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Study Amount of Dissolved Salts and Its Effect on Quality of Groundwater

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Abstract. In the Diyla area, fourteen ground water samples were obtained for this study in high density polyethylene (HDPE) bottles that had been previously cleaned and filled with diluted HNO_3. To prevent cross-contamination, the bottles were rinsed with the same water that would be used for the sample. Using normal procedures, each sample was examined for five parameters: TDS, Ca, Mg, No3, So4, and F (APHA, 1995). Using the Trimble Juno 3D Handheld, the sampling sites' geographic locations are gathered. Electrical Conductivity (EC) was used to compute Total Dissolved Solids (TDS). Spectronics 21 (Model BAUSCH AND LOMB) was used to measure nitrates (NO3-) and sulphates (SO4-2). Standard AgNO3 was used to estimate chloride (Cl-), and an Orion ion analyzer with an ion-selective electrode was used to analyse fluoride (F-). With the exception of pH and EC, all the parameters are presented in mg/l. The unit of EC expression is micro Siemens/cm.

Keyword :

INTRODUCTION

An increasingly important problem for water resource management in semiarid areas of the world is the increasing salinization of surface and groundwater [1]. Large sedimentary aquifers' soil fluids and groundwaters contain dissolved sodium chloride, which makes up a significant portion of the salt in Australia's terrain. Understanding the origins of salt and the processes by which it moves through the hydrosphere is essential for the long-term preservation and management of water resources that may be impacted by salt. The mobilisation of significant amounts of stored salts in groundwater and soil water through changes in hydrological balances is the ultimate cause of the issue, despite the fact that the most noticeable form of salinization is evident in saline streams and surficial efflorescences in shallow waterlogged depressions in the landscape [1-3]. Therefore, a key component of salinity amelioration at the watershed scale is comprehending the origins and evolution of solutes. This study focuses on the mechanisms governing the chemical evolution of the dissolved salts and the formation of solutes in the semi-arid Murray Basin groundwater systems in southeast Australia. Since European colonisation more than 150 years ago, the basin has experienced significant changes in land use, including the removal of native vegetation and its replacement by pastures and crops, as well as significant surface diversion for irrigation. The resultant salinization of streams and drylands as a result of higher water tables and greater recharge has been extensively studied [4-6]. Two primary topics are covered in this paper: First, where do solutes, especially Na and Cl, come from in the end? Secondly, how does the salinity of groundwater change from being comparatively fresh to becoming a salinity that is comparable to sea water as it moves from the basin boundary to the discharge zones? A similar sequence of main ion development from low-salinity Ca–Mg–HCO3 type groundwater to very saline Na–SO4–Cl groundwater over a hydraulic gradient is reported in many investigations of regional aquifer systems [7-9]. In sub-humid to arid basins around the world, low relief, little to no surface drainage, and high rates of evaporation can concentrate atmospherically derived salts to very high concentrations in soil water that eventually recharges the aquifer [12, 13]. These may be caused by reaction with aquifer minerals, which in turn may be related to increasing residence time of groundwater [10, 11].

MATERIALS AND METHOD

A. Sample Collection

In the Diyla area, fourteen ground water samples were gathered in high density polyethylene (HDPE) bottles that had been previously cleaned and filled with diluted HNO3. To prevent cross-contamination, the bottles were rinsed with the same water that would be used for the sample. Using normal procedures, each sample was examined for four parameters: TDS, NO3-, SO4-2, and Cl- (APHA, 1995). A digital pH metre (Model 802 Systronics) was used in the field to measure pH, and a conductivity metre (Model 304 Systronics) was used to assess EC. Using the Trimble Juno 3D Handheld, the sampling sites' geographic locations are gathered.

B. Sample Analysis

Electrical Conductivity (EC) was used to compute Total Dissolved Solids (TDS). An instrument called Spectronics 21 (Model BAUSCH AND LOMB) was used to measure nitrates (NO3-) and sulphates (SO4-2). Fluoride (F-) was measured using an Orion ion analyzer equipped with an ion-selective electrode, whereas Chloride (Cl-) was approximated using standard AgNO3. With the exception of pH and EC, all the parameters are presented in mg/l. The unit of EC expression is micro Siemens/cm.

