**Lesson 10-14**

**Permanganatometry method**

Permanganatometry is one of the most commonly used redox titration methods. As a titrant, a solution of potassium permanganate is used, the oxidizing properties of which can be controlled depending on the acidity of the solution.

The method is based on the reaction

MnO4- + 8H+ + 5e = Mn2+ + 4H2O φo = 1.51 V

Potassium permanganate has a high oxidizing ability, it is cheap, KMnO4 solutions have sufficient stability if stored properly - this explains the widespread use of KMnO4 in volumetric analysis. Oxidation with a KMnO4 solution can take place in an acidic, neutral or alkaline environment. However, quantitative determinations by the method of permanganatometry are most often carried out in an acidic environment for the following reasons:

1) in an acidic environment, MnO4- goes into Mn2+, showing the greatest oxidative activity;

2) in an acidic environment, all reaction products are colorless and soluble.

In a neutral or weakly alkaline medium, MnO4- is reduced to the sparingly soluble oxide MnO2, which makes it difficult to determine the equivalent point.

Sulfuric acid is used to acidify titratable solutions. Nitric acid cannot be taken, because. she herself is a strong oxidizing agent and can react with determined reducing agents. Hydrochloric acid is a reducing agent and can react with KMnO4.

A special indicator for determining the equivalence point in permanganatometry is not required, because. solutions containing MnO4- ions are intensely colored and the first excess drop of it stains the titrated solution pink.

According to the method of permanganatometry, it is possible to determine:

a) direct titration - the amount of reducing agents - Fe2+, H2O2, oxalic acid and its salts, nitrites, etc.;

b) substitution titration - the amount of substances that react with reducing agents, for example Ca2 +, which reacts with C2O42-. To carry out the determination, Ca2+ is precipitated in the form of CaC2.O4, the precipitate is separated by filtration, dissolved in 2N H2SO4 and the released H2C2O4, titrated with a KmnO4 solution;

c) titration on the residue - the amount of substances that react with reducing agents, for example K2Cr2O7, which reacts with Mohr's salt (NH4) 2Fe (SO4) 2 6H2O - an excess of Mohr's salt is titrated with a solution of KMnO4.

Acid-base titration

In an aqueous medium, the reaction between an acid and a base can be represented by the equation:

H3O+ + OH– = 2H2O

Strong acids (hydrochloric acid, sulfuric acid) are used as titrants - acidimetry; or strong bases (caustic soda, caustic potash) - alkalimetry.

Alkalimetry is used for the quantitative determination of medicinal substances, which are:

– inorganic and organic acids;

– salts of organic bases (hydrochlorides, nitrates, hydrophosphates, lactates, hydrotartrates, etc.).

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Acidimetry is used to determine:

- organic bases that exhibit basic properties in aqueous or alcoholic media;

- sodium salts of weak inorganic and organic acids.

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One of the acid-base titrations used is the combination of a neutralization reaction with pre-esterification or hydrolysis. Some medicinal substances, derivatives of alcohols or phenols are acetylated with acetic anhydride (an ester is formed). Excess acetic anhydride is converted to acetic acid and titrated with alkali. The possibility of using the acid-base titration method for the analysis of medicinal substances is determined by the dissociation constant of the titratable substance and its concentration in the solution.

The magnitude of the titration jump on the titration curve depends significantly on the dissociation constant. When determining medicinal substances by the method of neutralization, Ka and Kv of acids and bases must be at least 10-7. Thus, when titrating 0.1 mol/l solutions of strong acids and alkalis, the titration jump is about 6 pH units; if Ka(Kv) = 10-3, then 3-4 pH units; at Ka(Kv) = 10-5, 2-2.5 pH units; at Ka(Kv) = 10-9–10-10, there is no titration jump and determination of the titration end point becomes practically impossible.

