CHAPTER - IV

PHYSICO-CHEMICAL ANALYSIS OF WATER
IV. PHYSICO-CHEMICAL ANALYSIS OF WATER

(A) INTRODUCTION

The quality of water is vital concern for mankind, since it is directly linked with human welfare. Generally, the functions directly relate to their physical, chemical and biological integrity. Water quality evaluation for wetlands leads to information about their misuse by indicating the pollution status. The quality of aquatic life depends on the water quality and a thorough assessment of the water quality is an integral part of wetland evaluation.

In order to characterize any water body, studies on physico-chemical and biological characteristics should be carried out. The physical and chemical properties of a fresh water body are characterized by the climatic, geochemical, geomorphological and pollution conditions. The biota in the surface water is governed entirely by various environmental conditions. The primary production of organic matter is in the form of phytoplankton and macrophytes which are more intense in lakes and reservoirs than in rivers. In contrast to the chemical quality of water bodies which can be measured by suitable analytical methods, the biological quality is measured by both qualitative and quantitative characterization.

Disturbances in the ecosystem are mostly the result of excessive use, misuse and mismanagement of biosphere resources which form a major part of earth's ecosystem. The human activities aimed at development are responsible for disturbances and thus are the major
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causes of pollution. In a broad sense, any "unwanted and undesirable change in physical, chemical and biological characters of water, land or air is called pollution". As a result of pollution, human life, useful organisms, living conditions, environment, raw materials, etc. are damaged and may even be destroyed permanently.

Physical properties of water in any aquatic system are largely regulated by the existing meteorological conditions and chemical properties. The effect of physical factors like light and heat is of great limnological significance as they are solely responsible for thermal and chemical stratification, diurnal and seasonal variations in the number and distribution of microorganism. The assessment of the chemical criteria of the water body helps in

i. evaluating the chemicals, that cause toxicity to aquatic life,

ii. studying the long-term effects on the ecosystem and

iii. conducting the status and monitoring of wetlands resources by studying their physico-chemical and biological parameters. In many cases, planktons also act as the biological indicators of pollution being very sensitive to change in water quality.

Chemical properties of water not only alter the physical properties of the medium but also have a significant influence on the distribution and metabolic activities of the life forms which in turn, tend to change the chemical quality of water in due course of time. The important parameters selected for this study are Hydrogen ion concentration (pH),

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Temperature, Chloride, Alkalinity, Total Hardness, Calcium Hardness, Magnesium Content, Free Carbon dioxide (Free CO₂), Dissolved Oxygen (DO), Biochemical Oxygen Demand (BOD), Chemical Oxygen Demand (COD), Phosphate and Nitrate, were analyzed following the standard methods (APHA 1985, 1989, Adoni 1985 and NEERI 1988).
(B) MATERIALS AND METHODS

Water samples were collected in sterile glass bottles from the selected sites of Sagar lake and brought to the laboratory for physico-chemical analysis. The experiment usually started at 8 am in the morning and stations were visited by boat. Precautions were made to minimize the disturbance in water around the sites while collecting the samples.

HYDROGEN ION CONCENTRATION (pH)

It is the measure of the relative acidity or alkalinity which represents the negative logarithm of the concentration of free hydrogen ions in a solution measured by systronic digital pH meter 335.

TEMPERATURE

It was measured with the help of precalibrated thermometer having a least count of 0.5 and a range of 50 °C.

CHLORIDE

Chloride in natural water usually present in low concentration and high concentration is considered to be the indicator of pollution which is either due to organic wastes of animal origin or industrial effluents. Chloride is titrated with soluble silver salt (in weak acid solution) in presence of chromate yielding relatively insoluble AgCl. The brick red silver chromate is formed at the end point (iodide and bromide, when present get registered as chloride).
To 50 mL of sample aliquot 5 drops of potassium chromate indicator was added, colour of the sample turned yellow. The treated sample was titrated with standard 0.0141 N silver nitrate titrant until brick red colour (end point) is obtained.

\[
\text{Chloride in mg/L}^{-1} = \frac{\text{mL of titrant used} \times N \times 35.46 \times 10^3}{\text{mL of sample}}
\]

\(N\) = normality of titrant. Results are expressed as chlorides mg/L\(^{-1}\).

Preparation of chemicals and reagents

(i) **Silver nitrate titrant (0.0141 N)**

Dissolve 2.395 g of silver nitrate in distilled water and dilute to 1 litre. Store in dark bottle.

(ii) **Potassium chromate indicator**

Dissolve 10 g potassium chromate in a little distilled water. Add silver nitrate solution to produce red precipitate and then kept overnight, filter and dilute to 200 mL with distilled water.

ALKALINITY

In most natural water bicarbonates and sometimes carbonates are present in appreciable amounts. Their salts get hydrolyzed, produced hydroxyl ions, consequently raising the pH. Alkalinity is determined by titrating the sample with a standard solution of strong acid. Alkalinity due to hydroxide and carbonate is determined to the first end point (pH 8.3)
using phenolphthalein indicator and bicarbonate alkalinity is determined to
the second end point (pH 4.5) using methyl orange indicator.

50 mL of sample was taken in a 150 mL Erlenmeyer flask, two drops of phenolphthalein indicator was added, pink colour appeared which was then titrated with acid titrant to a colourless end point, later two drops of methyl orange was added in the same flask and titration was continued till the colour changed to orange.

\[
\text{Total alkalinity (T) as mg/L}^{-1} \text{ CaCO}_3 = \frac{\text{mL of titrant 't' x 1000}}{\text{mL of sample}}
\]

Results are expressed as the total alkalinity and contribution of bicarbonates (HCO\textsubscript{3}^{-}), carbonates (CO\textsubscript{3}^{2-}) and hydroxide (OH\textsuperscript{-}) mg/L\textsuperscript{-1} CaCO\textsubscript{3}.

**Preparation of chemicals and reagents**

(i) **Sulphuric acid titrant (0.02 N)**

Prepare stock solution 0.1 N by diluting 2.8 mL of conc. Sulphuric acid to 1 litre, dilute 200mL of this stock solution to 1 litre to obtain 0.02 N acid titrant, standardize it.

(ii) **Phenolphthalein indicator**

Dissolve 1.25 g phenolphthalein in 125 mL ethyl alcohol and add 125 mL distilled water. Add 2 drops of 0.02 N NaOH until a faint pink colour appears.