C. Total Dissolved Solids (TDS)

It gauges how much salt has dissolved in the water. The main components of TDS in groundwater are usually Ca, Mg, Na, HCO3, Cl, SO4, K, F, and NO3. The unit of measurement is mg/l. Rainwater usually contains more than 10 mg/l of TDS. Its content in groundwater can vary from 1,000 mg/l in regions with heavy rainfall to more than a million mg/l in desert brines. There is more than 3,000,000 mg/l of TDS in brine areas. The TDS content of groundwater can rise as a result of a number of events, including movements through rocks that contain soluble mineral matter, concentration through evaporation and concentration brought on by an influx of seawater, and the dumping of industrial, agricultural, and municipal waste.

D. Measurement Of Total Dissolved Solids Procedure

1. Clean the filter paper.
2. Weigh & use a dry evaporating dish
3. Mix the specimen.
4. Pipette 50 ml and stir.
5. Wash and filter three times.
6. Pour filtrate into an evaporating dish and pat dry.
7. Let cool and weigh 8. Determine the mg/L.

E. Calculating total dissolved solids concentration:

TDS mg / l = (A - B) X 1000 / ml sample

Where

A = Weight of dried residue + dish, mg

B = Weight of dish, mg.

F. Chloride (Cl-) Principle

100 mL of high-quality sample, or an appropriate fraction diluted to 100 mL, is required for the determination of chloride. Potassium chromate is used as an indicator for titrating standard silver nitrate in a naturally occurring or slightly alkaline solution to measure the chloride content. Red silver chromate is created quantitatively after the formation of silver chloride. The following list contains the chemical processes that this approach involves:

1. Ag^{++} Cl^{-} ® AgCl (White precipitate)
2. $2Ag^{++}$ CrO_4^{--}® Ag_2CrO_4 (Red precipitate)

G. Apparatus

1. Porcelain dish 200mL
2. Pipettes
3. Burettes
4. Glass rod

H. Reagents and standards

1. One method to test for potassium chromate is to dissolve 50g of K2Cr2O7 in

distilled water. AgNO3 should be added until a clear red precipitate forms. Let stand for twelve hours. Dilute to 1000 mL after filtering.

2. Silver nitrate, 0.0141N: Dilute 2.395g of AgNO3 to 1000mL by dissolving it. Standardise 1 mL of 0.0141N AgNO3 = 0.5 mg Cl- and compare to NaCl, 0.0141N.
3. 0.0141N sodium chloride: dissolve 824.1 mg of dried at 40°C sodium chloride and dilute to 1000 mL (1 mL = 0.5 mg Cl-).
4. Use a special reagent (125g AlK(SO4)2.12H2O) or AINH4(SO4)2.12H2O) to dissolve colour and turbidity, then dilute the mixture to 1000mL. Heat to 60°C and gradually add 55 mL of concentrated NH4OH. Allow to stand for one hour. Move the precipitate to a big bottle, then thoroughly mix each addition before decanting it with distilled water to remove any remaining chloride. A freshly made suspension takes up about one litre of space.

I. Calibration

It is necessary to standardise the silver nitrate solution against a 0.0141N sodium chloride solution. The silver nitrate solution's strength is indicated as 1 mL = 0.5 mg of chlorides, or Cl-.

J. Procedure

1. Add 1.0 mL of K2Cr2O7 to a 50 mL well-mixed sample that has been adjusted to a pH of 7.0–8.0.
2. Titrate using a normal AgNO3 solution until a pale red precipitate known as AgCrO4 begins to form.
3. Compare standard AgNO3 to normal NaCl
4. Titrate 50mL of distilled water the same way to establish a reagent blank for improved accuracy. A 0.2 to 0.3 mL blank is usual.

