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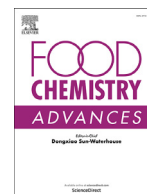


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Investigating pH and other electrical properties of potassium salt solutions

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ABSTRACT

A recent study of ours reported that pH of common salts varies with concentration at room temperature. It can either increase in case of basic salts or decrease in case of acidic salts. In this study, pH values of four kind of potassium salts were investigated and their pH values increased with salt concentrations. It was reasoned that the concentration of hydrogen ions equalled the concentration of positive charges/protons in a solution, so a relationship between the activity of hydrogen ions and charge activity was developed. Further, the well-established equations for pH with hydrogen-ion activity and electrostatic/cell potential were extended to charge activity. Potential energy, the product of charge activity and electrostatic/cell potential, was presented and discussed.

Introduction

pH is an essential parameter for quantifying acidity and alkalinity in solutions (Chakraborty et al., 2021; O'Sullivan et al., 2021) as well as in soils (Dong et al., 2023; Ghazali et al., 2020; Jin et al., 2022; Johansen et al., 2021; Kang et al., 2021; Li et al., 2021; Liu et al., 2021; Makungwe et al., 2021; Singh et al., 2021; Xu et al., 2020; Yao et al., 2022). Its widespread usage can be found in industries (Nguyen et al., 2023; Shiner et al., 2022), agriculture (Yang et al., 2021), medicine (Arafa et al., 2021; Bychkova et al., 2022, Emam et al., 2023; Ma et al., 2021; Mohammadzadeh et al., 2021; Park et al., 2023; Wei et al., 2022; Yan et al., 2022), food technology (Agalya et al., 2023; Geng et al., 2022; Kornet et al., 2023; Liu et al., 2022, Tan et al., 2021, 2022; Wang et al., 2022; Wu et al., 2022; Zhang et al., 2022, 2023), corrosion science (Ibrahimi et al., 2019; Wang et al., 2022), geotechnology, (Sadeghian et al., 2023) etc. When it comes to food industry, quality assurance is the utmost important factor as the food products that consumers buy should only yield positive impacts on their health. One of the key factors governing the quality of food products is pH as it helps to meet the above requirements by allowing the food industry to produce and maintain properties and regulatory requirements of the food products. For example, food industries producing milk and dairy products (cheese and yoghurt), fruit juices, marmalades and beer, wine, seafood (including the fresh consumption of fish, shellfish etc.) and bread all will have regulated pH requirements, for different production processes as well as in the final products, that allow the final products to be consistent in quality.

Salts are used in food production for preserving the food as well as processing agents or buffers. For example, potassium carbonate is used in chocolate making, dried fruit production, bakery and in pharmaceutical industry. It is also used as a processing agent in cooked products, alfalfa drying, brewing beer, oriental noodles and as additives in drying raisins. Further, it is used as a buffer in the production of fruit or wine (<https://persianutab.com/potassium-carbonate-in-the-food-industry/?lang=en>). Potassium nitrate is used as preservative for cured pork sausages and salami; potassium chloride is used as salt substitute for sodium chloride in food to reduce the risk of high blood pressure and potassium sulfate is used as additive in food to improve the taste and increase the shelf life. Thus, the presence of salts during the production processes and in the final products could alter the pH and cause undesirable qualities (if added at unsuitable concentrations). One unit change in the pH can increase or decrease the hydrogen ion concentration by 10 folds and therefore it is important to know how the changes in the concentrations of salts will impacts the pH. This study investigates the changes in pH of four salts (KCl, KNO₃, K₂CO₃ and K₂SO₄) with their concentrations so that the food industries can utilize this information while using those salts during their food production processes.

To understand the importance of pH values of some common salts, Shu et al. investigated pHs of NaCl, CaCl₂·2H₂O and Na₂CO₃ at room temperature and reported that the pH of all salts was related to their charge types and charge numbers as well as concentration. The pH of NaCl increased with salt concentration. CaCl₂·2H₂O with cation of two positive charges had a very low pH at high salt concentration. Na₂CO₃ with anion of two negative charges had a very high pH even at a concen-

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tration as low as 0.1 M (Shu et al., 2016). Additionally, water clusters shrink during salt dissolution, and it was shown that water molecules (H_2O) existed in bulk as water clusters, which became smaller and smaller once a salt particle entered them (Shu et al., 2020).