(iii) **Methyl orange indicator**

Dissolve 0.1 g methyl orange in 200 mL of distilled water.
TOTAL HARDNESS

Total hardness in water is the sum of the concentration of alkaline earth metal cations (e.g. Ca\(^{++}\) and Mg\(^{++}\)). In most fresh water, nearly all the hardness is imparted by the calcium and magnesium ions which are in combination with bicarbonates and carbonates (temporary hardness) apart from sulphates, chlorides and nitrates. Determination of this parameter helps in deciding the quality of water.

Eriochrome black T forms wine red complex compound with metal ions (Ca\(^{++}\) and Mg\(^{++}\)). The di-sodium salt of EDTA extracts the metal ions from the dye-metal ion complex as colourless chelate complexes leaving a blue coloured aqueous solution of the dye.

50 mL of sample was taken in a flask, to it, 1 mL of ammonia buffer and five drops of indicator was added. The colour of the sample turned wine red which was titrated with EDTA solution until a clear blue colour appeared.

\[
\text{Total hardness as mg/L}\text{ }\text{CaCO}_3 = \frac{\text{mL of titrant used } \times 1000}{\text{mL of sample}}
\]

Results expressed as the total hardness in mg/L\textsuperscript{−1} CaCO\(_3\).

Preparation of chemicals and reagents

(i) **Standard EDTA titrant (0.01 M)**

Dissolve 3.723 g of di-sodium salt of EDTA in distilled water to prepare 1 litre of titrant. Store in plastic bottle.
(ii) **Ammonia buffer solution**

Add 114 mL conc. NH₄OH to 13.5 g NH₄Cl and make the volume up to 200 mL.

(iii) **Eriochrome black - T indicator**

Dissolve 0.5 g dye in 100 mL of 80% ethyl alcohol.

**CALCIUM HARDNESS** $\text{mg/L}^{-1} \text{CaCO}_3$

Calcium is essential for all organisms, being an important cell wall constituent and regulates various physiological functions in animals too. It has a direct effect on pH and carbonate system. Murexide indicator forms pink coloured complexes with Ca ions with the addition of the di-sodium salt of EDTA. The Ca$$^{2+}$$ forms a colourless chelate complex leaving behind a purple solution of the dye.

50 mL sample was taken in a flask. One mL of 8% sodium hydroxide solution and a pinch of murexide indicator was added to it. Colour of sample turned to salmon pink. This was titrated with standard EDTA solution until the colour changed to purple.

Calcium hardness as $\text{mg/L}^{-1} \text{CaCO}_3 = \frac{\text{mL of titrant} \times 1000 \times 1.05}{\text{mL of sample}}$

Calcium hardness as $\text{mg/L}^{-1} \text{Ca} = \frac{\text{mL of titrant} \times 400.5 \times 1.05}{\text{mL of sample}}$

Calcium hardness expressed as $\text{mg/L}^{-1} \text{CaCO}_3$ and also as $\text{mg/L}^{-1} \text{Ca}$. 
Preparation of chemicals and reagents

(i) **Standard EDTA titrant (0.01 M)**
Dissolve 3.723 g of di-sodium salt of EDTA in distilled water to prepare 1 litre of titrant. Store in polythene bottle.

(ii) **Sodium hydroxide solution (8%)**
Dissolve 16 g NaOH in distilled water and dilute to 200mL.

(iii) **Murexide indicator mixture**
It is prepared by grinding thoroughly 0.2g ammonium purpurate with 100 g sodium chloride (AR grade).

**MAGNESIUM CONTENT**
Magnesium forms the nucleus of the porphyrin ring of the chlorophyll and therefore is an essential constituent of all the primary producers. If the calcium hardness as calcium carbonate is known, the magnesium content can be calculated as follows:

\[
\text{Magnesium in mg/L}^{-1} = (\text{Total hardness} - \text{Calcium hardness}) \times 0.243
\]

**FREE CARBON DIOXIDE (Free CO}_2\)**
Free carbon dioxide dissolved in water (CO}_2 aq.) is essentially the only source of carbon that can be assimilated and incorporated into the 'skeleton' of the living matter of all the aquatic autotrophs, once fixed it can further be utilized by the organisms of other categories. In absence of free CO}_2, plants utilize the bicarbonates. Carbon dioxide dissolved in natural water actively participates in the carbonate system.
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\[
\begin{align*}
\text{CO}_2 \ (g) & \quad = \quad \text{CO}_2^- \quad \text{(aq.)} \\
\text{CO}_2 + \text{H}_2\text{O} & \quad = \quad \text{H}_2\text{CO}_3 \\
\text{H}_2\text{CO}_3 & \quad = \quad \text{HCO}_3^- + \text{H}^+ \\
\text{HCO}_3^- & \quad = \quad \text{CO}_3^{2-} + \text{H}^+ \\
\text{CO}_2 + \text{OH}^- & \quad = \quad \text{HCO}_3^- \\
\text{H}_2\text{O} & \quad = \quad \text{H}^+ + \text{OH}^- \\
\end{align*}
\]

It is the normal practice to distinguish free carbon dioxide as the concentration of \( \text{CO}_2 + \text{H}_2\text{CO}_3 \) which is estimated by titrating the sample with standard alkali titrant to \( \text{pH} \) 8.3.

50 mL of sample was taken in a flask and 2 drops of phenolphthalein indicator was added (if slight pink colour develops then free carbon dioxide is absent). If the solution remains colourless, tritate with standard alkali titrant to slight pink end point.

\[
\text{Free CO}_2 \ \text{mg/L}^{-1} = \frac{\text{mL of titrant} \times 1000}{\text{mL of sample}}
\]

Results expressed as free \( \text{CO}_2 \ \text{mg/L}^{-1} \)

Preparation of chemicals and reagent

(i) **Standard sodium hydroxide titrant (0.02273 N)**

Dissolve 0.909 g sodium hydroxide in carbon dioxide free (freshly boiled and cooled) distilled water and raise the volume to 1 litre. Standardize it with oxalic acid (always use freshly prepared titrant).
Table - 4A: Physico-chemical characteristics of Sagar lake from July 2001 to December 2001

<table>
<thead>
<tr>
<th>S.No.</th>
<th>Month</th>
<th>Site</th>
<th>pH</th>
<th>Temp (°C)</th>
<th>Chloride*</th>
<th>Alkalinity*</th>
<th>Total Hardness (CaCO₃)*</th>
<th>Calcium hardness*</th>
<th>Mg* content</th>
<th>Free CO₂*</th>
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SITE I: Chakra Ghat  
II: Temple Site  
III: Dhobi Ghat  
* : mg/L⁻¹
Table - 4B : Physico-chemical characteristics of Sagar lake from July 2001 to December 2001

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<th>S.No.</th>
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<th>Nitrate*</th>
<th>DO*</th>
<th>BOD*</th>
<th>COD*</th>
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</table>

Site I : Chakra Ghat
II : Temple Site
III : Dhobi Ghat

* : mg/L⁻¹

DO : Dissolved oxygen
BOD: Biochemical oxygen demand
COD: Chemical oxygen demand
(ii) Phenolphthalein indicator

Dissolve 1.25 g phenolphthalein in 125 mL ethyl alcohol and add 125 mL distilled water. Add 0.02 N NaOH drop wise until a faint pink colour appears.