K. Calculation

Chloride mg/L as Cl^- = (A –B) x N x 35.45 x 1000 / mL sample

Where,

A = mL $AgNO_3$ required for sample

B = mL $AgNO_3$ required for blank, and

N = Normality of $AgNO_3$ used

L. Fluoride (F-) Principle

Any modern pH metre with an extended millivolt scale may detect the presence of fluoride ions, which establishes a potential when the fluoride electrode is dipped in the sample whose concentration needs to be measured. Using the suitable calibration curve, the fluoride ion selective electrode can be used to determine the fluoride activity or concentration in an aqueous sample. Fluoride activity, however, is contingent upon the sample's overall ionic strength. Fluoride that is complexed or bonded does not react with the electrode. These problems are solved by adding a buffer solution with a high total ionic strength that contains a chelate to compound aluminium preferentiality.

M. Apparatus and equipment

For accurate laboratory measurements, use an ion metre (field/laboratory mode) or a pH/mV metre. Fluoride-sensitive electrodes, a caromelle reference electrode, a magnetic stirrer, and plastic labware (samples and standards should always be stored in plastic containers as fluoride reacts with glass).

N. Reagents and standards

a. Dissolve 221 mg of anhydrous NaF and dilute to 1000 mL to get the stock fluoride solution. One millilitre is equal to one hundred grammes.-
b. Standard fluoride solution: To get 1 mL = 10µg F, dilute stock solution ten times with distilled water.-
c. The Total Ionic Strength Adjustment Buffer (TISAB) is made by adding 4g of 1, 2-cyclohexylenediamine tetraacetic acid, 57mL of glacial acetic acid, 58g of NaCl, and around 500mL of distilled water to a 1-liter beaker. To dissolve, stir. After submerging the beaker in a bath of cool water, gradually add 125 mL of 6N NaOH while stirring, bringing the pH up to 5.0–5.5. Move to a 1 litre volumetric flask and adjust the volume accordingly.

O. Calibration

Measure out 50 millilitres of the 1 ppm and 10 ppm fluoride standards. To calibrate the instrument, add 50 mL of TISAB (or 5 mL if conc. TISAB is being utilised).

P. Procedure

a. Refer to the manufacturer's instruction manual for instructions on attaching the electrodes to the metre and operating the instrument further.
b. Using an ion metre (59.16 mV for monovalent ions and 29.58 mV for divalent ions at 25°C), check the electrode slope.
c. Pour the sample into a 150 mL plastic beaker, using 50 to 100 mL. Include TISAB.
d. Clean the electrode, pat it dry, and insert it into the sample. After giving it a good stir, record the meter's consistent reading.
e. Adjust every one to two hours.
f. A straightforward method for measuring a big number of samples is direct

measurement. Samples and the standard should have the same temperature, and TISAB should be added to all solutions to bring the ionic strength of the samples and the standard to the same level.

g. A recognised addition process can be used to verify the findings of a direct measurement. A standard with a known concentration is added to a sample solution as part of the known addition technique. The initial concentration of the sample can be ascertained from the difference in electrode potential before and after addition.

Q. Calculation

The particular ion metre is the source of the concentration in mg/L.

R. Nitrate (NO3-) UV spectrophotometer method

The technique works best with water that is pure enough to drink and does not include any organic pollutants. The quick measurement of nitrate is made possible by measuring the UV absorption at 220 nm. Up to 11 mg/L N, the nitrate calibration curve conforms to Beer's law. The purpose of 1N hydrochloric acid acidification is to produce CaCO3 at concentrations of up to 1,000 mg/L or hydroxide interference. Chloride has no bearing on the outcome. NO32-N has a minimum detectable value of 40µg/L.

S. Principle

The absorbance at 420 nm in a sample containing 1 mL of hydrochloric acid (1N) in 100 mL of sample is used to calculate nitrate. From a graph of a standard nitrate solution in the range of 1–11 mg/L, the concentration is determined as N.

T. Apparatus and equipment

1. A spectrophotometer with matching silica cells of one centimetre or a longer light path, intended for usage at 420 nm and 275 nm.
2. Filter: You must use one of the options below.
   Filter assembly: 0.45µm membrane filter and suitable filter assembly
3. Paper: Hard-finish filter paper with an acid wash that is ashless and suitably retentive for fine precipitates.
4. Short form, 50 mL Nessler tubes.

