A Highly Efficient Hydrogen Generation Electrolysis System using Alkaline Zinc Hydroxide Solution

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Abstract

Alkaline water electrolysis is a well-established conventional technique for hydrogen production. However, due to its relatively high energy consumption, the cost of hydrogen produced by this technique is still high. Here in this work, we report for the first time the application of alkaline zinc hydroxide solution (composed of sodium zincate and potassium zincate in NaOH and KOH solutions, respectively) as an efficient, simple and recursive electrolyte for producing clean hydrogen through a continuous dual-step electrolysis process. The ionic conductivity, electrodes current density, and hydrogen evolution rate were measured in a wide range of the electrolyte concentrations (0.1-0.59 M). Also, the cell efficiency was studied at different ranges of current density (0.09-0.25 A/cm²) and applied potential (1.8-2.2 V). Results indicated that the application of alkaline zinc hydroxide solution at the optimum electrolyte concentration can enhance the hydrogen evolution rate minimally by a factor of 2.74 (using sodium zincate) and 1.47 (using potassium zincate) compared to the conventional alkaline water electrolyzers. The results of this study could be helpful to better understand the electrochemical behaviour of the alkaline water electrolyzers when sodium zincate and potassium zincate are used as ionic activators for enhancing hydrogen evolution.

Keywords

Hydrogen, Water Electrolysis, Alkaline Electrolyte, Sodium Zincate, Potassium Zincate, Ionic Activator

1. Introduction

Hydrogen as a non-toxic, renewable, transportable and emission-free energy carrier is becoming a popular alternative fuel in the energy sector [1, 2]. It is the simplest, lightest, and the most abundant element in the universe with the highest specific energy content compared to all other conventional fuels [3, 4]. Unlike fossil fuels, hydrogen is not available freely in nature as a primary fuel meaning it must be first produced from its sources such as natural gas or water, and then used as an energy source [5, 6]. Currently, reforming of fossil fuels (i.e. natural gas) is the most common process used for commercial production of hydrogen [7]. However, this process is arguably less efficient due to the highly endothermic reactions linked with steam- or dry-reforming of hydrocarbons [8]. In addition, the use of fossil fuels in reforming processes is associated with emission of carbon-dioxide (CO_2), which is the major cause of the environmental phenomena so-called “greenhouse effect” [9, 10]. These issues have caused a great interest in developing alternative hydrogen production processes based on water splitting which are environmentally-friendlier and more sustainable, such as thermolysis processes, thermochemical processes, electrochemical processes, photochemical processes, photocatalytic processes, and photoelectrochemical processes [11-14].

Considering the current environmental and energy issues, water electrolysis can be an efficient, clean, and promising alternative technology for hydrogen production [15]. This technique could be well integrated with the domestically available renewable sources (such as solar, wind, or hydropower), giving it the advantage of sustainability. In addition, water electrolysis is capable of producing emission-free hydrogen gas with a purity reaching up to 99.999 vol. %, when its moisture content and oxygen impurity are eliminated [16-18]. Although water electrolysis has not yet become economical for large-scale hydrogen production, it is conceptually feasible for small-scale systems or remote services such as home-appliance, food industries, medical applications, spacecraft, and etc., especially where the renewable resources are abundantly available [19-21]. The major limitation of large-scale water electrolysis for mass production of hydrogen is the rather high over potential required for the reactions [22]. This over potential plus the ohmic losses in the electrolysis process makes the actual potential to exceed the standard 1.23 V potential of water electrolysis and reach values of around 1.8 to 2.0 V [23, 24]. Water electrolysis could be carried out in different

types of electrolytic systems such as alkaline electrolyzers, polymer electrolyte membrane (PEM) electrolyzers, and the solid-oxide electrolyzers (SOEs) [25-27]. However, the application of expensive material (such as precious metals) in fabrication of PEM or SOE electrolyzers along with their shorter operational lifetime make them relatively more expensive compared to the alkaline cells for small scale hydrogen production systems [18, 28].