The relationship between pH in solution and electrostatic/cell potential of a pH electrode is an important factor as measuring pH involves comparing a potential of a solution of unknown hydrogen concentration $[H^+]$ with a known reference potential. pH meters convert the voltage ratio between a reference half-cell and a sensing half-cell to pH values (HACH, 2018). Reference half-cells contained a conductor (usually silver with a silver chloride coating) immersed in a solution with known $[H^+]$. The potential between this internal conductor and the known solution was constant, providing a stable reference potential. Sensing half-cells (measuring half-cells) were made of a nonconducting glass (or epoxy) tube sealed to a conductive glass membrane. Like the reference half-cell, the sensing half-cell also contained a conductor immersed in a buffered electrolyte solution, ensuring constant voltages on the inner surface of the glass membrane and the sensing conductor. When the pH electrode was immersed in the solution to be measured, a potential was established on the surface of the sensing glass membrane. If the unknown solution was neutral, the sum of fixed voltages on the inner surface of the glass membrane and on the sensing conductor approximately balanced the voltage on the outer surface of the glass membrane and the reference half-cell. This resulted in a total potential difference of 0 mV and a pH value of 7. In acidic or alkaline solutions, the voltage on the outer membrane surface changed proportionally to changes in $[H^+]$. The pH meter detects the change in potential and determined $[H^+]$ of the unknown by using the Nernst equation:

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} + 2.3(RT/nF) \cdot \log [a_{H^+}] \quad (1)$$

Where E_{cell} is measured potential difference; E°_{cell} is reference potential, a constant; R is the universal gas constant; T is temperature; n is ionic charge/number of electrons, $n = 1$; F is Faraday's constant; $[a_{H^+}]$ is hydrogen ion activity.

The theoretical gradient/slope and intercept are given by the Nernst equation:

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - 2.3(RT/nF) \cdot \text{pH} \quad (2)$$

Where the gradient/slope is equal to $-2.3RT/nF$ and E_{cell} is 0 mV at pH 7.00. With these conditions, Fig. 1 was drawn and the equation for calculating E_{cell} from pH was:

$$E_{\text{cell}} = -59.16\text{pH} + 414.12 \quad (3)$$

Where 414.12 was the intercept which is equal to E°_{cell} . Fig. 1 shows conversion of E_{cell} in mV and pH at 25 °C or 298.15 K. The pH of any solution was a function of its temperature. Voltage output from the electrode changes linearly in relation to pH, and the temperature of the solution determines the slope of the graph. One pH unit corresponded to 59.16 mV at 25 °C, the standard voltage and temperature to which all calibrations were referenced. The electrode voltage decreased to 54.20 mV/pH unit at 0.0 °C and increases to 74.04 mV/pH unit at 100.0 °C.

In physics, the product of charge and electrostatic/cell potential denotes the potential energy.

Materials and methods

Deminerlized water was used for all the tests. The salts used were of analytical grade. The pH meter had an error of 7.00 ± 0.02 . It was calibrated with standard buffer solutions every day before use according to the instructions given by the manufacturer. Solutions were mixed thoroughly before the determination of pH. In the case of saturated solutions, precipitates were allowed to settle on the bottom of a beaker and