DISSOLVED OXYGEN (DO)

The occurrence of dissolved oxygen in water may be mainly attributed to direct diffusion from the air and photosynthetic evolution by aquatic autotrophs. Eutrophic water bodies have a wide range of dissolved oxygen.

Oxygen combines with manganous hydroxide to form higher hydroxides, which on acidification liberate iodine equivalent to that of oxygen fixed. This iodine is titrated with standard sodium thiosulphate solution using starch as an indicator.

\[
\begin{align*}
\text{MnSO}_4 + 2\text{KOH} & \longrightarrow \text{Mn(OH)}_2 + \text{K}_2\text{SO}_4 \\
\text{Mn(OH)}_2 + \text{O} & \longrightarrow \text{MnO(OH)}_2 \\
\text{MnO(OH)}_2 + 2\text{H}_2\text{SO}_4 + 2\text{KI} & \longrightarrow \text{MnSO}_4 + \text{K}_2\text{SO}_4 + 3\text{H}_2\text{O} + \text{I}_2
\end{align*}
\]

The sample was collected in 250 mL BOD bottle without bubbling. 2 mL manganous sulphate and 2 mL alkaline iodide azide reagent was added to it. The bottle was stoppered and shaken vigorously to mix the contents well. The precipitate was allowed to settle down. Then, 2 mL conc. sulphuric acid was also added. Bottle was restoppered.
and mixed by inverting it several times until the precipitate dissolves completely. 50 mL aliquot of this treated sample was titrated with standard sodium thio-sulphate solution (0.025 N) to a pale straw colour. Then a few drops of starch indicator was added and the titration was continued till the solution turned colourless.

Since, 1 mL of 0.025 N sodium thiosulphate solution is equivalent to 0.2 mg oxygen:

\[
\text{DO in mg/L}^{-1} = \left( \frac{8 \times 1000 \times N}{V} \right) V
\]

Where,

\[
v = \text{Volume of sample (mL)}
\]

\[
V = \text{Volume of titrant used (mL)}
\]

\[
N = \text{Normality of the titrant.}
\]

Results expressed as DO in mg/L\(^{-1}\).

Preparation of chemicals and reagents

(i) **Manganous sulphate solution**

Dissolve 182 g MnSO\(_4\).H\(_2\)O in distilled water, filter and dilute to 500 mL.

(ii) **Alkaline iodide azide solution**

Dissolve separately 350 g KOH and 75 g KI in distilled water. Mix and make the volume upto 500 mL. Dissolve separately 5 g of sodium azide (Na\(_3\)N) in 20 mL of distilled water. Add this azide solution to the alkaline iodide reagent.
(iii) **Sodium thiosulphate titrant (0.025 N)**

Dissolve 6.205 g sodium thiosulphate (AR grade) in freshly boiled and cooled distilled water and dilute to 1 litre. Add one pellet of NaOH as preservative.

(iv) **Starch indicator**

Dissolve 1 g starch (soluble) in 200 mL distilled water and add few drops of toluene as preservative.

(v) **Sulphuric acid concentrated**

sp.gr.1.84; 18 M.

(vi) **Sodium sulphite solution (5%)**

Dissolve 5 g Na₂SO₃ in 100 mL distilled water (to get oxygen free solution, DO = zero).

**BIOCHEMICAL OXYGEN DEMAND (BOD)**

Biochemical oxygen demand is the amount of dissolved oxygen required in milligrams per litre for stabilizing the biodegradable organic matter by microorganisms of the sample under aerobic conditions in a stated time. The BOD of a water sample is the decrease in oxygen concentration after incubation in the dark at a certain temperature for a period of time.

Two BOD bottles were filled with water sample avoiding air bubbles. One bottle was immediately fixed for its initial dissolved oxygen content (D₁), the other bottle was incubated in dark at 28±2 °C for 3 days.
(D₂), after incubation the DO content of the second bottle analyzed and was considered as final dissolved oxygen

\[ \text{BOD}_3 \text{ in mg/L}^{-1} = D_1 - D_2 \]

Where,

- \( D_1 \) = initial DO in sample (mg/L⁻¹)
- \( D_2 \) = DO after 3 days incubation (mg/L⁻¹)

Results expressed as \( \text{BOD}_3 \) mg/L⁻¹.

(Formula of measuring dissolved oxygen already mentioned).

**Preparation of chemicals and reagents**

*(i)*  **Sulphuric acid (1 N)**

Add 2.8 mL of conc. sulphuric acid to 100 mL distilled water.

*(ii)*  **Sodium hydroxide (1 N)**

Add 4.0 g sodium hydroxide to 100 mL distilled water.

**CHEMICAL OXYGEN DEMAND (COD)**

The amount of organic matter in water is estimated by their oxidability by chemical oxidants such as potassium permanganate or potassium dichromate (the constituent carbon and hydrogen are oxidized and not the nitrogen). In the permanganate method, the organic matter is first oxidized with a known amount of KMnO₄ and then the excess of oxygen is allowed to react with potassium iodide to liberate iodine in amounts equal to the excess oxygen which is estimated titrimetrically with sodium thiosulphate solution using starch as an indicator.
It was analyzed titrimetrically by potassium permanganate digestion method using standard 0.1 M sodium thiosulphate as a titrant and starch as an indicator (Adoni 1985).

To 50 mL sample aliquot in an Erlenmeyer flask, 5 mL of 0.1 N potassium permanganate reagent was added and the flask was placed in a boiling water-bath for 1 h. Then the flasks were cooled for 10 min. Later 5 mL of 10% potassium iodide reagent and 10 mL of 2 M sulphuric acid were added and then titrated with 0.1 M sodium thiosulphate reagent using starch as an indicator. A glass distilled water blank was run with each set.

\[
\text{COD in mg/L}^{-1} = \frac{8 \times C \times (A-B)}{S} \times 0.23
\]

Where,

\( C = \) Concentration of titrant (m mol \( L \)^{-1})

\( A = \) Volume of titrant used for blank (mL)

\( B = \) Volume of titrant used for sample (mL)

\( S = \) Volume of water sample (mL)

Results expressed as mg/L^{-1}.