When titrating a 0.1 mol/l solution of a strong acid with an alkali solution and vice versa, the titration jump is about 6 pH units, at a concentration of 0.01 mol/l, respectively, 3.4 pH units; at 0.001 mol/l - 1.4 pH units; at 0.0001 mol/l there is no titration jump.

Mixed solvents are used to enhance the acid-base properties of analytes and also when the drug is poorly soluble in water (for example, titration of sulfanilamide preparations with a dissociation constant of 10-7-10-8 (norsulfazol).

Titration in non-aqueous solvents.

The method of acid-base titration in non-aqueous solvents is used for the quantitative determination of weak acids (barbiturates, sulfonamides), weak bases (caffeine, reserpine). Salts of organic bases. This method allows the determination of many medicinal substances that, when titrated in aqueous solutions, do not have a clearly defined end point of the titration. Under the influence of non-aqueous solvents, the acid-base properties of various substances change. Depending on the solvent, the same substance can become an acid, a base, an amphoric or neutral compound, a strong or weak electrolyte. The strength or weakness of an acid or base is determined by the nature of its interaction with the solvent. In the acid-base process, all solvents are divided into two large groups: APROTIC and PROTOLYTIC.

Aprotic solvents are chemical compounds of a neutral nature, the molecules of which are not ionized and are not capable of either donating or adding a proton. Aprotic solvents do not interact with the substance dissolved in them. Such solvents include hydrocarbons (benzene, toluene, hexane) and their halogen derivatives. Aprotic solvents are often added to the titrated solution to suppress the process of solvolysis of neutralization products, which contributes to a clearer determination of the end point of the titration.

Protolytic solvents are chemical compounds whose molecules are capable of donating or accepting protons. They are involved in the acid-base process. Protolytic solvents, in turn, can be divided into three groups:

Amphiprotic - amphoteric, capable of both donating and accepting a proton. Water, alcohols.

Protogenic or acidic solvents. Substances in which the ability to donate a proton significantly exceeds the ability to attach it. Acetic acid, formic acid. Protogenic solvents enhance the basic properties of chemical compounds.

Protophilic or basic solvents. Liquid ammonia, pyridine, DMF, and other protophilic solvents enhance the acidic properties of weak acids and amphoteric compounds.

A typical example is the titration of potassium acetate in anhydrous acetic acid with perchloric acid.

Titration in protophilic solvents is carried out with potassium or sodium methylates in pyridine.

Argentometry

Argentometric methods of analysis are among the precipitation methods that have found application in the analysis of drug substances. Argentometry is based on the reactions of precipitation of halides (chlorides, bromides and iodides) and thiocyanates with a solution of silver nitrate. This method determines not only inorganic medicinal substances, which are alkali metal halides, but also organic ones: salts of hydrohalic acids - hydrochlorides, hydrobromides and hydroiodides of organic bases (for example, ephedrine hydrochloride, scopolamine hydrobromide, pachycarpine hydroiodide, etc.); quaternary ammonium halides (eg, carbachol, pentamine, and others); organic compounds with a covalently bound halogen after the conversion of the halogen into an ionic state (for example, iodoform, bromisoval, bromocamphor, and others).

Argentometry is used for the analysis of thiocyanates, chlorides, bromides, iodides of alkaline earth metals and organic bases. The working solution is a 0.1M AgNO3 solution, and in the Volhard method, a 0.1M NH4CNS solution.

To determine the equivalence point depending on the pH of the medium in argentometry, a number of indicators are used:

1. K2CrO4 - 5% potassium chromate solution;

2. NH4Fe(SO4)2 \* 12H2O - saturated solution of iron-ammonium alum, acidified with nitric acid;

3. Adsorption indicators: sodium eosinate, bromophenol blue, fluorescein;

4. Freshly prepared starch solution;

5. Nitroso-starch paper.

The essence of argentometric methods is to isolate the analyte in the form of sparingly soluble precipitates of silver salts: AgCl, AgBr, AgJ, AgCN, AgSCN. After all halogen ions have been precipitated, an excess drop of 0.1M AgNO3 will react with the indicator to form colored precipitates or colored solutions at the equivalence point.