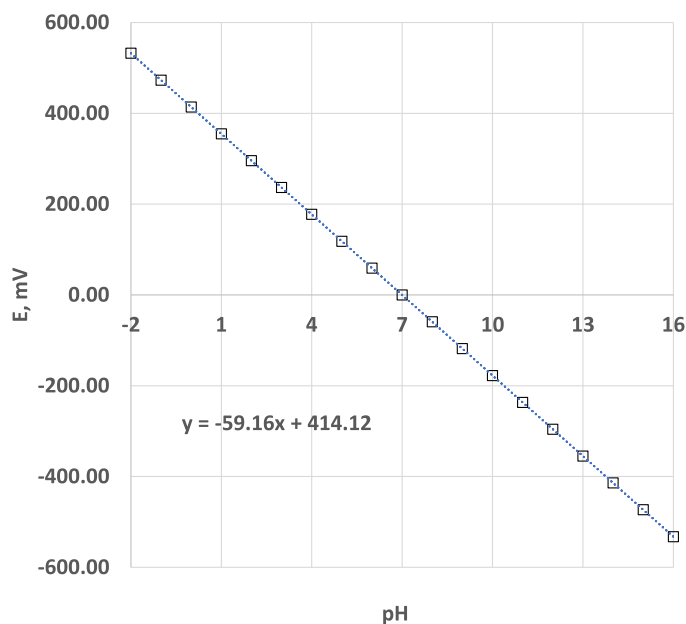


Fig. 1. Relationship between the electrostatic potential (E_{cell}) and pH at 25 °C.

the pH of the supernatant was measured. We prepared solutions with salt concentrations exceeding the saturation concentration as their pH were different. pH was measured in duplicates and the average value was taken. Tissue paper used was Kimberley Clark Kimwipes which are lint free wipes. We worked under an ambient temperature of 22.0 ± 2 °C in the wet laboratory at Deakin University, Geelong, Australia.

Results and discussion

pH of potassium salt solutions. pH values of four kinds of potassium salts, namely, KCl, KNO_3 , K_2CO_3 and K_2SO_4 were determined and displayed in Fig. 2. pH of pure water used in the study was approximately 5.7. Results showed that pH values of all salts increased with concentration, though at different rates. pH values of K_2CO_3 were the highest among the four of salts. They were 9.59 and 10.38 at salt concentrations of 0.0001 and 0.001 M respectively. The pH values had increased 0.79 units while the concentration increased 10-fold from 0.0001 to 0.001 M. The pH value was 10.92 at 0.01 M. The pH had increased by 0.54 when the concentration of the salt in the solution changed from 0.001 to 0.1 M. The pH values increased more rapidly when the concentration changed from 0.0001 to 0.001 M. The pH of K_2CO_3 increased linearly from 0.1 to 6 M according to the following expression: $\text{pH} = 0.4139$ (concentration of K_2CO_3 in M) + 11.48 ($R^2 = 0.9984$). K_2CO_3 is a basic salt with pH values more than 11 even at 0.1 M. pH of KCl solution was 5.62 at the salt concentration of 0.0001 M and 5.64 at 0.1 M. The pH was mostly unchanged even while the salt concentration increased 1000 times. The pH of KCl solution increased linearly from 0.1 to 6 M. The following equation describes the trend: $\text{pH} = 0.1246$ (concentration of KCl in M) + 5.6537 ($R^2 = 0.9794$). At 6 M the pH of KCl was 6.43. The increase in pH was 0.73 units compared to that of pure water without KCl addition. The pH increment due to the change in concentration from 0.0001 to 6 M was 0.81 units.

The pH of KNO_3 was 5.53 at a salt concentration of 0.0001 M and 5.68 at 0.1 M. The pH values increased 0.15 units while the salt concentration increased 1000 times. Between 0.1 and 4 M, pH values of KNO_3 could be expressed as 5.68 ± 0.01 , which was almost the same as the pure water's. At 6 M, the pH of KNO_3 was 5.74. The pH change of KNO_3 was 0.21 units when the concentration changed from 0.0001 to 6 M. The pH difference between KNO_3 solution at 6 M and the pure water was 0.04. The main increment happened before 0.1 M and af-