Alkaline water electrolysis is one of the oldest and easiest techniques for producing hydrogen [29-31]. The current commercial alkaline water electrolyzers mainly use aqueous solutions of sodium or potassium hydroxide as the state-of-the-art electrolyte [32]. Although alkaline water electrolysis exhibits an inherent low-cost characteristic, due to the simplicity of the process, the development of large-scale hydrogen production systems based on this technology still needs potential improvements in the overall cell efficiency through reducing the energy consumption and increasing the electrolyte ionic conductivity [17, 33, 34]. Compared to the research on development of new electrolytes, the majority of the recent studies have concerned the improvement of alkaline electrolysis processes by focusing on the use of advanced electrode materials to increase the overall cell efficiency [17, 35].

Adding ionic activators is an efficient method for improving the properties of alkaline electrolyte solutions in water electrolysis [17]. The ionic activators can improve the cell energy consumption by electrodepositing an in-situ metal composite on the surface of the cathode, exhibiting a better electrode catalytic activity for hydrogen evolution reaction. In addition, ionic activators can enhance the electrolyte ionic conductivity and improve the electrode corrosion protection [36, 37]. The application of some ionic activators such dialkylimidazolium [38], Co-W [39], Mo-Pt [36], Ni-Co-Mo [40], and Ni-Cu-Mo [41] have been already reported in literature with obtaining an improved cell energy consumption.

Here in this paper, we report for the first time the use of zinc oxide (ZnO) as a simple an efficient ionic activator precursor for improving the hydrogen evolution rate for the alkaline water electrolysis systems using the conventional electrolyte solutions (NaOH and KOH). Adding ZnO powder to the aqueous NaOH and KOH solutions results in formation of the equilibrium alkaline zinc hydroxide

solutions (sodium zincate, $\text{Na}_2\text{Zn}(\text{OH})_4$ and potassium zincate, $\text{K}_2\text{Zn}(\text{OH})_4$, respectively) that could exhibit a better electrolyte ionic conductivity, increased hydrogen evolution rate, and less overall energy consumption during the water electrolysis process.

The aim of this paper is to investigate the use of alkaline zinc hydroxide solution as an efficient, simple, and recursive electrolyte for producing a clean hydrogen stream through a continuous water electrolysis system. Zinc oxide (ZnO) nanopowder was used as a precursor to mix with aqueous solutions of sodium hydroxide (NaOH) and potassium hydroxide (KOH) for preparing the alkaline zinc hydroxide solutions (i. e. sodium zincate and potassium zincate in NaOH and KOH , respectively as the electrolyte ionic activators). A handmade electrolyser using cylindrical graphite/zinc electrodes was used for analysing the influence of sodium zincate and potassium zincate concentrations on the electrolyte ionic conductivity, current density, hydrogen evolution rate, and cell efficiency were investigated in detail. The results of this study could help to understand better the ionic activation behaviour of sodium zincate and potassium zincate in the conventional alkaline water electrolyzers for enhancing the hydrogen generation rate.

2. Materials and methods

2.1 Apparatus

Batch experiments were carried out in a closed electrolysis cell, shown in Figure 1. The cell comprised of an open-top rectangular acrylic (Perspex) container, which was 25 cm high with a length and width of 20 cm x 12 cm. This ensured enough capacity to contain different volumes of solution for all the experiments. Cylindrical graphite electrodes (1 cm in diameter and 3 cm in height) and zinc electrodes (0.9 cm in diameter and 2.4 cm in height) were utilised separately. Each electrode was attached to the base of the container and were spaced apart by a 6.0 cm gap (centre-to-centre). A one litre gas collecting tube was placed over each electrode to capture hydrogen gas produced in the experiments. A scale on the side of the tubes allowed the volume of hydrogen gas produced to be measured. Nylon tubing was connected to the top of the tubes, allowing them to be filled with an electrolyte using a vacuum pump prior to the start of an experiment. Valves could then

be closed ensuring that any gas collected during the experiment would remain in the tubes. A conventional DC power supply (DIGIMESS HY3010, 0-30V/0-10A) was used to apply the required voltage/current to the system.