The ongoing reactions must satisfy the following conditions:

1. The precipitate should fall out quickly and be practically insoluble.

2. The results of the titration should not be affected by side reactions.

3. The equivalence point should be easy to fix.

Depending on the method of titration and the indicator used, argentometry methods are divided into:

§ non-indicator:

- Gay-Lussac method (equal haze method)

- method to the point of enlightenment

§ indicator:

- Mohr's method

- Faience-Fischer-Khodakov method

- Volgard method

The essence of argentometric methods is to isolate the analyte in the form of sparingly soluble precipitates of silver salts.

Gay-Lussac method

The Gay-Lussac method is a method of direct titration of chlorides and bromides with a 0.1M AgNO3 solution without an indicator. The equivalent point is observed when the formation of a precipitate stops and the titrated solution becomes clear.

Mohr method. Direct Argentometry.

Only for Cl–, Br–.

Titrant - AgNO3, indicator - K2CrO4, medium: pH=6–8.

NaBr + AgNO3 → AgBr↓ + NaNO3,

K2CrO4 + 2AgNO3 → Ag2CrO4↓ + 2KNO3.

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fequiv(NaBr)=1,

Not allowed:

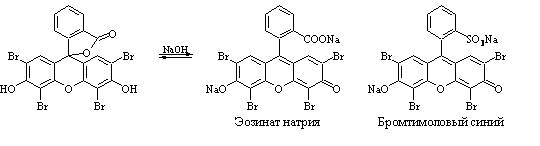
pH<6: CrO72– + 2H+ → Сr2O72– + H2O,

pH>8: 2Ag+ + 2OH– → Ag2O + H2O.

Presence of ions: SO42–, S2–, PO43–, AsO33–, BO33–, CO32–; Bi3+, Ba2+, Pb2+.

Faience method. Direct Argentometry.

Titrant – AgNO3, indicators – bromthymol blue (Cl–), sodium eosinate (Br–, I-), medium – CH3COOH (30%). Fluorescein - in a neutral and slightly alkaline medium.



NaI + AgNO3 → AgI↓ + NaNO3.

Silver iodide adsorbs the ions of the same name on itself; a bright pink color appears:

{m(AgI)∙nI–(n–x)K+}x–

At the equivalence point, the colloidal particle becomes electrically neutral; in the CTT, it begins to adsorb Ag+; micelles are recharged, the precipitate coagulates, the solution becomes clear:

{m(AgI)∙nI–(n–x)K+}x–∙xAg+,

{m(AgI)∙nI–(n–x)K+}x–∙xAg+ + Ind2– → (x/2)Ag2Ind + {m(AgI)∙nI–(n–x)K+}x–.

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fequiv(NaI)=1, .

Foldguard method. Reverse argentometry.

Titrants - 1) AgNO3, 2) NH4SCN;

Indicators – NH4Fe(SO4)2 (Cl–, Br–), FeCl3 (I–);

Medium - HNO3 (pH=3).

HNO3, the exact excess volume of AgNO3 and the indicator are added to the weighed solution of the preparation:

KBr + AgNO3 → AgBr↓ + KNO3,

AgNO3 + NH4SCN → AgSCN↓ + NH4NO3;

In KTT, the indicator interacts with the titrant:

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fequiv(KI)=1, .

Features of chloride titration:

1. Low titration rate [Ks(AgCl)>Ks(AgSCN)],

2. Very weak agitation,

3. Adding toluene or benzene to extract AgCl.

Features of titration of iodides:

After adding silver nitrate, complete precipitation of AgI is expected, the second indicator is added just before titration:

2KI + 2FeCl3 → I2 + 2FeCl2 + 2KCl.

Calthoff-Stangler method. /Only for Br–/

Titrants - 1) AgNO3, 2) NH4SCN, indicator - NH4Fe(SO4)2, medium - HNO3 (pH=3).

