as an indicator.

WHY?



When unknown $[\text{Ag}^+](\text{aq})$ is slowly added into a beaker containing both 0.10 M Cl^- (aq) and 0.10 M CrO_4^{2-} (aq). What will the first ppt to form be? ($Q = K_{sp}$)

Re-arrange the above K_{sp} expressions and Solve for $[\text{Ag}^+]$:

$$K_{sp} = [\text{Ag}^+][\text{Cl}^-] = 1.8 \cdot 10^{-10}$$

$$\frac{[x][0.10]}{0.10} = \frac{1.8 \cdot 10^{-10}}{0.10}$$

$$[\text{Ag}^+] = x = \frac{1.8 \cdot 10^{-9} \text{ M}}{\text{AgCl}}$$

$$K_{sp} = [\text{Ag}^+]^2 [\text{CrO}_4^{2-}] = 1.1 \cdot 10^{-12}$$

$$\frac{[x]^2 [0.10]}{0.10} = \frac{1.1 \cdot 10^{-12}}{0.10}$$

$$\sqrt{x^2} = \sqrt{\frac{1.1 \times 10^{-11}}{0.10}}$$

$$[\text{Ag}^+] = x = 3.3 \cdot 10^{-6} \text{ M}$$

The first ppt to form will be AgCl(s) as it requires a SMALLER $[\text{Ag}^+]$.

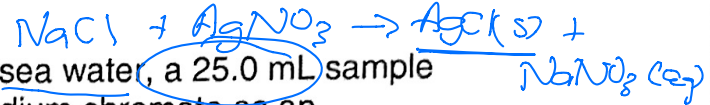
As more and more Ag^+ is added to the beaker the Cl^- is eventually all used up, and at that point the CrO_4^{2-} will begin to combine with the Ag^+ and there will be a distinctive orange colour produced as Ag₂CrO₄ is formed. At this point the titration is **STOPPED**.

AT THIS POINT THE MOLES OF Ag^+ ADDED = MOLES OF Cl^- PRESENT

↳ you have reached the "EQUIVALENCE POINT" of the rxn

IN SUMMARY:

when the colour ORANGE appears the titration is STOPPED as the equivalence point is reached



Example 1. In order to find the $[\text{Cl}^-]$ in a sample of sea water, a 25.0 mL sample was titrated with 0.500 M AgNO_3 solution, using sodium chromate as an indicator. At the EQUIVALENCE POINT 26.8 mL of AgNO_3 had been added. What was the $[\text{Cl}^-]$ in the sea water?

Step 1. Balanced equation



Step 2. Solve for moles of KNOWN

$0.0268 \text{ L} \left(\frac{0.500 \text{ mol AgNO}_3}{1 \text{ L}} \right) \left(\frac{1 \text{ mol Ag}^+}{1 \text{ mol AgNO}_3} \right) \left(\frac{1 \text{ mol Cl}^-}{1 \text{ mol Ag}^+} \right) \left(\frac{1}{0.0250 \text{ L}} \right)$

Step 3. Convert to moles of UNKNOWN

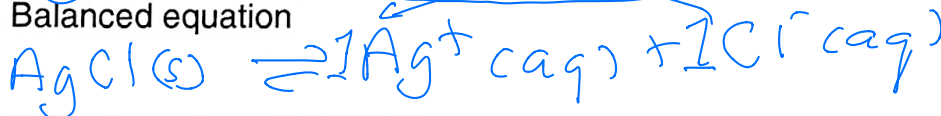
$= 0.536 \text{ M Cl}^-$

Step 4. Solve for $[\text{Cl}^-]$ of UNKNOWN

$M = \text{mol} \div \text{L}$

Example 2. What volume of 0.125 M AgNO_3 will be required to titrate 50.0 mL of 0.0500 M Cl^- solution, using the chromate indicator?

Step 1. Balanced equation



Step 2. Solve for moles of KNOWN

$0.0500 \text{ L} \left(\frac{0.0500 \text{ mol Cl}^-}{1 \text{ L}} \right) \left(\frac{1 \text{ mol Ag}^+}{1 \text{ mol Cl}^-} \right) \left(\frac{1 \text{ mol AgNO}_3}{1 \text{ mol Ag}^+} \right) \left(\frac{1 \text{ L}}{0.125 \text{ mol AgNO}_3} \right)$

Step 3. Convert to moles of UNKNOWN

$0.0200 \text{ mol AgNO}_3 = 0.0200 \text{ L AgNO}_3$

Step 4. Solve for volume of UNKNOWN

$\text{L} = \text{mol} \div M \leftarrow \frac{0.0200 \text{ mol}}{0.125 \text{ M}} = 0.160 \text{ L} \text{ or } 160 \text{ mL AgNO}_3$

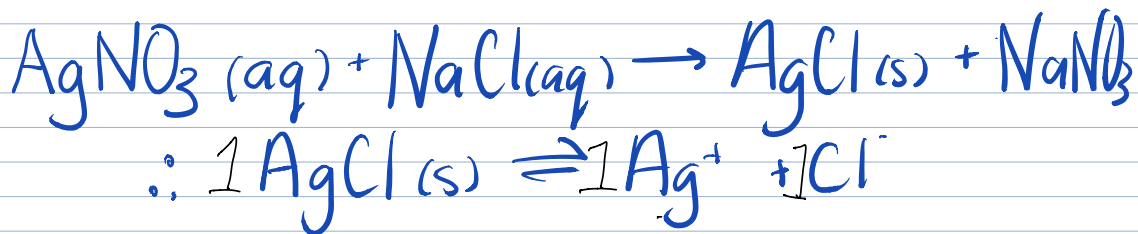
Sample Problem. A 5.29 g sample of impure NaCl was dissolved and diluted to a total volume of 250.0 mL. If 25.0 mL of the NaCl solution required 28.5 mL of 0.300 M AgNO_3 solution to reach the equivalence point, using the chromate indicator, what was the percentage purity of the original NaCl solution?

Recall Percent Purity = actual / expected x 100 %

Seatwork/Homework: Exercises 70-75 pgs 101-102

PLO's: I7

Sample Problem: $P.P. = \frac{\text{actual}}{\text{expected}} \cdot 100$



actual \rightarrow (titration) $0.0285 \text{ L} \cdot \left(\frac{0.300 \text{ mol Ag}^+}{1 \text{ L}} \right) \left(\frac{1 \text{ mol Cl}^-}{1 \text{ mol Ag}^+} \right)$
 $= 8.55 \cdot 10^{-3} \text{ mol Cl}^- \Rightarrow 8.55 \cdot 10^{-3} \text{ mol NaCl}$

$[\text{NaCl}] = \frac{8.55 \cdot 10^{-3} \text{ mol NaCl}}{0.0250 \text{ L}} = 0.342 \text{ M NaCl}$ (actual)

expected \rightarrow (impure) $5.29 \text{ g NaCl} \left(\frac{1 \text{ mol}}{58.5 \text{ g}} \right) \left(\frac{1}{0.2500 \text{ L}} \right)$
 $= 0.362 \text{ M NaCl}$ (expected)

Percent Purity $= \frac{0.342 \text{ M}}{0.362 \text{ M}} \cdot 100 = 94.6\%$