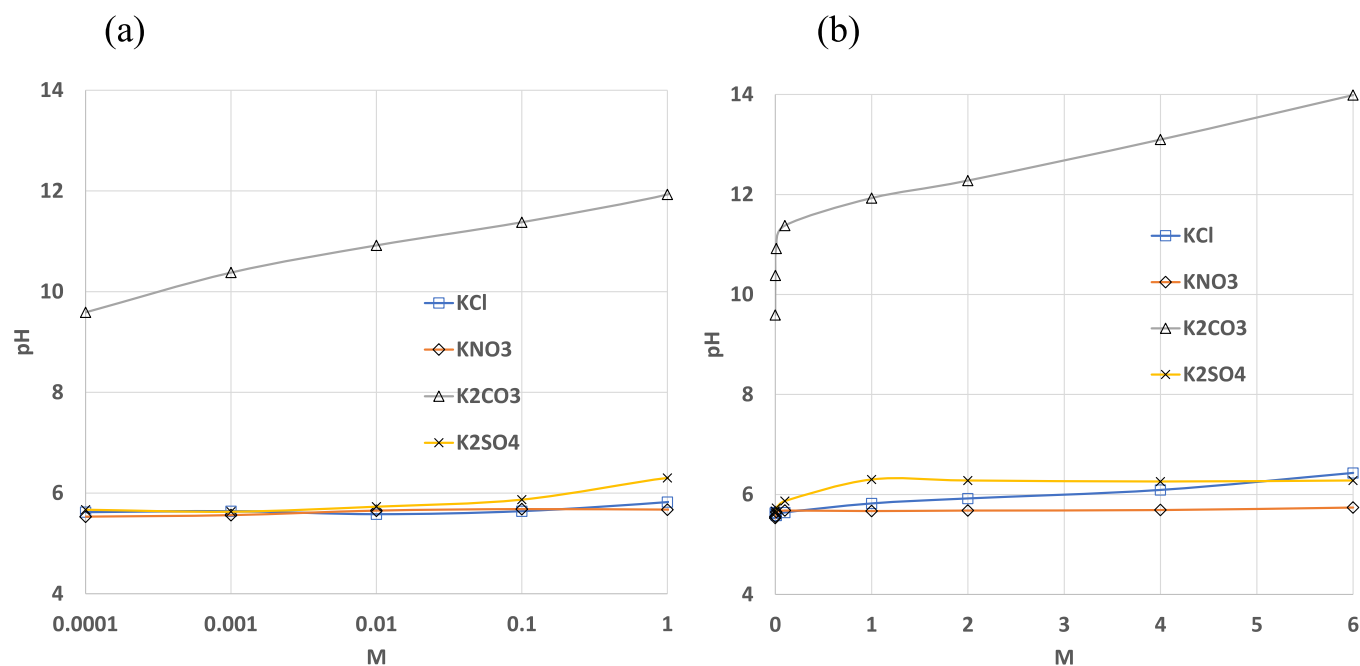


Fig. 2. pH of potassium salt solutions against their concentration at room temperature.

ter 4 M. KNO₃ had the minimum pH increment among the four salts studied.

K₂SO₄ had the second highest pH values among the four salts. Its pH was 5.67 at a salt concentration of 0.0001 M which was 0.05 and 0.14 higher than those of KCl and KNO₃ at the same concentration respectively. From 0.001 M to 1 M the pH values of K₂SO₄ increased linearly according to $\text{pH} = 4.7778 (\text{concentration of K}_2\text{SO}_4 \text{ in M}) + 5.8222$ ($R^2 = 1$). K₂SO₄ reached its highest pH value of 6.3 at 1 M. Then its pH value decreased to 6.27 ± 0.01 when its concentration was increased further. Saturation points for KCl, KNO₃, K₂CO₃ and K₂SO₄ are 3.2, 2.5, 3.4 and 0.6 M respectively. The pH values of K₂SO₄ stopped increasing from 1 M after reaching the saturation point at 0.6 M. This showed that precipitation of salt from the solution would not contribute to the total charge. However, the pH values of KCl, KNO₃ and K₂CO₃ were still increasing even after salt concentrations reached saturation. The following argument could also be included as the reason for the increase in pH even after the saturation point is reached in some solutions. Common salts can conduct electricity in aqueous solutions, but the dry salt will not allow the charges to flow. Shu et al. (2020) have shown that salt particles can exist in water in two forms. One form as salt particles alone and another as salt particles with water layers surrounding them. In this study, the pH value of K₂SO₄ becomes constant after the concentration of the solution reaches around 1 M. This indicates that the salt particles of K₂SO₄ exist in the solution as particles alone and which could not conduct electricity even they present in microscales. The increase in pH values of KCl, KNO₃ and K₂CO₃ solution after saturation can be due water layers surrounding those three kinds of salt particles and enabling the current pass through.

