Name: Key

Blk:_____ Date:_____

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 (unknown) 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?**

$AgCl(s) \leftrightarrow Ag^{+}(aq) + Cl^{-}(aq)$ **Ksp = [Ag⁺][Cl]=1.8 x 10⁻¹⁰**

( WHITE ) (colourless) (colourless)

$Ag_2CrO_4(s) \leftrightarrow 2Ag^{+}(aq) + CrO_4^{2-}(aq)$ **Ksp =** $[Ag^{+}]^{2}[CrO_4^{2-}]$ **= 1.1 x 10⁻¹²**

( RED ) (colourless) ( YELLOW )

(unknown)

When [redacted] $Ag^{+}$ (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?

**Re-arrange** the above Ksp expressions and Solve for $[Ag^{+}]$:

$$K_{sp} = [Ag^{+}][Cl^{-}] \Rightarrow [Ag^{+}] = \frac{1.8\times10^{-10}}{0.10} = \boxed{1.8\times10^{-9}}$$

$$K_{sp} = [Ag^{+}]^{2}[CrO_4^{2-}] \Rightarrow [Ag^{+}] = \sqrt{\frac{1.1\times10^{-12}}{0.10}} = 3.3\times10^{-6}$$

$[x]^2$

The first ppt to form will be $AgCl(s)$ as is requires a SMALLER $[Ag^{+}]$. 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 $Ag_2CrO_4(s)$ is formed. At this point the titration is **STOPPED**.

**AT THIS POINT THE MOLES OF Ag⁺ ADDED = MOLES OF Cl⁻ PRESENT**

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

$AgNO_3 \longrightarrow Ag^+ + NO_3^-$

**Example 1.** In order to find the [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 [$Cl^-$ ] in the sea water?

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

Step 2. Solve for moles of KNOWN

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

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 → $\frac{0.0134 \text{ mol } Cl^-}{0.0250 \text{ L}} = \boxed{0.536 \text{ M } Cl^-}$

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 $AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$

Step 2. Solve for moles of KNOWN

$$\frac{0.0500 \text{ mol } Cl^-}{1 \text{ L}} \times 0.0500 \text{ L} = 2.50 \times 10^{-3} \text{ mol } Cl^-$$

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


$$L = \frac{mol}{M} = \frac{2.50 \times 10^{-3} \text{ mol } AgNO_3}{0.125 \text{ M } AgNO_3} = \boxed{0.0200 \text{ L } AgNO_3} \text{ or } 20.0 \text{ mL}$$

**Sample Problem**. A 5.29 g sample of impure NaCl was dissolved and diluted to at 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