U. Reagents and Standards

1. Redistilled water: For the preparation of all solutions and dilutions, use redistilled water.
2. Stock nitrate solution: Using distilled water, dissolve 721.8 mg of anhydrous potassium nitrate and dilute to 1000 ml. 100 µg N = 443 µg NO3- = 1 mL.
3. Standard nitrate solution: dilute 100mL stock nitrate solution to 1000mL with

distilled water. 10µg NO_3 N = 44.3µg NO_3 in 1 mL.

4. HCl, 1N hydrochloric acid solution.
5. Suspension of aluminium hydroxide: disperse 125g of potash alum in 1000mL of purified water. Heat to 60°C, then add 55–60 mL of NH_4OH and let it stand for one hour. After decanting the supernatant, repeatedly wash the precipitate to remove any remaining Cl, NO_2, and NO_3. Decant as much of the clean liquid as you can once it has set, leaving only the concentrated suspension.

V. Calibration

First, dilute 1, 2, 4, 7,..35 mL of the standard nitrate solution to 50 mL in order to prepare nitrate calibration standards in the range of 0 to 350µg N. The nitrate standards should be handled similarly to the samples.

W. Procedure

Compare the absorbance or transmittance to 100% transmittance or zero absorbance in redistilled water. For the nitrate measurement, use a wavelength of 220 nm, and if necessary, use a wavelength of 275 nm to measure interference from dissolved organic materials.

X. Calculation

For correction for dissolved organic matter, subtract 2 times the reading at 275nm from the reading at 220nm to obtain the absorbance due to nitrate. Convert this absorbance value into equivalent nitrate by reading the nitrate value from a standard calibration curve.

Nitrate N, mg/L = mg nitrate-N / mL of sample

NO_3^-, mg/L = Nitrate N mg/L x 4.43

Y. Pollution prevention and waste management

One crucial component of waste management is the elimination of trash produced during analysis. The analytical chemical is a micro-quantity (1 mL) of 1N HCl in 50 mL of sample, therefore waste management is not an issue.

Z. Sulphate (SO_4^{2-}) Turbidimetric method Principle

The determination of sulfate ions is done using this approach. Barium chloride ($BaCl_2$) and sulfate ion (SO_4^{--}) are precipitated in an acetic acid solution to produce uniformly sized crystals of barium sulphate ($BaSO_4$). The following is the reaction in question:

$BaSO_4$ (White suspension) ® Ba^{++} + SO_4. A photometer or a nephelometer is used to measure the $BaSO_4$ suspension's light absorbance or light scattering.

AA. Apparatus and equipment

1. Magnetic stirrer
2. Colorimeter for use at 420mm or turbidimeter/nephelometer
3. Stopwatch
4. Nessler tubes, 100mL
5. Measuring spoon 0.2 to 0.3mL

BB. Reagents and standards

a. Dissolve 30g of magnesium chloride, MgCl2.6H2O, and 5g of sodium acetate in buffer solution A. CH3COONa.To manufacture up to 1000mL, mix 3H2O, 1g potassium nitrate, KNO3, 20mL acetic acid, and CH3COOH (99%) in 500mL of distilled water.
b. Buffer solution B: this is necessary if the sample's sulphate (SO4--) content is below 10 mg/L. Dissolve 5g of sodium acetate (CH3COONa) and 30g of magnesium chloride (MgCl2.6H2O).Combine 3H2O, 1.0g potassium nitrate (KNO3), 0.111g sodium sulfate (Na2SO4), and 20mL (99%) acetic acid with 500mL of distilled water to create a 1000mL solution.
c. 20–30 mesh crystals of barium chloride.
d. Standard sulfate solution (dilute to 1000mL): dissolve 0.1479g anhydrous sodium sulphate (Na2SO4) in distilled water. One liter is equal to one hundred micrograms of SO4.

(OR)

Dilute 10.4mL standard 0.02 N H_2SO_4 titrant specified in Alkalinity, Section 10.1.3, with distilled water; 1mL = 100 µg SO_4^{2-}

CC. Calibration

Carry the standard sulphate solution through the entire process to prepare the standard curve. In the range of 0 to 40 mg/L, space standards at 5 mg/L increments. Read the mg SO42-resent from the standard curve for the sample.