HNO3, an indicator and the exact excess volume of NH4SCN (0.1 ml) are added to the weighed solution of the drug:

3NH4SCN + NH4Fe(SO4)2 → Fe(SСN)3 + 2(NH4)2SO4.

Titrate with silver nitrate until the red color disappears:

KBr + AgNO3 → AgBr↓ + KNO3,

AgNO3 + NH4SCN → AgSCN↓ + NH4NO3,

3AgNO3 + Fe(SCN)3 → 3AgSCN↓ + Fe(NO3)3.

*f*экв(KBr)=1, http://nesterovdmitriy.narod.ru/03-Nauchnaya_Deyat/pharm_chem/o_argentometr.files/image024.gif.

Kaltgof method (Starch iodine). /For I– in the presence of Cl– and Br–/

Titrant - AgNO3, indicator - starch, medium - H2SO4.

Add 1 drop of KIO3 (0.1 mol/l), starch solution and H2SO4 (1:5) to the weighed solution of the drug until a blue color appears:

5KI + KIO3 + 3H2SO4 → 3I2 + 3K2SO4 + 3H2O,

Titrate with AgNO3 (0.1 mol/l) until the blue color disappears (after removal of iodide):

KI + AgNO3 → AgI↓ + KNO3.

*f*экв(KI)=1, http://nesterovdmitriy.narod.ru/03-Nauchnaya_Deyat/pharm_chem/o_argentometr.files/image026.gif.

Argentometry with an external indicator. /For I–/

Titrant - AgNO3, indicator - starch nitrite paper, medium - H2SO4.

H2SO4 (1:5) is added to the weighed solution of the preparation, AgNO3 is titrated. Near the equivalence point, after each addition of the titrant, place a drop of the solution on starch nitrite paper. The titration is completed when there is no blue color on the paper from the addition of the solution:

NaI + AgNO3 → AgI↓ + NaNO3,

2KI + 2KNO2 +2H2SO4 → I2 + 2NO + 2K2SO4 + 2H2O.

**Thiocyanatometry**

Determined substances: drugs that contain silver (protargol, collargol, silver nitrate).

Titrant: ammonium or potassium thiocyanate NH4SCN, KSCN - secondary standard solutions

Indicator for standardization of ammonium or potassium thiocyanates - ferum salts (ІІІ):

**Fe3+ + SCN-= [Fe(SCN)]2+**

Medium: nitrate

Method indicator: ferum salts (ІІІ) NH4Fe(SO4)212H2O in the presence of nitrate acid

Standardization: according to AgNO3 standard solution:

**AgNO3 + NH4SCN = AgSCN + NH4NO3**

iodometry method

I. The essence of the method

This method uses the oxidizing properties of free iodine and the reducing properties of iodide ions:

I2 + 2e = 2I-φo = + 0,54 B

1. Using a working solution of iodine, it is possible to determine the number of various reducing agents, the redox potential of which is less than the potential of the I2 / 2I- system.

2. Using solutions of iodides, for example KI, it is possible to determine the amount of oxidizing agents whose oxidizing potential is higher than the potential of the I2 / 2I- system.

The oxidation of reducing agents is carried out by direct titration of the reducing agent solution with a working solution of iodine. An example is the determination of sodium sulfite, which reacts with I2 according to the equation:

I2 + Na2SO3 + H2O = Na2SO4 + 2HI

Similarly, you can determine the amount of SnCl2, H2S and sulfides, H3AsO3. and other restorers.

However, the iodometric determination of reducing agents by direct titration with a working solution of iodine is not widely used. Much more often they are determined by titration on the residue. To do this, a working solution of I2 is added in excess to a solution containing a reducing agent, as is done, for example, when determining the content of sulfuric acid in hydrogen sulfide water:

H2S + I2 = S0 + 2HI

The residue I2 is titrated with a working solution of sodium thiosulfate.