**Preparation of chemicals and reagents**

(i) **Potassium permanganate solution (0.1 N)**

Dissolve 3.16 g KMnO\(_4\) and 16 g NaOH in glass distilled water free from COD to make 1 litre.
(ii) **Sodium thiosulphate solution (0.1 M)**

Dissolve 15.81 g of Na$_2$S$_2$O$_3$ in COD free distilled water and dilute to 1 litre. Add 1 g of Na$_2$CO$_3$ as preservative.

(iii) **Sulphuric acid (2 M)**

To 900 mL distilled water add cautiously 100 mL of 18 M sulphuric acid.

(iv) **Potassium iodide solution (10%)**

Dissolve 10 g KI in 90 mL distilled water.

(v) **Starch solution (1%)**

Add 1 g soluble starch in 100 mL distilled water and heat until it dissolves.

**PHOSPHATE**

Phosphorous is one of the most important major nutrients required by biota. However, in comparison to other major elements it is required in smaller amounts. Phosphorous is abundant in the sediments but its movements to the over lying water largely governed by the physico-chemical conditions and under aerobic state the equilibrium is towards the sediment site.

Orthophosphate (PO$_4$-P) in an acidified ammonium molybdate solution produces blue colour when stannous chloride is added. This colour is measured by spectrophotometer at 690 nm.
25 mL sample was taken in an Erlenmeyer flask (also prepare a distilled water blank simultaneously). 1 mL ammonium molybdate reagent was added and the flask was shaken vigorously after which 0.12 M (3 drops) of stannous chloride reagent was added and the flask was shaken again. The intensity of blue colour that developed after 10 min but before 12 to 15 min against a blank was measured at 690 nm and compared with a calibration curve.

Preparation of chemicals and reagents:

(i) **Ammonium molybdate strong acid solution**

Dissolve 5 g ammonium molybdate \((\text{NH}_4)_5\text{MO}_7\text{O}_{24}.4\text{H}_2\text{O}\) in 35 mL distilled water. Cautiously add 62 mL con. \(\text{H}_2\text{SO}_4\) to 80 mL distilled water. Cool, add the molybdate solution and dilute to 200 mL.

(ii) **Stannous chloride solution**

Dissolve 0.5 g fresh \(\text{SnCl}_2.2\text{H}_2\text{O}\) in 2 mL conc. \(\text{HCl}\), dilute to 20 mL with distilled water. Use freshly prepared solution whenever required.

(iii) **Standard phosphate solution**

Dissolve 0.1757 g of potassium dihydrogen phosphate (oven dried at 105°C) in distilled water and dilute to 1 litre, 1 mL of this solution contains 40 \(\mu\)g PO₄.

**NITRATE**

Nitrates is the most oxidized form of nitrogen and is an important plant nutrient. Due to its higher mobility as compared to other vital nutrients, its concentration in fresh water apart from autochthonous production is largely regulated by waste water loading, agricultural runoff and ground water inputs.
The reaction between nitrate and 1,2,4 phenoldisulphonic acid produces 6-nitro-1,2,4 phenoldisulphonic acid which on conversion to the alkaline salt yields yellow colour. Wavelength required for determination is 410 nm.

25 mL of sample in a casserole was evaporated on a hot water bath to dryness, after cooling residue was rubbed and 0.5 mL phenoldisulphonic acid reagent (use spatula), 5 mL distilled water and 1.5 ml conc. NH₄OH (or 12 N KOH) was added and stirred. The intensity of yellow colour developed was measured at 410 nm against a distilled water blank and compared with a calibration curve prepared with KNO₃.

Results expressed as NO₃ - N in mg/L⁻¹.

Preparation of chemicals and reagents

(i) **Phenoldisulphonic acid**

Dissolve 25 g pure white phenol in 150 mL conc. H₂SO₄. Add cautiously 75 mL fuming sulphuric acid (15% free SO₃), stir well, and heat on a water bath for 2 h.

(ii) **Potassium hydroxide solution (12 N)**

Dissolve 336.5 g KOH in distilled water and dilute to 500 mL.

(iii) **Ammonium hydroxide (32% w/v NH₃)**

It is available as lab reagent in the market.

(iv) **Standard potassium nitrate solution**

Dissolve 7.22 g anhydrous potassium nitrate in nitrate free distilled water and make upto 1 litre. 1 mL of this solution contains 1 mg NO₃ ions.
(C) RESULTS AND DISCUSSION

HYDROGEN ION CONCENTRATION (pH)

In the present investigation, the pH value ranged from 8.4 to 8.9 at SITE-I, 7.1 to 8.7 at SITE-II and 7.0 to 7.7 at SITE-III (Table 4A and Fig.3). High pH (8.9) was recorded in July at SITE-I while low pH (7.0) in December was recorded at SITE-III. Thus when comparing the results, SITE-I and SITE-II were found to be alkaline while SITE-III was neutral. Adoni (1975) reported the variation of pH in surface water of the lake which ranged from 8.2 to 9.8 exhibiting a general pattern of low values in winter and high values in summer. Similar observation was also reported by Yadav (1986) and Gupta (1987).

The measurement of pH provides an important means to understand the chemical conditions prevailing in natural water. This gives an idea of the concentration of the ionized hydrogen which in turn yields indirect information of free CO₂ content, alkalinity, dissolved O₂, residual pH, dissolved solids and thus may serve as the test for several environmental conditions (Wetzel 1975). It is one of the most important parameter in water chemistry and is defined as -log (H⁺) and measured as intensity of acidity or alkalinity on a scale ranging from 0-14. If free H⁺ is more, it is expressed acidic (i.e. pH<7) and if more OH⁻ ions is expressed as alkaline (i.e. pH>7). It tends to increase during day largely due to the photosynthetic activity. Waste water and polluted natural water have pH values lower or higher than 7.
Fig. 3: Variation in pH of different sampling sites of Sagar lake from July 2001 to December 2001

Fig. 4: Variation in Temperature (°C) of different sampling sites of Sagar lake from July 2001 to December 2001
based on the nature of the pollutant. The higher values at surface were probably associated with the photosynthetic activity.