DD. Procedure

A appropriate volume of sample should be taken and diluted to 100 mL in a 250 mL Erlenmeyer flask. Next, add 20 mL of buffer solution and well mix.Use the stirrer to ensure that the flask is continuously stirred. Stir in one spatula of BaCl2 crystals. After adding the BaCl2 d, stir for one additional minute.Transfer the suspension into a photometer's absorption cell and gauge the turbidity after 5 to 0.5 minutes.In order to adjust for sample color and turbidity, perform a blank experiment without adding BaCl2.

EE. Calculation

mg SO4--/L = {mg (SO4--) x 1000} / {mL of sample}

FF.Pollution prevention and waste management

Small amounts of the compounds are utilized. There are no substances that pose a health risk. Since no dangerous chemicals or substantial amounts of solvents are utilized, waste management is not really necessary. It is possible to protect water and land by reducing and regulating all discharges from bench activities if laboratory waste management practices are followed.

RESULTS AND DISCUSSION

A. Major Ions in Groundwater

Key quality metrics and ratios are found using the collected data. Because of the location and historical flow of water, salts are present in all groundwater. The kind and amount of salts in water are determined by the surrounding conditions, groundwater flow, and sources. As water moves from precipitation to groundwater throughout the hydrologic cycle, its quality varies. [14]. The chemical, physical, and bacteriological characteristics of groundwater determine its usefulness for industrial, residential, and irrigation uses. The examined main ions are addressed in light of the aforementioned information in order to determine their suitability for consistent use.

Table 1. Comparision Of Chemical Analysis Data Of Groundwater In The Study Area With Who (2004) Standards

S.NO	Water Parameters	Min	Max	Average	WHO(2011) Standards
		(mg/l)	(mg/l)	(mg/l)	(mg/l)
1	TDS	611	1980	955	500
2	Na^{2+}	131	400	209	250
3	K^+	1	4	2.3	10
4	Ca^{2+}	24	148	50	100
5	Mg^{2+}	15	90	31	50
6	HCO_3^-	311	512	398	500
7	Cl^-	60	545	207	250
8	SO_4^{2-}	35	190	87	250
9	F^-	0.63	2	1.02	1.5
11	TA	255	420	326	

Table 2. Major Element Analytical Data Of Groundwater In The Study Area (All the parameters are in mg/l except pH and EC in microsiemens/cm)

S.No.	Turbidity	Odour	PH	EC	TDS	TA	TH	Ca H	So4	Cl	F	Ca	Mg	Hco3	Na	K
1	8.5	U.O	7.8	1460	964	385	285	140	80	220	1.13	56	35	470	131	1.40

2	0.6	U.O	7.3	3000	1980	400	575	205	190	530	1.85	148	90	488	264	3.00
3	0.8	U.O	7.6	2860	1887	295	870	640	140	545	0.86	92	56	360	400	4.00
4	0.6	U.O	7.2	1800	1188	375	580	480	80	315	0.68	41	24	457	267	3.00
5	0.5	U.O	7.3	1384	913	260	460	355	118	185	0.63	42	26	317	212	3.00
6	0.6	U.O	7.3	1870	1234	270	600	485	80	315	0.69	46	28	329	276	2.00
7	0.7	U.O	7.2	1235	815	255	445	350	117	340	0.63	38	23	311	205	2.00
8	0.5	U.O	7.4	1089	719	285	400	305	140	120	1.09	38	23	348	184	2.00
9	0.6	U.O	7.3	970	640	300	360	265	51	80	0.81	38	23	366	166	2.00
10	0.7	U.O	7.3	1183	781	360	450	355	80	115	1.15	38	23	439	207	2.00
11	0.6	U.O	7.6	933	615	355	345	235	35	60	1.49	44	27	432	159	2.00
12	0.7	U.O	7.4	960	633	295	370	280	55	70	0.98	36	22	360	170	2.00
13	0.5	U.O	7.6	1004	663	325	375	290	58	75	0.99	34	21	396	173	2.00
14	0.5	U.O	7.3	926	611	310	340	280	46	65	0.84	24	15	378	156	2.00
15	0.6	U.O	7.3	1052	694	420	345	225	45	65	1.55	48	29	512	159	2.00
		Min	7.2	926	611	255	285	140	35	60	0.63	24	15	311	131	1.4
		Max	7.8	3000	1980	420	870	640	190	545	2	148	90	512	400	4
		Avg.	7.39	1448	955.8	326	453	326	87	207	1.02	50.9	31	398	209	2.3

