# ****TOTAL ACIDITY****

**Introduction**

 Acidity of water is its quantitative capacity to react with a strong base to a designated pH. The measured value may vary significantly with the end-point pH used in the determination. Acidity is a measure of an aggregate property of water and can be interpreted in terms of specific substances only when the chemcical composition of the sample is known. Strong mineral acids, weak acids such as carbonic and acetic, and hydrolyzing salts such as iron or aluminum sulfates may contribute to the measured acidity according to the method of determination. Acids contribute to corrosiveness and influence chemical reaction rates, chemical speciation, and biological processes.

**Titration Method**

**1. General Discussion**

a. Principle: Hydrogen ions present in a sample as a result of dissociation or hydrolysis of solutes react with additions of standard alkali. Acidity thus depends on the end-point pH or indicator used. The construction of a titration curve by recording sample pH after successive small measured additions of titrant permits identification of inflection points and buffering capacity, if any, and allows the acidity to be determined with respect to any pH of interest. In the titration of a single acidic species, as in the standardization of reagents, the most accurate end point is obtained from the inflection point of a titration curve. The inflection point is the pH at which curvature changes from convex to concave or vice versa. Because accurate identification of inflection points may be difficult or impossible in buffered or complex mixtures, the titration in such cases is carried to an arbitrary end-point pH based on practical considerations. For routine control titrations or rapid preliminary estimates of acidity, the color change of an indicator may be used for the end point. Samples of industrial wastes, acid mine drainage, or other solutions that contain appreciable amounts of hydrolyzable metal ions such as iron, aluminum, or manganese are treated with hydrogen peroxide to ensure oxidation of any reduced forms of polyvalent cations, and boiled to hasten hydrolysis. Acidity results may be highly variable if this procedure is not followed exactly.

b. End points: Ideally the end point of the acidity titration should correspond to the stoichiometric equivalence point for neutralization of acids present. The pH at the equivalence point will depend on the sample, the choice among multiple inflection points, and the intended use of the data. Dissolved carbon dioxide (CO2) usually is the major acidic component of unpolluted surface waters; handle samples from such sources carefully to minimize the loss of dissolved gases. In a sample containing only carbon dioxide-bicarbonates-carbonates, titration to pH 8.3 at 25°C corresponds to stoichiometric neutralization of carbonic acid to bicarbonate. Because the color change of phenolphthalein indicator is close to pH 8.3, this value generally is accepted as a standard end point for titration of total acidity, including CO2 and most weak acids. Metacresol purple also has an end point at pH 8.3 and gives a sharper color change.

c. Interferences: Dissolved gases contributing to acidity or alkalinity, such as CO2, hydrogen sulfide, or ammonia, may be lost or gained during sampling, storage, or titration. Minimize such effects by titrating to the end point promptly after opening sample container, avoiding vigorous shaking or mixing, protecting sample from the atmosphere during titration, and letting sample become no warmer than it was at collection.

In the potentiometric titration, oily matter, suspended solids, precipitates, or other waste matter may coat the glass electrode and cause a sluggish response. Difficulty from this source is likely to be revealed in an erratic titration curve. Do not remove interferences from sample because they may contribute to its acidity. Briefly pause between titrant additions to let electrode come to equilibrium or clean the electrodes occasionally.

In samples containing oxidizable or hydrolyzable ions such as ferrous or ferric iron, aluminum, and manganese, the reaction rates at room temperature may be slow enough to cause drifting end points.

Do not use indicator titrations with colored or turbid samples that may obscure the color change at the end point. Residual free available chlorine in the sample may bleach the indicator. Eliminate this source of interference by adding 1 drop of 0.1M sodium thiosulfate (Na2S2O3).

d. Sampling and storage: Collect samples in polyethylene or borosilicate glass bottles and store at a low temperature. Fill bottles completely and cap tightly. Because waste samples may be subject to microbial action and to loss or gain of CO2 or other gases when exposed to air, analyze samples without delay, preferably within 1 d. If biological activity is suspected analyze within 6h. Avoid sample agitation and prolonged exposure to air.