The water of Sagar lake remained alkaline throughout the study period and this may be due to abundance of buffering substances brought with rain water from the adjoining hills (Moore 1981). Generally the pH of the water varies between 6.5 to 8.5. Higher pH occur regularly in water of arid regions and also temporarily in hard water lakes due to vigorous photosynthesis. Value of pH below 7 occurs in seepage tanks and lakes of acid igneous rocks.

Hydrogen ion concentration of water is linked to the species composition and their life processes. As Sagar lake has alkaline pH the blue green algae form a major plankton flora (Adoni 1975). Pandoy et al., (1991) and Bhatt and Hedge (1997) observed that the pH of the water appears to be dependent upon the relative quantities of calcium carbonates and bicarbonates. The water tends to be more alkaline when it possesses carbonates and less alkaline when it contains large quantities of bicarbonates, carbon dioxide and calcium. Pandey et al., (2003) reported alkaline pH of sub surface water of Makrana area from 18 different sites.

TEMPERATURE

In the present study water temperature ranged from 19 to 26 °C at SITE-I, 20 to 26 °C at SITE-II and 19 to 26 °C at SITE-III (Table 4A and Fig.4). Maximum temperature of 26 °C in October was recorded at
all the three sites while minimum temperature of 19 °C in December was recorded at SITE-I and SITE-III. The studies of thermal condition of this lake reveal that it is a homothermal lake. The homothermal condition of Sagar lake is due to the shallowness of the lake and severe wind action. Another important factor in this regard is that the lake is only rainfed and has no underground or surface inlet or any perennial streams joining it. The surface water temperature ranged between 16.5 to 29.4 °C (Adoni 1975). Yadav (1986) reported that the surface water temperature varied from 20 °C to 34 °C during summer, 27.5 to 29.5 °C during rainy and from 17.2 to 22.0 °C during winter. The water temperature of Sagar lake ranged from 17.0 to 30.5°C, and minimum temperature was recorded in February and maximum in May (Ghosh 1986). Gupta (1987) also observed that the value of water temperature ranged from 18.0 to 33.3 °C.

Air and water temperature plays a crucial role in physico-chemical, metabolic and physiological behaviour of aquatic ecosystems. According to the modified classification of lakes (Whipple 1927), Sagar lake can be placed under the tropical lakes. The lake had a surface water temperature of more than 4 °C. Hutchinson and Loffer (1956) classified lakes as warm polymictic (tropical lakes with frequent periods of turn over). Water is characterized by several unique thermal properties. The change in water temperature is found to be lesser than in air (Pandoy et al., 1991 and Bhattacharya et al.,1997). Generally the water temperature is very low in the morning but with the advance of the day it gradually started rising and reaching to its maximum in the afternoon.
The surface water temperature of lake was always found to be more than the bottom water but was always less than the air temperature. The change in water temperature is confined to the upper surface layer due to its colour and transparency. At times due to low transparency temperature of the bottom water remained considerably less than the upper layers. Vyas (1968) observed that during rainy season the water temperature decreased due to frequent clouds, high humidity, high velocity, increased turbidity and high water level. Temperature is the prime factor in determining the seasonal changes of flora and fauna (Born 1979).

CHLORIDE

Chloride contents ranged from 60.29 mg/L to 96.29 mg/L at SITE-I, 68.29 mg/L to 141.99 mg/L at SITE-II and 110.29 mg/L to 189.99 mg/L at SITE-III (Table 4A and Fig.5). Highest chloride content (189.99 mg/L) in November was recorded at SITE-III while lowest chloride content (60.29 mg/L) in August was recorded at SITE-I. On the basis of above results it was concluded that during the study period high value of chloride content was observed from SITE-III in comparison to SITE-I and SITE-II which recorded lower values. The chloride content tend to increase during summer and thereafter decreased during rainy season (Gupta 1987). Similar observation was reported by Yadav (1986). The chloride content ranged from 20.2 to 48.0 ppm with minimum and maximum values during September and May respectively (Adoni 1975).
Fig. 5: Variation in Chloride content (mg/L$^{-1}$) of different sampling sites of Sagar lake from July 2001 to December 2001

Fig. 6: Variation in Alkalinity (mg/L$^{-1}$) of different sampling sites of Sagar lake from July 2001 to December 2001
Chlorides are relatively conservative and undergo minor spatial and temporal variations due to least biotic utilization or biotically modified changes in the environment (Wetzel 1975). Pollution sources like industries and municipal wastes can greatly modify the natural concentration of chlorides (Ownbey and Kee 1967, Pandoy et al., 1991, Bhattacharya et al., 1997).

The presence of chloride in natural water can mainly be attributed to dissolution of salt deposits in the form of ions (Cl⁻) otherwise high concentrations may indicate pollution by sewage, industrial waste, intrusion of sea water or other saline water. It is the major form of inorganic anions in water for aquatic life. High chloride content has a deleterious effect on metallic pipes and structures as well as agricultural plants.

The weathering of the soil and rocks, atmospheric precipitation and environmental factors directly influence the salinity (Moss and Moss 1969). The presence of chloride ions in high quantities in natural water without any specific geological reason indicate the extent of organic pollution especially of animal origin (Joshi 1987). Munwar (1970) suggested that higher levels of chloride in water is an index of pollution of animal origin. Bhattacharya et al., (1997) directly correlated chloride concentration with pollution.

**ALKALINITY**

Alkalinity ranged from 7.2 mg/L⁻¹ to 23.2 mg/L⁻¹ at SITE-I, 12.0 mg/L⁻¹ to 48.6 mg/L⁻¹ at SITE-II and 31.2 mg/L⁻¹ to 63.2 mg/L⁻¹ at SITE-III (Table 4A and Fig.6). Highest alkalinity (63.2 mg/L⁻¹) in October
was recorded at SITE-III while lowest value (7.2 mg/L\(^{-1}\)) was found in July at SITE-I. The bicarbonate alkalinity values of Sagar lake water ranged from 42.02 to 108.0 ppm during summer 42.0 to 78.0 ppm during rainy and 58.0 to 130.0 ppm during winter (Yadav 1986) while Gupta (1987) observed variation of bicarbonate alkalinity from 44.0 ppm in October to 122.0 ppm in December. The carbonate alkalinity was also reported which ranged from 8 ppm in October to 88 ppm in March. Ghosh (1986) reported that the bicarbonate alkalinity ranged from 6.0 mg/L\(^{-1}\) in April to 118.0 mg/L\(^{-1}\) in July of Sagar lake.

Alkalinity of water refers to the quantity and kind of compounds present. The alkalinity shifts the pH to the alkaline side of neutrality. It is usually imparted by the presence of bicarbonates, carbonates, hydroxides, borates, silicates and phosphates (Wetzel 1975).