B. Total Dissolved Solids (TDS):

It gauges how much salt has dissolved in the water. The main components of TDS in groundwater are usually Ca, Mg, Na, HCO3, Cl, SO4, K, F, and NO3. The unit of measurement is mg/l. Rainwater usually contains more than 10 mg/l of TDS. Its content in groundwater can vary from 1,000 mg/l in regions with heavy rainfall to more than a million mg/l in desert brines. There is more than 3,000,000 mg/l of TDS in brine areas. The TDS content of groundwater can rise as a result of a number of events, including movements through rocks that contain soluble mineral matter, concentration through evaporation and concentration brought on by an influx of seawater, and the dumping of industrial, agricultural, and municipal waste.

The study area's groundwater has an average TDS value of 955 mg/l, with variations ranging from 611 mg/l to 1980 mg/l. High TDS water, however, may make it less suitable for use as a drinking or irrigation source. High TDS levels can also affect the produced product's purity, color, and flavor. A TDS level of 500 mg/l is suitable for drinking water.

Tabel 3. Groundwater Classification According To The Tds Content

Water Type	TDS (mg/l)
Fresh water	<1000
Brackish water	1000 – 10,000

Salty water	10,000 – 1,00,000
Brine Water	>1,00,000

C. Sodium (Na+):

In water, sodium is the most prevalent cation. A content of sodium over 250 mg/l (WHO 2011) renders the water unfit for residential usage. People with heart, kidney, or circulation disorders may be at risk because to the increased sodium content. Elevated levels of Na+ in groundwater are ascribed to several factors such as mineral dissolution, ion exchange processes, sewage contamination, infiltration from agricultural runoff, and mining operations. The frequent usage of water in irrigated areas may also be the cause of high quantities of salt. Groundwater in the research area has an average sodium concentration of 209 mg/l, with a range of 131 mg/l to 400 mg/l.

D. Calcium (Ca2+):

The majority of igneous, metamorphic, and sedimentary rocks contain a significant amount of it. The main sources of calcium in groundwater include minerals found in sedimentary rocks such as limestone, dolomite, and gypsum, and in igneous and metamorphic rocks such as amphibole, pyroclase, and plagioclase. In addition, it is an absorbed ion in soils and rocks and is found as a cementing ingredient in sandstone and shale. Water containing 20 to 100 mg/l of Ca can dissolve in the presence of CO2. The exchange of Ca ions with Na ions during the ion exchange process can raise the Ca content of water. Ca levels in groundwater typically range from 10 to 100 mg/l. Brine has a maximum concentration of 75,000 mg/l. The calcium concentration in the study area ranges from 24 to 148 mg/l, with an average of roughly 50.9 mg/l.

E. Magnesium (Mg2+):

The primary sources of magnesium in groundwater include seawater, clay minerals, amphibolites and talk-tremolite schists (metamorphic rocks), dunites and pyroxenites basalts (igneous rocks), and dolomite (sedimentary rock). Like Ca content, groundwater's Mg concentration can rise by up to 200 mg/l due to CO2. According to their relative abundance in rocks, the Ca and Mg contents in groundwater are often higher. In groundwater, the average concentration of mg is less than 50 mg/l; however, in brines, the concentration can reach 57,000 mg/l. Healthy persons easily eliminate magnesium from their bodies. On the other hand, coma, confusion, hypertension, and muscle weakness can occur in those with kidney disease. Magnesium levels range from 15 mg/l to 90 mg/l at the minimum and highest, with an average content of 31 mg/l.