**2. Apparatus**

a. pH meter

b. Titration vessel

c. Magnetic stirrer.

d. Pipets, volumetric.

e. Flasks, volumetric,

f. Buret

**3. Reagents**

0.02N NaOH solution

Phenolphthalein indicator

**4. Procedure**

Potentiometric titration to pH 8.3:

Prepare sample and titration assembly. Titrate to preselected end-point pH 8.3 As the end point is approached make smaller additions of alkali and be sure that pH equilibrium is reached before making the next addition.

**5. Calculation**

$$Acidity, as mg\frac{CaCO3}{L}=A X B X\frac{50000}{ml Sample}$$

where:

 A = mL NaOH titrant used,

 B = normality of NaOH,

**Referencess**

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# ****TOTAL ALKALINITY****

**1. Theory**

Alkalinity of a water is its acid-neutralizing capacity. It is the sum of all the titratable bases. The measured value may vary significantly with the end-point pH used. Alkalinity is a measure of an aggregate property of water and can be interpreted in terms of specific substances only when the chemical composition of the sample is known.

Alkalinity is significant in many uses and treatments of natural waters and wastewaters. Because the alkalinity of many surface waters is primarily a function of carbonate, bicarbonate, and hydroxide content, it is taken as an indication of the concentration of these constitutents. The measured values also may include contributions from borates, phosphates, silicates, or other bases if these are present. Alkalinity in excess of alkaline earth metal concentrations is significant in determining the suitability of a water for irrigation. Alkalinity measurements are used in the interpretation and control of water and wastewater treatment processes. Raw domestic wastewater has an alkalinity less than, or only slightly greater than, that of the water supply. Properly operating anaerobic digesters typically have supernatant alkalinities in the range of 2000 to 4000 mg calcium carbonate (CaCO3)/L.

a. Principle: Hydroxyl ions present in a sample as a result of dissociation or hydrolysis of

solutes react with additions of standard acid. Alkalinity thus depends on the end-point pH used.

b. End points: When alkalinity is due entirely to carbonate or bicarbonate content, the pH at

the equivalence point of the titration is determined by the concentration of carbon dioxide (CO2) at that stage. CO2 concentration depends, in turn, on the total carbonate species originally present and any losses that may have occurred during titration.

‘‘Phenolphthalein alkalinity’’ is the term traditionally used for the quantity measured by titration to pH 8.3 irrespective of the colored indicator, if any, used in the determination. Phenolphthalein or metacresol purple may be used for alkalinity titration to pH 8.3. Bromcresol green or a mixed bromcresol green-methyl red indicator may be used for pH 4.5.

c. Interferences: Soaps, oily matter, suspended solids, or precipitates may coat the glass electrode and cause a sluggish response. Allow additional time between titrant additions to let electrode come to equilibrium or clean the electrodes occasionally. Do not filter, dilute, concentrate, or alter sample.

d. Selection of procedure: Titrate at room temperature with a properly calibrated pH meter or electrically operated titrator, or use color indicators.

 **2. Reagents**

a. Standard sulfuric acid or hydrochloric acid, 0.02N: Dilute 200.00 mL 0.1000N standard

acid to 1000 mL with distilled or deionized water.

b. Methyl orange indicator solution, pH 4.5 indicator: Dissolve 100 mg methyl orange, in 100 mL distilled water.

**3. Apparatus**

a. Buret.

b. Titration vessel

c. Magnetic stirrer.

d. Pipets, volumetric.

e. Flasks, volumetric.

**4. Procedure**

 Potentiometric titration to preselected pH:

Prepare sample and titration assembly. Titrate to the end-point pH without recording intermediate pH values and without undue delay. As the end point is approached make smaller additions of acid and be sure that pH equilibrium is reached before adding more titrant.

**Calculations**



**References**

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