The alkalinity of the lake water is maintained and does not seem to be affected either by oxygen or by carbon dioxide content (Singh 1960 and Verma 1969). This is probably due to effective buffering capacity of water which avoids large variations. Below pH 5, total CO\(_2\) is of quantitative importance while pH between 7 to 9 bicarbonates are of greatest significance and above pH 9.5 carbonates are of biological importance.

**TOTAL HARDNESS**

Total hardness ranged from 129.3 mg/L\(^{-1}\) to 167.3 mg/L\(^{-1}\) at SITE-I, 120.0 mg/L\(^{-1}\) to 172.2 mg/L\(^{-1}\) at SITE-II and 137.30 mg/L\(^{-1}\) to 178.6 mg/L\(^{-1}\) at SITE-III (Table 4A and Fig.7). Maximum hardness (178.6 mg/L\(^{-1}\))
Fig. 7: Variation in Total hardness (mg/L\(^{-1}\)) of different sampling sites of Sagar lake from July 2001 to December 2001

- **Month**
  - July
  - August
  - September
  - October
  - November
  - December

- **Total hardness (mg/L\(^{-1}\))**
  - 0.00
  - 20.00
  - 40.00
  - 60.00
  - 80.00
  - 100.00
  - 120.00
  - 140.00
  - 160.00
  - 180.00
  - 200.00

- **SITE-I**
- **SITE-II**
- **SITE-III**
was recorded in July at SITE-III while minimum value (120.0 mg/L$^{-1}$) was recorded in October at SITE-II. Gupta (1987) reported that total hardness ranged from 86.0 ppm in June to 142.4 ppm in November while Ghosh (1986) reported 81.0 mg/L$^{-1}$ in September and 108.0 mg/L$^{-1}$ in December. Yadav (1986) reported minimum value in summer and maximum value of total hardness in rainy season.

Hardness is predominantly caused by divalent cations (calcium and magnesium) and alkaline earth metals (iron, manganese and strontium). The total hardness is defined as the sum of calcium and magnesium concentration both expressed as CaCO$_3$ (mg/L$^{-1}$). Carbonates and bicarbonates of calcium and magnesium cause temporary hardness while sulphate and chloride cause permanent hardness (Banerjee 1991, Mishra 1991).

Hardness is frequently used to express the quality of water. According to (Wetzel 1975) hard water contains large conc. of alkaline earth derived from the rainage of calcareous deposits. Moss (1973) reported the increase in hardness with increased eutrophication. Singh (1979) also observed a positive correlation between calcium, total hardness and magnesium content.

**CALCIUM HARDNESS**

The calcium hardness in all the sites during the study period ranged from 43.26 mg/L$^{-1}$ to 89.4 mg/L$^{-1}$ as CaCO$_3$ and 17.32 mg/L$^{-1}$ to 35.82 mg/L$^{-1}$ as Ca (Table 4A and Fig.8A and 8B). The higher concentration
Fig. 8A: Variation in Calcium hardness as CaCO₃ (mg/L⁻¹) of different sampling sites of Sagar lake from July 2001 to December 2001

Fig. 8B: Variation in Calcium hardness as Ca (mg/L⁻¹) of different sampling sites of Sagar lake from July 2001 to December 2001
Chapter IV

Physico-chemical analysis of water

89.4 mg/L$^{-1}$ as CaCO$_3$ in December was recorded at SITE-III while lower 43.26 mg/L$^{-1}$ concentration as CaCO$_3$ in July was recorded at SITE-I. Calcium hardness ranged from 43.26 mg/L$^{-1}$ to 78.86 mg/L$^{-1}$ as CaCO$_3$ and 17.32 mg/L$^{-1}$ to 38.78 mg/L$^{-1}$ as Ca at SITE-I, 63.0 mg/L$^{-1}$ to 86.73 mg/L$^{-1}$ as CaCO$_3$ and 25.23 mg/L$^{-1}$ to 34.73 mg/L$^{-1}$ as Ca at SITE-II and 60.9 mg/L$^{-1}$ to 89.4 mg/L$^{-1}$ as CaCO$_3$ and 24.39 mg/L$^{-1}$ to 35.82 mg/L$^{-1}$ as Ca at SITE-III. During study period it was found that the value of calcium hardness as CaCO$_3$ and Ca was recorded highest at SITE-III in comparison to SITE-I and SITE-II. Gupta (1987) observed that calcium hardness ranged from 31.5 ppm in May to 85.05 ppm in December and at times the low value of calcium hardness in summer was also reported. Ghosh (1986) reported that calcium hardness ranged from 36.7 mg/L$^{-1}$ in August to 65.1 mg/L$^{-1}$ in December. Yadav (1986) found minimum value of calcium hardness in summer while maximum during rainy season.

The presence of calcium in water results from passage through deposits of limestone, dolomite, gypsum and other calcium bearing rocks. Calcium contributes to the total hardness of water and is an important micronutrient of aquatic environment which is especially needed in large quantities by mollusks and vertebrates. Small concentration of calcium carbonate prevents corrosion of metal pipes by forming a protective coating. Increased amount of calcium precipitates on heating and forms harmful scales in pipes and utensils (Ramanujula et al., 1992). Munawar (1970) observed high values of calcium during winter and summer. It is important in buffering of lake water and in the precipitation

MAGNESIUM CONTENT (HARDNESS)

The magnesium content in all the study sites ranged from 9.97 mg/L\(^{-1}\) to 30.14 mg/L\(^{-1}\) (Table 4A and Fig.9). It ranged from 14.27 mg/L\(^{-1}\) to 30.14 mg/L\(^{-1}\) at SITE-I, 9.97 mg/L\(^{-1}\) to 26.32 mg/L\(^{-1}\) at SITE-II and 11.93 mg/L\(^{-1}\) to 24.26 mg/L\(^{-1}\) at SITE-III. Higher concentration (30.14 mg/L\(^{-1}\)) in July was recorded at SITE-I while lower concentration (9.97 mg/L\(^{-1}\)) in October was recorded at SITE-II. Yadav (1986) observed that magnesium content in surface water of Sagar lake ranged from 6.56 ppm to 18.83 ppm. Minimum and maximum value of magnesium content was observed in the rainy season while Ghosh (1986) reported that magnesium content ranged from 6.1 mg/L\(^{-1}\) in September to 14.6 mg/L\(^{-1}\) in May. The calcium and magnesium hardness started increasing from July onwards. This may be probably due to initial phase of leaching and drainage following first monsoon (Joshi 1987). Munawar (1970) and Rao (1971) have also found a direct correlation between magnesium content and organic matter of the water bodies.