F. Potassium (K+):

Potash fertilizers, feldspathoids, clay minerals, certain micas, and feldspars (orthoclase and microcline) are major sources of potassium. Despite being both less and almost as prevalent as sodium in the minerals that make up the earth's crust. Its concentration in groundwater is typically 10 mg/l and seldom goes above 15 mg/l; in thermal and brine waters, it can be as high as 100 and 25,000 mg/l, respectively. In water, the potassium ion has no color. Hexachloroplatinic acid, sodium cobalt nitrite, and sodium tetraphenylborate are three precipitation-based techniques for potassium separation that are occasionally employed in gravimetric analyses. Potassium concentrations in the study region range from 1.4 to 4 mg/l at the lowest and greatest levels, with an average of 2.3 mg/l.

G. Chloride (Cl-):

For the most part, rocks have minimal levels of chloride. Through sprays caught in rain traps, ocean sprays provide groundwater with its Cl. Evaporation deposits, hot springs, volcanic gases, connate water, clay material, urban, agricultural, and industrial wastes, desert regions, and coastal locations are other sources of Cl in groundwater.

The research area's groundwater has an average content of 136 mg/l of Cl, with a minimum of 46 mg/l and a maximum of 461 mg/l detected in the bore well near Nyalkal and Zaheerabad town. As seen in Figure 4.10, these concentrations exceed the 200 mg/l allowable limit. Laxative effects could result from excess. Most metals corrode when exposed to high quantities.

H. Bicarbonate (HCO3-):

The dissolved CO2 in water is the main source of HCO3 ions in groundwater. When organic matter breaks down in the soil, CO2 is released for dissolution. HCO3 typically varies from 250 to 500 mg/l in groundwater. Agricultural practices or the seepage of domestic sewage are two possible causes of high alkalinity. The research area's HCO3 values range from 311 mg/l at the least to 512 mg/l at the maximum, with an average of 398 mg/l being recorded.

I. Sulphate (SO4):

Because of the oxidation of sulphide ores, atmospheric precipitation, water passing through lignite, coal, and gypsiferrous beds, low precipitation, insufficient drainage in mining areas, industrial wastes, and the application of sulphatic soil conditions, it is plentiful in groundwater. The research area's groundwater samples range in value from 35 to 190 mg/l, with an average of 87 mg/l.

J. Fluoride (F-):

An element that occurs naturally in some igneous and metamorphic rocks is fluoride. It may enter the ecosystem by naturally occurring erosive processes or land disturbances brought on by building and development. Fluoride is occasionally added to drinking water to assist prevent tooth decay in children. Fluoride is typically found in low concentrations in surface and groundwater sources. High amounts of fluoride (1.5 mg/l or more) have the potential to be harmful, however additions to drinking water seldom go over 1.5 to 2 mg/l.

Tooth discolorations and mottling may occur from concentrations higher than 2 mg/l, and skeletal deformities may arise from high dosages.

Water in fractured hard rock zones with pegmatite veins frequently has abnormally high fluoride levels. It coexists with the fluoride-rich mineral bearing rocks in the earth's crust. Groundwater fluoride is influenced by minerals such as topaz, fluorite, villuamite, fluor-apatite, cryolite, and fluoride replaceable hydroxide ion in ferromagnesium silicates [15]. These minerals' fluoride ions seep into groundwater and contribute to elevated fluoride levels. The average fluoride concentration in the study area is 1.02 mg/l, with a range of 0.63 to 1.85 mg/l.

CONCLUSSION

This study evaluated the groundwater quality in a specific area by analyzing the concentrations of major ions, including TDS, Na^+, Ca^{2+}, Mg^{2+}, K^+, Cl^-, HCO_3^-, SO_4^{2-}, and F^-. The results revealed significant variations in the chemical composition of groundwater, with some parameters exceeding the standards set by the WHO. Specifically, TDS, chloride, and sodium levels were found to be close to or above recommended limits, potentially limiting their suitability for domestic, irrigation, and industrial use. Fluoride concentrations were also near the safety threshold, posing potential health risks. These findings highlight the importance of continuous monitoring and management of groundwater quality to ensure its suitability for various applications.

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