Will impurities affect the results of the pH values measured? Impurities could come from DI water, gases in the air, especially CO₂ as well as from the analytical grade salts. The conductivity of DI water is 1.10 $\mu\text{S}/\text{cm}$. This conductivity value includes the effect of impurities in the pure water and from air including CO₂. At 0.01 M, the conductivity of KCl is 1408.23 $\mu\text{S}/\text{cm}$ (Haynes, 2016). The conductivity of DI water is less than 0.1% of 0.01 M of KCl. The pH of KCl solution was 5.62 at the salt concentration of 0.0001 M and 5.64 at 0.1 M. Further, the difference of pH from 0.0001 M to 0.1 M is 0.02 unit, thus the effect of impurities on pH could not be significant. Will impurities cause the increase of pH?

Impurities in salts might be one of the reasons for pH increment after saturation for KCl, KNO₃ and K₂CO₃. An analytical grade salt will contain more than 99% of original salt and less than 1% of impurities. The impurities will be composed of various soluble and insoluble chemicals. Thus, it will be difficult to quantify the effect of the impurities in salt on pH since the information on the type and quantity of impurities and their pH after mixing are not available. One way to answer the question is to use a salt with 100% purity, which is not done in this study. But it will be safe to say that the impact of impurities will be insignificant as discussed in the previous section. The contradicting phenomena of pH increase after saturation need further investigation. Devising experiments to understand the interaction between the liquid and solid-state salts will be beneficial. The pH of K₂SO₄ was overtaken by that of KCl at around 5 M. The pH range of K₂SO₄ was from 5.67 to 6.3. The difference of its peak pH to pure water was 0.6 units. Like K₂CO₃, K₂SO₄ also possessed two negative charges but CO₃²⁻ was much more negative than those of SO₄²⁻ assuming K⁺ contributed positive charges to both potassium salts in a similar manner. Further in-depth research is needed to understand the vast difference in pH between K₂CO₃ and K₂SO₄.

At lower salt concentrations, some pH values of solutions were lower than that of pure water's. For example, the pH values of KCl were less than 5.64 units when its concentration changed from 0.0001 to 0.1 M; pH values of KNO₃ were less than 5.65 when its concentration changed from 0.0001 to 0.01 M and the pH values of K₂SO₄ were less than 5.67 when its concentration changed from 0.0001 to 0.001 M. Four kinds of salts used in this study were basic salts. Why did pH become acidic when basic salts were added in pure water? Pure water used during the study was 5.7, which is acidic. When salts dissolve in water, water clusters shrink (Shu et al., 2020). When the size of water clusters became smaller, the surface charge density became larger so water was more acidic resulting in the solution pH being less than that of initial water pH. The lower pH of salt solutions at low concentration was due to the size decrease of water clusters.

Relationship between hydrogen ions and protons/positive charges. In ideal solutions, there is neither chemical nor physical interactions between cations and anions so the concentration of hydrogen ions in a solution is the same as the concentration of protons since every hydrogen ion (H⁺) possesses a positive charge (e⁺). An equation could be

Table 1
pH and electrical properties of solutions at 25 °C.