2.2 Experimental procedure

Sodium zincate solution (0.59 mole/litre) was prepared by first dissolving 660 g sodium hydroxide pellets (certified grade Sigma-Aldrich, 06203, 98% purity) in 1386 ml of distilled water. This would cause an exothermic reaction which would raise the temperature of the solution in which 66 g zinc oxide powder (certified grade Honeywell, 205532, 99.9% purity) was dissolved. The reaction between zinc oxide and sodium hydroxide to produce sodium zincate could be written as [42]:

In the above reaction, the aqueous NaOH solution was used in excess in order to dissolve the produced sodium zincate. The resulting solution was transferred into the electrolysis cell as the electrolyte solution after diluting with distilled water to achieve the desired sodium zincate concentrations. According to the literature, this solution may contain the equilibrium amount of zinc hydroxide, $\text{Zn}(\text{OH})_{2(\text{aq})}$ along with the formation of a range of cations and the other zinc hydrated oxoanion with different charges such as $\text{Zn}^{2+}_{(\text{aq})}$, $\text{Zn}(\text{OH})^{+}_{(\text{aq})}$, $\text{Zn}(\text{OH})^{3-}_{(\text{aq})}$, and $\text{Zn}(\text{OH})^{2-}_4$ [42, 43]. Sodium zincate could also form an equilibrium colloidal solution of zinc hydroxide and sodium hydroxide as [44]:

During the electrolysis, zinc hydroxide could be reduced for electrodeposition of a zinc layer over the surface of the cathode along with a simultaneous water reduction as follows [45]:

Similarly, the anode reaction could be written as [45]:

where, the overall reaction could be summarized as:

For the above set of reactions, the reversible (standard) cell potential, E°_{Cell} , which is the minimum cell voltage defined as the equilibrium potential difference between the standard anode and cathode potentials ($E^\circ_{\text{Anode}} - E^\circ_{\text{Cathode}}$) would be -0.848 V. Different values of the electrical potential (1.8, 2.0, and 2.2 V) were applied in the electrolysis experiments of this study. The cell did not show a proper rate of hydrogen generation below 1.8V that is reasonable considering the obtained minimum cell voltage.

In order to stop continuous deposition of zinc over the cathode, the electrolysis experiments were carried out within two steps. In the primary step, a uniform layer of zinc was allowed to be electrodeposited over the cathode surface during a fixed time (4 h), in accordance to Equation 3. In the secondary step, the electrodes were switched by swapping the power supply ports (+/- poles). During the secondary step, there is enough amount of zinc on the anode surface (previously used as the cathode in the primary step) to react with the oxygen generated by oxidizing the hydroxide ions (Eq. 5) to make zinc oxide (ZnO). The produced zinc oxide then could return back to the electrolyte solution by repeating the cycle through Equation 1. Accordingly, the overall reaction for the secondary step could be written as:

Therefore, the overall cell equation could be obtained by adding the summary equations of the primary step (Eq. 7) and the secondary step (Eq. 8) as the following:

In the electrolysis experiments, the solution was mixed regularly by inserting a glass rod from the open top of the tank in to the solution, to ensure uniform gradient of zinc hydroxide all over the cell. The readings were carried out in 5 minute intervals to monitor the hydrogen production over time. To maintain the electrolyte concentration to be constant during the measurement, since water is electrolytically dissociated to produce hydrogen, 1 ml of water was added to the electrolyte for every 1244 cm³ of hydrogen produced.

As an alternative electrolyte solution, potassium zincate (1.25 mole/litre) was prepared by first dissolving 1430 g of potassium hydroxide in 980 ml of distilled water followed by adding 98 g zinc oxide powder to the resulting solution. The related reactions for potassium zincate solution will be as described for sodium zincate, except that the sodium would have been replaced by potassium. Apart from these differences, the apparatus and method used were as described for sodium zincate solution. A conductivity measuring instrument (Mettler Toledo, FEP30-Kit) was used to measure the ionic conductivity of both electrolyte solutions within the whole range of concentration (0.1-0.59 M). The resistance of the cell was then calculated based on the electrodes distance (graphite) and surface area (6 cm and 10.2 cm², respectively).

2.3 Electrode switching

As mentioned above, in the secondary step the charge of electrodes have been changed by switching the power supply ports. During this step, as the electrolysis reaction progressed the zinc disposed on the anode could react with the oxygen evolving on the anode to produce zinc oxide. Accordingly, a reduction of the anode zinc layer was observed as the reaction progressed, Figure 2. Figure 2b(i) represents the anode at the start of a reaction with a thick layer of zinc just after the primary step. Figure 2b(ii) represents the anode status after few times of switching during the electrolysis, indicating the fact that the zinc layer has been substantially decreased. Finally, Figure 2b(iii) represents the anode surface at the end of the secondary step, showing that the layer of zinc has been completely depleted. Accordingly, oxygen bubbles could be evolved on the surface of the anode in accordance with Equation 5. It was noted that when the layer of zinc depletes, the observed current and the rate of hydrogen production decrease significantly (Figure 2a).

