

Name: Key  
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Chemistry 12  
Solubility Lesson # 8  
Chloride Titrations

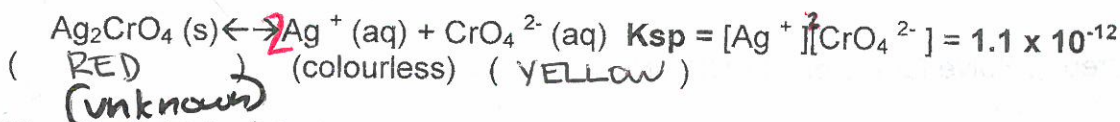
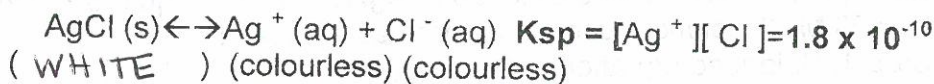
Recall from Unit V + Chem 11:

A Titration is a process where a measured amount of a solution is reacted with a known volume of another until a desired EQUIVALENCE PT is reached.

The purpose of carrying out a TITRATION is to determine the [ ] of an unknown substance.

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

WHY?



When ~~0.10 M~~  $\text{Ag}^+$  (aq) is slowly added into a beaker containing both 0.10 M  $\text{Cl}^-$  (aq) and 0.10 M  $\text{CrO}_4^{2-}$  (aq). What will the first ppt to form be?

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

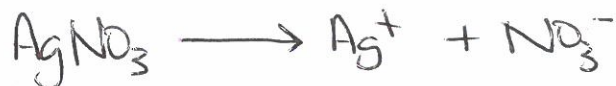
$$K_{sp} = [\text{Ag}^+][\text{Cl}^-] \Rightarrow [\text{Ag}^+] = \frac{1.8 \times 10^{-10}}{0.10} = 1.8 \times 10^{-9}$$

$$K_{sp} = [\text{Ag}^+]^2[\text{CrO}_4^{2-}] \Rightarrow [\text{Ag}^+] = \sqrt{\frac{1.1 \times 10^{-12}}{0.10}} = 3.3 \times 10^{-6}$$

The first ppt to form will be  $\text{AgCl (s)}$  as it requires a SMALLER  $[\text{Ag}^+]$ . As more and more  $\text{Ag}^+$  is added to the beaker the  $\text{Cl}^-$  is eventually all used up, and at that point the  $\text{CrO}_4^{2-}$  will begin to combine with the  $\text{Ag}^+$  and there will be a distinctive ORANGE colour produced as  $\text{Ag}_2\text{CrO}_4 \text{ (s)}$  is formed. At this point the titration is STOPPED.

AT THIS POINT THE MOLES OF  $\text{Ag}^+$  ADDED = MOLES OF  $\text{Cl}^-$  PRESENT

IN SUMMARY: When the colour "orange" is observed the equivalence pt is reached where moles  $\text{Ag}^+$  = moles  $\text{Cl}^-$ .



**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  $\text{AgNO}_3$  solution, using sodium chromate as an indicator. At the EQUIVALENCE POINT 26.8 mL of  $\text{AgNO}_3$  had been added. What was the  $[\text{Cl}^-]$  in the sea water?

Step 1. Balanced equation  $\text{AgCl}(s) \rightleftharpoons \text{Ag}^+(aq) + \text{Cl}^-(aq)$

Step 2. Solve for moles of KNOWN

$$\frac{0.500 \text{ mol Ag}^+}{\cancel{\text{L}}} \times 0.0268 \cancel{\text{L}} = 0.0134 \text{ mol Ag}^+$$

Step 3. Convert to moles of UNKNOWN

$$0.0134 \text{ mol Ag}^+ \times \frac{1 \text{ mol Cl}^-}{1 \text{ mol Ag}^+} = 0.0134 \text{ mol Cl}^-$$

Step 4. Solve for  $[\ ]$  of UNKNOWN  $\rightarrow \frac{0.0134 \text{ mol Cl}^-}{0.0250 \text{ L}} = \boxed{0.536 \text{ M Cl}^-}$

**Example 2.** What volume of 0.125 M  $\text{AgNO}_3$  will be required to titrate 50.0 mL of 0.0500 M  $\text{Cl}^-$  solution, using the chromate indicator?

Step 1. Balanced equation  $\text{AgCl}(s) \rightleftharpoons \text{Ag}^+(aq) + \text{Cl}^-(aq)$

Step 2. Solve for moles of KNOWN

$$\frac{0.0500 \text{ mol Cl}^-}{\cancel{\text{L}}} \times 0.0500 \cancel{\text{L}} = 2.50 \times 10^{-3} \text{ mol Cl}^-$$

Step 3. Convert to moles of UNKNOWN

$$2.50 \times 10^{-3} \text{ mol Cl}^- \times \frac{1 \text{ mol Ag}^+}{1 \text{ mol Cl}^-} = 2.50 \times 10^{-3} \text{ mol Ag}^+ = 2.50 \times 10^{-3} \text{ mol AgNO}_3$$

Step 4. Solve for volume of UNKNOWN

$$L = \frac{\text{mol}}{\text{M}} = \frac{2.50 \times 10^{-3} \text{ mol AgNO}_3}{0.125 \text{ M AgNO}_3} = \boxed{0.0200 \text{ L AgNO}_3}$$

or 20.0 mL



**Sample Problem.** A 5.29 g sample of impure  $\text{NaCl}$  was dissolved and diluted to a total volume of 250.0 mL. If 25.0 mL of the  $\text{NaCl}$  solution required 28.5 mL of 0.300 M  $\text{AgNO}_3$  solution to reach the equivalence point, using the chromate indicator, what was the percentage purity of the original  $\text{NaCl}$  solution?

Recall Percent Purity = actual / expected x 100 %

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

PLO's:

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