Measurement of Chloride in Natural Waters

Background

Chloride is one of the major negatively charged ions found in water and sewage. The salty taste produced by chloride concentrations is variable and depends on what other components represent in the water. Some waters containing 250 milligrams per liter (mg/L) may have a detectable salty taste when sodium ions are present. On the other hand, waters containing as much as 1000 mg/L may not taste salty if calcium and magnesium ions are present. The chloride content of natural surface waters will depend to a great extent on the geology of the area. In a limestone area like Richmond, the natural surface water has very little chloride in it (10-50 mg/L). Any appreciably higher chloride concentration would suggest contamination. High concentrations of chloride can be damaging to metal pipes and structures, as well as to agricultural crops. Note that chlor ***ide*** , Cl^{--}, is quite different from chlor ***ine*** , Cl_2, as an aqueous species. Chlorine is added to drinking water in small amount (<1 ppm) and to swimming pools in somewhat larger amounts (ca. 5 ppm) for disinfecting purposes. Chlorine will produce some chloride as it reacts, but in most cases, except for pool water, it is not a major source of chloride.

In this study we will analyze a number of water samples from the local area for chloride content. A characteristic property of compounds containing chloride ion, Cl^-, is that it will react with silver nitrate, $AgNO_3$, to form an insoluble compound, silver chloride, AgCl. One way to write an equation for the reaction is as follows:

$$AgNO_3\ (aq) + Cl^-\ (aq)\ ?\quad AgCl(s)? + NO_3^-\ (aq)$$

This reaction provides the basis for a method of analysis called a titration. In a titration, a sample of chloride-containing solution is measured out, then a solution of silver nitrate is added in small portions until enough has been added to just exactly react with all of the chloride. If one knows the volume of silver nitrate added and its concentration, then it possible to calculate how much chloride must have been present to react with the silver nitrate. If the volume of the unknown solution is known, it is possible, in turn, to calculate the concentration of chloride in the solution. For the calculation of this analysis, concentrations are expressed as *molarity* (*M*), defined as ***moles of a substance per liter of solution.***

One problem remains. We need a way to know when just enough silver nitrate has been added. We solve this by adding a drop of sodium chromate solution (Na_2CrO_4) before starting the titration. As silver nitrate is slowly added, it reacts first with the chloride ions until they are used up. After they are all used up, the silver nitrate reacts with the chromate ion to form a deep red-colored solid compound. The point of color change is called the " end point" and the sodium chromate is often called an indicator because it is used to " indicate" the end point.

Write a balanced chemical equation for the end-point reaction below.

Procedure

1. Use the following data table to record your data.

Water Source:_________________________________

Trial	1	2	3	4	5
Number of Drops of Water					
Number of Drops of Indicator					
Number of Drops of $AgNO_3$ Solution					
Comments					

Concentration of the Silver Nitrate Solution used:______________________________

2. Place between ten and twenty drops of sample water in your test tube using one of your micropipets. Count the drops carefully so you know exactly how much water you added. This may not be as easy as it seems so you may have to start over until you get the hang of it.

3. Add one trop of Na2CrO4 indicator solution. The result should be a clear yellow-green solution.

4. Carefully and the Silver Nitrate one drop at a time (Count the number of drops) and observe the reaction. Swirl the test tub to mix everything between drops. Silver Chloride will form first and it is white. When all of the chloride is gone, Silver Chromate will form - it is red.

5. Stop adding Silver Nitrate drops when the solution stays red after swirling.

Calculations

1. Look carefully at the four (or more) results for each titration. Is the number of drops similar within each set? Does any one result seem way out of line from the others? If so, you probably did something wrong and therefore it is legitimate to omit this result. Simply indicate this in your notebook. Using all the results that you think are valid, calculate the average drops of $AgNO_3$ for each set and record.

2. The key idea in titrations is that at the end point the number of moles of one reagent used corresponds exactly to the number of moles of the other reagent used, according to the balanced chemical equation. Looking at the equation on the first page for this particular titration, we can see that the number of moles of $AgNO_3$ added from one pipet must equal the number of moles of chloride dispensed from the other pipet: moles of $AgNO_3$ added = moles of chloride in sample But the number of moles of each is equal to the volume times the concentration. If we assume that the drops are all the same size and therefore are a

fixed but unknown fraction of a liter, then we can say that drops of sample x *M* of Cl^- in sample = drops of $AgNO_3$ x *M* of $AgNO_3$ By rearranging we get the working equation: *M* of Cl^- = *M* of $AgNO_3$ x (drops of $AgNO_3$/drops of sample) You can now use this equation to calculate the molarity of chloride in each sample.

$$\text{molarity of chloride} = \text{molarity of AgNO3} \times \frac{\text{Drops of } AgNO_3}{\text{Drops of water sample}}$$

Perform your calculation in the space below:

Finally convert the answer to milligrams of chloride per liter (mg/L) which is the same as parts-per-million (ppm). To do this, multiply the molarity of the chloride by 35,500 mg Cl^- per 1 mole of Cl^-

Questions

1. Suppose that during the titrations you consistently added more Silver Nitrate ($AgNO_3$) to change the color of the indicator to a darker red-orange. Would this change the results? Explain.

2. What is the advantage of using 3 or more trials, especially since you are doing exactly the same thing each time?

3. What are the possible sources for the chloride in the water you tested?

4. Do you consider chloride a serious water pollutant? Why or why not?

Lab Exercise adapted from Stratton's Laboratory Manual to Accompany Chemistry in Context – Applying Chemistry to Society McGraw Hill. (2003) pp99-105.