Na2S2O3 according to the reaction:

2Na2S2O3 + I2 = Na2S4O6 + 2NaI

f(Na2S2O3) = 1; f(I2) = 1/2.

The iodometric method finds the widest application for determining the amount of oxidizing agents by the substitution method. This is how permanganate, dichromate, Cu2+, Fe3+, ClO-, etc. are determined.

For example, when determining the amount of KMnO4, a solution of KI (which is an auxiliary solution) is added to a permanganate solution in an acidic medium, and I2 is released, the amount of which is equivalent to the content of KMnO4, (first stage):

2KMnO4 + 10KI + 8H2SO4 = 5I2 + 2MnSO4 + 6K2SO4 + 8H2O.

f(KMnO4) = 1/5; f(KI) = 1.

The released iodine is titrated with a working solution of sodium thiosulfate (second stage):

2Na2S2O3 + I2 = 2NaI + Na2S4O6

It is impossible to carry out such a determination directly by titration of oxidizing agents with a KI solution, since it is not possible to notice the end of the formation of 12, a sharp change in the appearance of the solution is not observed.

Direct titration of oxidizing agents with a solution of sodium thiosulfate is also impossible, because the reaction proceeds very complicatedly and a mixture of products of an indeterminate composition is formed, as a result of which it is impossible to make a calculation.

3. In addition to oxidizing and reducing agents, it is possible to determine the amount of strong acids by iodometry. Their definition is based on the fact that KI and KIO3 do not interact in a neutral medium, but if acids are added to the KI + K103 mixture, then free iodine is released according to the equation:

KIO3 + 5KI + 6HCl = 3I2 + 6KCl + 3H2O

f(KIO3) = 1/5; f(KI) = 1; f(HCl) = 1/6.

It follows from the reaction equation that the acid participates in the reaction and the amount of released iodine is equivalent to the amount of acid present in the solution

Definition of the equivalence point

The equivalence point in iodometry can be determined by the appearance or disappearance of 12, the aqueous solution of which, in the presence of KI, is quite intensely colored yellow-brown (KI3 complex). However, much more accurate results are obtained with the introduction of an indicator - starch solution. Starch with I3 is a dark blue adsorption compound. In this compound, the iodine molecule is deformed, due to which it changes its color from brown to dark blue.

For direct titration of reducing agents. for example, SnCI2, Na2HAsO3, etc., starch is added before starting the titration.

When titrating reducing agents by residue, as well as when determining oxidizing agents and acids by the substitution method, starch is added to the reaction mixture at the end of the titration, when the solution turns from brown to pale yellow. Otherwise, the starch will adsorb large amounts of iodine and slowly release it into the solution, as a result of which an excess of sodium thiosulfate will be consumed and the titration results will be distorted.

Conditions for performing iodometric determinations

1) Free iodine is volatile and when heated, its volatility increases. The sensitivity of starch as an indicator (adsorption) decreases with increasing temperature.

Therefore, titration should be carried out in the cold.

2) Iodine reacts with alkalis according to the equation:

3I2 + 6NaOH = NaIO3 + 5NaI + 3H2O

Therefore, iodometric determinations cannot be carried out in a strongly alkaline medium. The result of the analysis in this case cannot be accurate.

3) The reactions occurring during iodometric determinations are not very fast.

To increase the rate of reaction of potassium iodide with an oxidizing agent, an excess of sulfuric acid and iodide is taken. The reaction mixture is allowed to stand for 5-6 min.

4) The end of the titration is determined by the disappearance (or appearance) of the blue color inherent in starch in the presence of free iodine.

It should be remembered that the solution should discolor from one drop of sodium thiosulfate. Further addition of the latter will not change the color of the titrated solution, but will make the analysis incorrect.

With the help of iodometric titrations, the amount of sugar in the blood is found, the rate constant of the oxidation reaction of KI is determined by the action of H2O2, etc.