pH	E_{cell} , mV	v	Potential energy, vFE_{cell} , mJ/mol
16	-532.44	-100	5.14×10^9
15	-473.28	-10	4.57×10^8
14	-414.12	-1	4.00×10^7
13	-354.96	-0.1	3.42×10^6
12	-295.80	-0.01	2.58×10^5
11	-236.64	-0.001	2.28×10^4
10	-177.48	10^{-4}	1.71×10^3
9	-118.32	10^{-5}	114.16
8	-59.16	10^{-6}	5.71
7	0	10^{-7}	0
6	59.16	10^{-6}	5.71
5	118.32	10^{-5}	114.16
4	177.48	10^{-4}	1.71×10^3
3	236.64	+0.001	2.28×10^4
2	295.80	+0.01	2.58×10^5
1	354.96	+0.1	3.42×10^6
0	414.12	+1	4.00×10^7
-1	473.28	+10	4.57×10^8
-2	532.44	+100	5.14×10^9

Electrostatic/cell potential (E_{cell} , mV) and charge activity (v) were calculated using equations:

$$E_{\text{cell}}, \text{mV} = -59.12 \text{pH} + 414.12; v = 10^{-\text{pH}} \text{ for } \text{pH} < 7; \\ -v = 10^{\text{pH}-14} \text{ for } \text{pH} > 7.$$

developed from the argument, $[\text{H}^+] = [\text{e}^+]$, where $[\text{H}^+]$ is concentration of hydrogen ions in mole per liter (mol/L); $[\text{e}^+]$ is proton concentration in mol/L. However, the measured pH is related to hydrogen ion activity a_{H^+} . Sørensen (Haynes, 2016) defined $\text{pH} = -\log[a_{\text{H}^+}]$ and $a_{\text{H}^+} = c_{\text{H}^+}/c^0$, where c_{H^+} is the active hydrogen-ion concentration in mol/L and c^0 is equal to 1 mol/L is the standard amount concentration. Similarly, we can have charge activity $v = c_{\text{e}^+}/c^0$, where c_{e^+} is the proton concentration in mol/L. Therefore, $a_{\text{H}^+} = v$. This equation relates the hydrogen ions activity with the charge activity. Both a_{H^+} and v are dimensionless. From $\text{pH} = -\log[a_{\text{H}^+}]$ and $a_{\text{H}^+} = v$, now new equations can be derived as follows:

$$\text{pH} = -\log[v] \quad (4)$$

Eq. (4) also could be written as

$$v = 10^{-\text{pH}} \quad (5)$$

$$-v = 10^{-(14-\text{pH})} \quad (6)$$

Eq. (5) was for positive charges with pH less than 7 and Eq. (6) for negative charges at pH more than 7.

Table 1 displays the relationship of pH with electrostatic/cell potential, E_{cell} (mV), charge activity and potential energy, vFE_{cell} . pH values from -2 to 16 were listed although pH scale 0 to 14 was the most used. At pH 7 and 25 °C, its electrostatic/cell potential E_{cell} (mV) equals 0. As pH value decreased by 1 unit from pH 7 its potential E_{cell} increased by 59.16 mV whilst its pH value increased by 1 unit from pH 7 its potential decreased by 59.16 mV. At pH 7, the charge activity was 10^{-7} mol/L. The charge activity increased 10-fold as pH value decreased 1 unit. The electron activity increased 10-fold when pH value increased 1 unit. An increase of 1 unit in pH from pH 7 had the same value as a decrease of 1 unit from pH 7 but it had a positive sign in front of the charge values if pH decreased from 7 and a negative sign if the pH increased from pH 7. The positive sign showed that the charges were positive, or protons and the negative sign showed that the charges were negative or electrons. At pH 7, the potential energy is 0. There will be a more than 10-fold increase in potential energy with every unit of either a decrease or an increase of pH from 7. At pH 6 and 8, the potential energy was 5.71

Table 2.1
Positive charges using equation $v = 10^{-\text{pH}}$.