Magnesium is a relatively abundant element in the earth’s crust ranking eighth in abundance among the elements. It is found in all natural water and its source lies in rocks and generally present in lower concentration than calcium. It is also an important element contributing to
Fig. 9: Variation in Magnesium content (mg/L$^{-1}$) of different sampling sites of Sagar lake from July 2001 to December 2001
hardness and a necessary constituent of chlorophyll. Its concentration greater than 125 mg/L\(^{-1}\) can influence cathartic and diuretic actions. High concentration of magnesium proves to be diuretic and laxative and reduce the utility of water for domestic use while concentration above 500 mg/L\(^{-1}\) imparts an unpleasant taste to water and renders it unfit for drinking. Chemical softening, reverse osmosis and electro dialysis or ion exchange reduces the magnesium hardness to acceptable levels (Patel 1991, Mahal and Datta 1997).

FREE CARBON DIOXIDE (FREE CO\(_2\))

Throughout the investigation period it was found that free CO\(_2\) was absent in all the months except October (Table 4A). 2.0 mg/L\(^{-1}\) from SITE-I, 7.2 mg/L\(^{-1}\) from SITE-II and 1.8 mg/L\(^{-1}\) from SITE-III was recorded. Maximum value was recorded from SITE-II while minimum value was recorded at SITE-III. Throughout the period of investigation free CO\(_2\) was never detected in surface water except in stray samples. It was detected only in a few samples collected in December and January which ranged between 0.0 to 2.7 ppm (Adoni 1975). Yadav (1986) reported that in Sagar lake free CO\(_2\) rarely appeared in surface water, its presence during September, November and December may be due to death and decay of macrophytes. The absence of free CO\(_2\) during the major part of the study can be attributed to its complete utilization in photosynthetic activity. This lake differs in this respect from many Indian lakes and reservoirs. Sreenivasan (1964) also recorded total absence of free CO\(_2\) in surface and bottom water.
The important source of free CO$_2$ in surface water bodies is mainly due to respiration and decomposition by aquatic organisms. It reacts with water to form calcium bicarbonate and in the absence of free CO$_2$ they get converted to carbonates releasing carbon dioxide (Patel 1991, Kaur et al., 1996).

Free carbon dioxide present in atmosphere combines with the rain water to form carbonic acid which dissociates into H$^+$ and HCO$_3^-$ ions thus giving acidic nature to water. The carbonate present in the lake are converted into bicarbonates in the presence of free CO$_2$. The total absence of free CO$_2$ is due to its complete utilization in photosynthetic activity (Sreenivasan 1965, Sahai and Sinha 1969). In natural water reduction of CO$_2$ is carried out by photosynthetic activity, removal by the marlforming organisms, agitation of water, evaporation and rise of bubbles from the depths. Carbonates are appreciably low in bottom water when compared to surface water.

DISSOLVED OXYGEN (DO)

During the investigation dissolved oxygen in all the study sites of Sagar lake ranged from 1.2 mg/L$^{-1}$ to 2.9 mg/L$^{-1}$ (Table 4B and Fig.10). The DO ranged from 1.6 mg/L$^{-1}$ to 2.9 mg/L$^{-1}$ at SITE-I, 1.2 mg/L$^{-1}$ to 2.7 mg/L$^{-1}$ at SITE-II and 1.6 mg/L$^{-1}$ to 2.4 mg/L$^{-1}$ at SITE-III. Higher DO (2.9 mg/L$^{-1}$) was recorded at SITE-I in August while lower DO (1.2 mg/L$^{-1}$) in September was recorded at SITE-II. Ghosh (1986) noticed that the values of DO in Sagar lake ranged from 2.0 mg/L$^{-1}$ in March to 17.2 mg/L$^{-1}$
Fig. 10: Variation in Dissolved oxygen (mg/L\(^{-1}\)) of different sampling sites of Sagar lake from July 2001 to December 2001
in September. Gupta (1987) also reported the DO which ranged from 6.8 to 20.0 ppm. Maximum value (20.0 ppm) was observed in August while 6.8 ppm in September. The DO value decreased from late winter to summer season and showed to be in rising from late rainy season. In Sagar lake the DO in surface water varied from 5.20 ppm to 21.00 ppm. Values of DO were found much higher during summer as compared to those in winter and rainy season (Yadav 1986).

Dissolved oxygen is a very important parameter in water analysis as it serves as an indicator of the physical, chemical and biological activities of the water body. The two main sources of dissolved oxygen are (i) diffusion of oxygen from the air and (ii) photosynthetic activity. Diffusion of oxygen from the air into water depends on the solubility of oxygen and on the influence of many other factors (water movement, temperature, salinity etc). Photosynthesis is a biological phenomenon carried out by the autotrophs which depends on the plankton population, light condition, gases etc. Oxygen is considered as the major limiting factor in water bodies (Ghosh 1997, Pandey et al., 1991).

Atmospheric and hydrostatic pressure affects the solubility of gases in water. The saturation of oxygen is considered in relation to the pressure at the surface of the lake. Salinity also reduces the solubility of oxygen in water. Adequate supply of dissolved oxygen is essential for the physiological processes of the aquatic organisms having aerobic mode of metabolism, thus the dynamics of oxygen distribution, behaviour and physiological growth of aquatic organisms are affected (Ghosh 1986).
Joshi 1987 and Hutchinson 1957). Hutchinson (1957) remarked that "a skillful limnologist can probably learn more about the nature of the lake from the change in oxygen content than any other kind of chemical data".

**BIOCHEMICAL OXYGEN DEMAND (BOD)**

Biochemical oxygen demand ranged from 11.0 mg/L$^{-1}$ to 21.3 mg/L$^{-1}$ at SITE-I, 11.2 mg/L$^{-1}$ to 22.1 mg/L$^{-1}$ at SITE-II and 14.5 mg/L$^{-1}$ to 23.0 mg/L$^{-1}$ at SITE-III (Table 4B and Fig.11). Higher concentration of BOD (23.0 mg/L$^{-1}$) in July was recorded at SITE-III while lower concentration (11.0 mg/L$^{-1}$) in October was recorded at SITE-I. It was also found that SITE-III had the highest BOD values in comparison to the other sites. Higher value of BOD in summer was mainly attributed to the high biological activity and high temperature whereas the lowest BOD in winter indicated low biological activity and low temperature (Rai 1978). Fokmare and Musaddiq (2003) recorded that high BOD value of water was due to organic enrichment. The BOD is the amount of oxygen required by microorganisms for stabilizing biologically decomposable organic matter (carbonaceous) in water under aerobic conditions. The least BOD values are used to determine the pollution and the efficiency of wastewater treatment (Shaw et al., 1991, Bhattacharya et al., 1997).