It should be noted that while the layer of zinc is being depleted on the anode, a further layer of zinc (with a half thickness in theory) is being deposited on the cathode, according to Equation 3. As a result, once the level of zinc on the anode was depleted, it was possible to switch the electrodes by swapping the power supply ports. The new anode (the former cathode) would then have a layer of

zinc disposed thereon. It is mentioned that there was no oxygen bubble evolving from the anode as long as a zinc layer was available on its surface.

There were several observable indicators to suggest the time of switching the electrodes, namely the drop in the current, the decrease in the hydrogen production rate and the increase in the oxygen production rate. In theory, these indicators are expected to appear at the same time. However, it was important to have one main indicator to ensure consistency. Accordingly, the level of the current was used as the main indicator to switch the electrodes. It was observed that once the layer of zinc became so thin that the surface of the anode was exposed to the solution, a rapid drop was observed in the current, indicating the point that the electrodes had to be switched. It should be mentioned that the switch time varied between experiments, due to changes in the applied voltage and the concentration of the solution. Accordingly, during the experiments it was possible to predict at what time the current was expected to drop, but also live observation was maintained on the current so that as soon as any drop in current was detected, the electrodes were switched. This approach ensured that a consistent standard was maintained.

3. Result and Discussion

3.1 Effect of sodium/potassium zincate concentration

Ionic transfer within the electrolyte solution depends on the concentration of the solution and distance between the electrodes. Here, the ionic resistance due to the distance of the electrodes was fixed and minimised by reducing the gap between the electrodes as much as the gas collecting tubes on top of the electrodes could be accommodated. The ionic resistance of the solution as a function of the sodium concentration was estimated by measuring the conductivity of the solution. The direct relationship between conductivity and resistance is given by the equation $R=L/\sigma A$ (Eq. 9), where R is the electrical resistance (Ω), σ is the electrical conductivity (mS/cm), L is the distance between the electrodes (cm) and A is the electrodes surface area (cm²). As it has been mentioned in the experimental procedure, the electrodes L and A were fixed (6 cm and 10.2 cm², respectively),

therefore the effect of changing the concentration of the sodium/potassium zincate solution on the resistance was estimated using the conductivity equation (Eq. 9) that has been illustrated in Figure 3.

The ionic transfer in the cell is controlled by convective mass transfer in the solution. At lower concentrations ($<0.25\text{M}$ for sodium zincate and $<0.3\text{M}$ for potassium zincate), the conductivity is low due to the solution becoming heavily diluted resulting in a decrease in the total number of dissociated ions. Conversely, at higher concentrations, the conductivity also decreases because of less mobility of the ions due to high viscosity of the solution and also the formation of neutral ion-pairs not contributing to the overall cell conductivity. Accordingly, the optimum concentration for the sodium and potassium zincate solutions were around 0.25M and 0.3M , respectively as shown in Figure 3. In addition, since sodium ions have a smaller size but larger charge density, compared to potassium ions, they illustrate greater interaction with water molecules at low concentrations (lower than 0.3M) [46]. However, for concentrations higher than 0.3M , the great increase in these interactions lower the mobility of sodium ions in the solution, resulting in a decrease in the conductivity. The same trend was observed for KOH, however due to the lower charge density of potassium ions, the conductivity peaks at higher concentrations. In addition, since at a given concentration, KOH solution shows lower viscosity than NaOH, an increase in the solution concentration at a fixed viscosity could raise its ionic conductivity, and therefore KOH solution at higher concentrations could show a higher ionic mobility than NaOH solution [47].