Concentration	KCl	KNO ₃	K ₂ CO ₃	K ₂ SO ₄
0.0001	2.40×10^{-6}	2.95×10^{-6}	2.57×10^{-10}	2.14×10^{-6}
0.001	2.29×10^{-6}	2.75×10^{-6}	4.17×10^{-11}	2.34×10^{-6}
0.01	2.63×10^{-6}	2.24×10^{-6}	1.20×10^{-11}	1.86×10^{-6}
0.1	2.29×10^{-6}	2.09×10^{-6}	4.17×10^{-12}	1.35×10^{-6}
1	1.51×10^{-6}	2.14×10^{-6}	1.17×10^{-12}	5.01×10^{-7}
2	1.20×10^{-6}	2.09×10^{-6}	5.25×10^{-13}	5.25×10^{-7}
4	8.13×10^{-7}	2.04×10^{-6}	7.94×10^{-14}	5.50×10^{-7}
6	3.72×10^{-7}	1.82×10^{-6}	1.02×10^{-14}	5.25×10^{-7}

Table 2.2
Negative charges using equation $-v = 10^{(\text{pH}-14)}$.

Concentration mol/L	KCl	KNO ₃	K ₂ CO ₃	K ₂ SO ₄
0.0001	-4.17×10^{-9}	-3.39×10^{-9}	-3.89×10^{-5}	-4.68×10^{-9}
0.001	-4.37×10^{-9}	-3.63×10^{-9}	-2.40×10^{-4}	-4.27×10^{-9}
0.01	-3.80×10^{-9}	-4.47×10^{-9}	-8.32×10^{-4}	-5.37×10^{-9}
0.1	-4.37×10^{-9}	-4.79×10^{-9}	-2.40×10^{-3}	-7.41×10^{-9}
1	-6.61×10^{-9}	-4.68×10^{-9}	-8.51×10^{-3}	-2.00×10^{-8}
2	-8.32×10^{-9}	-4.79×10^{-9}	-1.91×10^{-2}	-1.91×10^{-8}
4	-1.23×10^{-8}	-4.90×10^{-9}	-1.26×10^{-1}	-1.91×10^{-8}
6	-2.69×10^{-8}	-5.50×10^{-9}	-9.77×10^{-1}	-1.91×10^{-8}

Table 3
Electrostatic/cell potential (mV), E_{cell} , of potassium salts at room temperature.

Concentration mol/L	KCl	KNO ₃	K ₂ CO ₃	K ₂ SO ₄
0.0001	81.62	86.95	-153.25	78.66
0.001	80.44	85.17	-199.99	81.03
0.01	83.99	79.85	-231.94	75.12
0.1	80.44	78.07	-259.16	66.83
1	69.79	78.66	-291.70	41.39
2	63.87	78.07	-312.40	42.58
4	53.82	77.48	-360.92	43.76
6	33.70	74.52	-413.57	42.58

Electrostatic/cell potential were computed from E_{cell} , mV = -59.12pH + 414.12.

mJ/mol and at pH 0 and 14 it was 4.00×10^7 mJ/mol. pH values in Fig. 2 were converted to either positive charge/proton activity or negative charge/electron activity and the results were listed in Tables 2.1 and 2.2. Equation $v = 10^{-\text{pH}}$ was used to calculate proton activity in solutions for pH less than 7 while electron activity was obtained using $-v = 10^{(\text{pH}-14)}$ for solutions with pH greater than 7. Comparing the two tables, the electron activity for K₂CO₃ in Table 2.1 was discarded and its value in Table 2.2 was chosen. The values of proton activity for the other three salts were adopted from Table 2.1.

Table 3 lists the electrostatic/cell potential, E_{cell} , of potassium salt solutions with their concentrations. For K₂CO₃ solutions, their E_{cell} values were negative. Their absolute values increased with the increase in concentration. It was -153.25 mV at 0.0001 M and -413.57 mV at 6 M. For the other three potassium salts, their E_{cell} were positive and the values decreased with concentration. This means that E_{cell} values for the three salts became less positive, and the solutions became more negatively charged. The maximum value of E_{cell} among the three kinds of potassium salts was 86.95 mV. This occurred for KNO₃ at 0.0001 M. The minimum value was 33.70 mV. It occurred for KCl at 6 M.