**CHEMICAL OXYGEN DEMAND (COD)**

In the present studies, the COD ranged from 21.1 mg/L$^{-1}$ to 48.1 mg/L$^{-1}$ at SITE-I, 21.2 mg/L$^{-1}$ to 51.6 mg/L$^{-1}$ at SITE-II and 20.5 mg/L$^{-1}$ to 55.2 mg/L$^{-1}$ at SITE-III (Table 4B and Fig.12). Higher COD value (55.2
Fig. 11: Variation in Biochemical oxygen demand (mg/L$^{-1}$) of different sampling sites of Sagar lake from July 2001 to December 2001

Fig. 12: Variation in Chemical oxygen demand (mg/L$^{-1}$) of different sampling sites of Sagar lake from July 2001 to December 2001
mg/L\(^{-1}\)) in December was recorded at SITE-III while lower COD value (20.5 mg/L\(^{-1}\)) at SITE-III in October.

Chemical oxygen demand (COD) is the measure of oxygen equivalent to the organic content of the sample that is susceptible to oxygenation by a strong chemical oxidant. The intrinsic limitation of the test lies in its ability to differentiate between the biologically oxidisable materials with inert materials (Shaw et al., 1991, Panday et al., 1991).

**PHOSPHATE**

Phosphate ranged from 0.03 mg/L\(^{-1}\) to 0.29 mg/L\(^{-1}\) at SITE-I, 0.03 mg/L\(^{-1}\) to 0.24 mg/L\(^{-1}\) at SITE-II and 0.05 mg/L\(^{-1}\) to 0.35 mg/L\(^{-1}\) at SITE-III (Table 4B and Fig.13). Higher phosphate content (0.35 mg/L\(^{-1}\)) in July was recorded at SITE-III while lower phosphate content (0.03 mg/L\(^{-1}\)) in November at SITE-I and SITE-II. Yadav (1986) reported that in Sagar lake the phosphate content ranged from 5.60 ppm to 200.1 ppm. He also reported minimum content in summer while maximum during rainy season. Ghosh (1986) observed that phosphate content ranged from 14.0 \(\mu\)g/L\(^{-1}\) in January to 484.0 \(\mu\)g/L\(^{-1}\) in July, while Gupta (1987) found that phosphate content ranged from 21.8 \(\mu\)g/L\(^{-1}\) in January to 483.0 \(\mu\)g/L\(^{-1}\) in July from Sagar lake. The phosphate used to be found mainly in surface water of Sagar lake although the values remained very low throughout the study. The high values were seen during rainy season and this was certainly due to addition of phosphate through drainage and sewage by the arrival of rain wash (Adoni 1975).
Fig. 13: Variation in Phosphate content (mg/L\(^{-1}\)) of different sampling sites of Sagar lake from July 2001 to December 2001

Fig. 14: Variation in Nitrate content (mg/L\(^{-1}\)) of different sampling sites of Sagar lake from July 2001 to December 2001
Phosphates occur in natural or waste water as orthophosphates, condensed phosphates and naturally found phosphates. Their presence in water is due to detergents, used boiler water, fertilizers and biological processes. They are essential for the growth of organisms and its presence limits the primary productivity of the water body (Jayangoudar 1964). Inorganic phosphorous play a dynamic role in aquatic ecosystems and when present in low concentration it forms the most important nutrient but when in excess along with nitrates and potassium, causes algal blooms (Somasekar et al., 1991)

In nature total phosphorous occurs as inorganic, dissolved, organic and particulate phosphorous. The only significant form of inorganic phosphorous is orthophosphate which is assimilated rapidly by the biota. Natural origin of phosphorous in aquatic system is mainly due to the erosion, the chemical and mechanical weathering of rocks, human excreta, detergents, urbanization, leaching and drainage of fertilizers and other soil nutrients (Barrette 1953). These factors also tend to increase total phosphate concentration. Phosphorus acts as a limiting nutrient in lakes of various types (Schindler et al., 1971). Forkmare and Musaddiq (2003) observed high values of phosphate in Purna river water which can be attributed to pollution from detergents and domestic wastes.

NITRATE

Nitrate content ranged from 3.1 mg/L\(^{-1}\) to 10.1 mg/L\(^{-1}\) at SITE-I, 4.8 mg/L\(^{-1}\) to 8.2 mg/L\(^{-1}\) at SITE-II and 4.3 mg/L\(^{-1}\) to 8.3 mg/L\(^{-1}\) at SITE-III (Table 4B and Fig.14). Maximum nitrate content (10.1 mg/L\(^{-1}\)) in
October at SITE-I while minimum value (3.1 mg/L$^{-1}$) in November at SITE-I was also recorded. Ghosh (1986) reported that nitrate content of Sagar lake ranged from traces to 1090.0 μg/L$^{-1}$ in June while Yadav (1986) reported that surface water of Sagar lake had a nitrate value of 20 ppm to 1090 ppm. Maximum value was observed during summer while minimum during rainy season.

Nitrites are the most oxidized forms of nitrogen and it is the end product of the aerobic decomposition of organic nitrogenous matter. The significant sources of nitrites are chemical fertilizers from cultivated lands, drainage from livestock feeds, as well as domestic and industries (Sondergaard and Jensen 1979). Natural water in their unpolluted state contains only minute quantities of nitrites. The plant growth is stimulated by nitrites, thus causing algal eutrophication. The subsequent death and decay of these plants cause secondary pollution (Hannan and Young 1975). Nitrites are important for biological oxidation of nitrogenous organic matter. Certain nitrogen fixing bacteria and algae have the capacity to fix molecular nitrogen into nitrites. The main source of polluting nitrites is through domestic sewage. Nitrites may find their way into ground water through leaching from soil and also by contamination (Somasekhar et al., 1991). Vollenweider (1968) concluded that in general, nitrate and nitrogen are easily available than ammonium while nitrites are either absent or present (traces) in well aerated water.

The important nitrogen sources in lake water are in the form of NO$_3$, NH$_4$ and N$_2$. The phytoplankton can only utilize NO$_3$ and NH$_4$ easily. The entry of elemental nitrogen into the aquatic habitat is entirely from the atmosphere.