The ability to sustain a passage of electrical current by the electrolyte solution depends on the mobility of its constituent charged ions in the electric field between electrodes immersed in the electrolyte. Better ion mobility leads to higher reaction rates that in turn increases hydrogen production rate. Accordingly, hydrogen production rates at different concentrations were also analysed and the results have been presented in Figure 4. As can be observed, the higher ionic conductivity of NaOH electrolyte at low concentrations, results in greater H_2 production than that of KOH electrolyte. Although it was observed that at concentrations above 0.3M , KOH electrolyte shows higher ionic conductivity than that of NaOH electrolyte (Figure 3), a sudden decrease in the

ZnO solubility at concentrations higher than 0.3M decreases the concentration of $K_2Zn(OH)_4$ in the electrolyte, leading to a considerable fall in the H_2 production [42, 48].

3.2 Effect of voltage

The highest hydrogen production rate observed for sodium (potassium) zincate was at 0.2M (0.3M), and the trends for all three voltages suggests that the optimum concentration is between 0.1 M and 0.3 M (0.1M and 0.4M). Below this concentration, the hydrogen production decreases as the solution becomes heavily diluted and the availability of zinc hydroxide ions at the electrode and electrolyte interface is restricted. Similarly, above this concentration, neutral ion-pairs form and these will not be drawn to the electrodes.

The effect of the applied operating cell voltage on the hydrogen production was experimented by tuning the voltage for each experiment. The hydrogen production increases at higher voltages as it increases the electrolysis process. The relationship between the cell voltage and current characterises the electrochemical behaviour of an electrolysis cell. The hydrogen produced in electrolysis is proportional to the amount of charge involved in the process. Therefore, according to Faraday's law, the hydrogen production is directly proportional to the charge transfer, which is the electric current. Hence, when the voltage is tuned on the power supply, the electric charge delivered to the electrolysis process by the power supply affects the current. Therefore, by increasing the voltage, current density also increases resulting in higher hydrogen production rates as shown in Table 1.

In general, it is desirable to work at lower voltages to reach greater energy efficiencies for the cell. Since in the presented method a layer of zinc is constantly covering the graphite electrode, the cell can operate at lower voltages than would otherwise be possible. For instance, the rate of hydrogen production was high when a voltage of 1.8 V was used. However, as explained above, when the anode did not comprise a zinc layer, a minimum voltage of 1.9 V was necessary to simply allow the reaction to proceed at all. Similarly in industry, water electrolysis needs the minimum voltage of 2.0V for hydrogen generation for current densities between 0.1-0.3 A/cm² [17].

3.3 Effect of current density

As it is observed in table 1, hydrogen production rate is dependent on the current density in the electrolysis cell. Accordingly, the change in current density over time at different concentrations of sodium/potassium zincate was investigated. Figure 5 represents an example of the results obtained at the concentration of 0.2M.

As shown in this figure, the average current density at each voltage closely reflects the values reported in table 1. It was also observed that the current decreased over time as the zinc layer became thicker on the electrode. The formation of a very thick zinc layer on the electrode surface increases the resistance, which is of two components; the resistance of the electrode and the resistance of the electrolyte. The thick zinc layer on the electrode causes an increase in resistance at the electrode and electrolyte interface, which means the freshly formed zinc hinders the evolution of hydrogen or the availability of zinc hydroxide ions at the interface is restricted, suspending the electrochemical reaction. The cell responds by decreasing the current, resulting in lower rates of hydrogen production. Therefore, live observation on the current was applied so that as soon as any drop in current was detected, the electrodes were switched. The process of switching the electrodes can clearly be observed in Figure 5.

3.4 Cell efficiency

To compare the results of this work with the conventional electrolyser technologies, the energy efficiency of the current study was studied. For low-temperature alkaline water electrolysis, the efficiency of a water electrolysis systems could be evaluated by considering the hydrogen production rate and the total electrical energy applied to the cell, as given by the following equation [17]:

$$\eta_{H_2 \text{ Production rate}} = \frac{V_{H_2}}{U_{it}} \quad \text{Eq. 10}$$

where V_{H_2} is the hydrogen production rate per unit volume of the electrolysis cell, U is the cell voltage, i is the cell current, and t is time. The unit of η_{H_2} Production rate could be expressed as $m^3 m^{-3} h^{-1} kWh^{-1}$. In this unit, $m^3 h^{-1}$ represents the rate of hydrogen production, m^{-3} stands for the volume of the electrolysis cell containing the electrolyte, and the group $(kWh)^{-1}$ represents the amount of electrical power used during the electrolysis experiments. The cell electrolysis efficiency values obtained when the electrolyte comprised 0.2M sodium zincate (and 0.3M potassium zincate) have been represented in table 2.