The product of charge activity and electrostatic/cell potential gives the potential energy of a solution. Table 4 lists the potential energy (vFE_{cell}) of potassium salt solutions. From the Table, it could be seen that the potential energy of K₂CO₃ solutions have positive values since both charge activity and electrostatic potential were negative values leading to the product to be positive. The positive values of potential energy

Table 4Potential energy (mJ/mol), vFE_{cell} , of potassium salts at room temperature.

Concentration mol/L	KCl	KNO ₃	K ₂ CO ₃	K ₂ SO ₄
0.0001	18.89	24.76	5.75×10^2	16.23
0.001	17.78	22.63	4.63×10^3	18.33
0.01	21.32	17.25	1.86×10^4	13.50
0.1	17.78	15.74	6.00×10^4	8.70
1	10.19	16.23	2.40×10^5	2.00
2	7.41	15.74	5.74×10^5	2.16
4	4.22	15.26	4.38×10^6	2.32
6	1.21	13.08	3.90×10^7	2.16

Electrostatic/cell potential (E_{cell} , mV) and pH values were calculated using equation:
$$E, \text{ mV} = -59.12 \text{ pH} + 414.12; v = 10^{-\text{pH}} \text{ for } \text{pH} < 7; -v = 10^{\text{pH}-14} \text{ for } \text{pH} > 7$$

might show the tendency for salt to dissolve in water since the potential energy is related to Gibbs energy. The higher the potential energy the higher the tendency of salt to dissolve in water. The potential energy of K₂CO₃ solutions increased significantly with the increase in concentration. The potential energy was 5.75×10^2 mJ/mol at 0.0001 M and 3.90×10^7 mJ/mol at 6 M. We generally observe that salt precipitates after saturation so that we perceive that the higher the concentration the lower the tendency of dissolution. This might be a superficial observation. The dissolution tendency might be affected by a combination of factors including charge activity, salt aggregation (Shu et al., 2015, 2005) and precipitation. Potential energy for the other three salts increased with the increase in concentration. KNO₃ solution at 0.0001 M might have the highest tendency to dissolve in water while KCl at 6 M had the least tendency to dissolve in water among the three salts.

Conclusions

In this study, pH of 4 common potassium salts were measured. The results demonstrate that every salt had its unique pH values and trends with the increase in their concentrations in the solution. Potassium carbonate with two negative charges had the highest pH increase with concentration. Its pH was more than 11 at 0.1 M and was a strong basic salt. The pH values of potassium sulfate were much lower than those of K₂CO₃ although K₂SO₄ also has two negative charges. The pH values of K₂SO₄ were slightly higher than those of KCl and KNO₃ before its saturation point (0.6 M) and they became stable from 1 M. pH values of KCl increased linearly and it overtook that of K₂SO₄ at around 5 M. pH values of KNO₃ were the lowest among the four kinds of potassium salts and had the least variability. Its pH value was 5.58 ± 0.01 from concentration 0.1 to 4 M, which was very close to the pure water pH (5.7 unit) used for making the salt solutions. The pH of KNO₃ increased from 5.53 units at concentration of 0.0001 M to 5.74 unit at 6 M. The experimental outcome showed that the contribution of positive ions and negative ions in a salt to the charge activity were not equal. In this study negative ions contributed more to the electrical nature of the salt particles than positive ions, so the pH of all the salts increased with concentration. pH is one of electrical properties of solutions. It related to charge activity of a solution and the system's electrostatic/cell potential can be computed with the well-established equation. Potential energy can be obtained from charge activity and electrostatic potential. This study highlights the diversity of salt properties as well as the importance of pH values and will lead to further research on thermodynamic properties of solutions.

Declaration of Competing Interest

None.

CRedit authorship contribution statement

Li Shu: Conceptualization, Methodology, Validation, Investigation, Resources, Writing – original draft, Writing – review & editing, Supervision, Project administration, Funding acquisition. **Xi Lu:** Investigation. **Veeriah Jegatheesan:** Resources, Writing – review & editing, Funding acquisition. **Leonardo Jegatheesan:** Writing – review & editing.

Data availability

Data will be made available on request.

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