According to the results, the cell is considered inefficient if a high voltage is used to produce the same hydrogen amount, while keeping the current constant. As shown in the Table 2, the obtained value for the system using sodium zincate (potassium zincate) solution at 2.2V is 6.3 (3.4) $m^3 m^{-3} h^{-1} kWh^{-1}$, that is almost 2.74 (1.47) times better than the value reported for the typical unipolar water electrolyzers ($2.3 m^3 m^{-3} h^{-1} kWh^{-1}$) [21, 49]. Accordingly, the electrolysis cell investigated by this study could provide much higher hydrogen production performance compared to conventional water electrolyzers if less power (voltages) is used to produce hydrogen.

3.5 Electrode material and surface condition

Selection of the correct electrode material is important for the efficient operation of the electrolysis cell. The electrode material was selected to be graphite based as this was understood to provide adequate strength and stability against physical attacks, such as erosion by the alkaline solution. At first, the electrodes had a smooth surface. However, when the electrodes were subjected to long hours of applied voltage, they corroded leaving a slightly porous surface. Repeating the initial experiments could determine the effect of this change. As shown in Figure 6, an increase in hydrogen production rate was observed when the electrodes had a porous surface, which was obtained after few times of using the electrodes. Quantitatively, an average increase of 96% in the hydrogen production rate was observed for the same voltage and electrolyte concentration as when performed on a smooth surface. This matter was due to the increase in the active surface area of the

electrodes after a slight corrosion taking place on their surfaces during the electrolysis experiments. This could clearly show the advantageous application of the porous electrodes (not applied in this study) for water electrolysis using aqueous sodium zincate and potassium zincate as the electrolyte solutions.

In order to study the effect of electrode material on the hydrogen production, the graphite electrode was replaced with zinc electrodes [2.4 cm (h) x 0.9 cm (d)]. All the experiments were repeated with the zinc electrode in the sodium zincate solution. The hydrogen production rates at different concentrations are shown in Figure 7. Again, the highest hydrogen production rate observed was at 0.2M, and the trends for all three voltages similarly suggested that the optimum concentration for maximizing the hydrogen production is between 0.1 M and 0.3 M.

In addition, it was found that in the case of using zinc electrodes, a better cell efficiency can be obtained if the cell operates at lower voltages (Table 3).

Similar to the graphite electrodes, the change of current density over time in a sodium zincate electrolyte was investigated using zinc electrodes, and the results are shown in Figure 8. Again, considering all the switching steps, the average current density at each voltage closely reflects the values reported in table 3.

The cell efficiency was also calculated for this system, and the values obtained are given in Table 4. Comparing the results obtained for zinc electrodes (table 4) with those of graphite electrodes (table 2), indicate a higher cell efficiency for the system using graphite electrodes. This matter could be explained by the presence of activated zinc particles already electrodeposited on the surface of graphite electrode, whereas the zinc available on the surface of metallic zinc electrode is not as active to react with hydroxide ions present on the surface of the anode [50].

4. Conclusion

The application of the alkaline zinc hydroxide solutions (including sodium zincate and potassium zincate) as ionic activator for improving the energy efficiency of the alkaline water electrolyzers have been studied. The optimum conditions for operation were investigated by analysing the

influences of applied voltage, current density and sodium/potassium zincate concentrations. The experimental results show that the hydrogen production peaks at about 0.2M for sodium zincate solution and 0.3M for potassium zincate solution. The measurement of the cell energy efficiency at the optimum zincate concentration demonstrated the fact that the use of alkaline zinc hydroxide solutions could increase the hydrogen evolution rate minimally by a factor of 2.74 (using sodium zincate) and 1.47 (using potassium zincate), compared to the to conventional alkaline water electrolyzers. Furthermore, it was found that the formation of a porous surface for the graphite electrodes positively affected the hydrogen production rate. In terms of the electrode material, similar trends were observed when the graphite electrodes were substituted with zinc electrodes.

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Tables

Table 1. Current density at different voltages for all concentrations.

Concentration (M)	Average current density (A/cm ²)					
	Na ₂ Zn(OH) ₄			K ₂ Zn(OH) ₄		
	2.2 V	2.0 V	1.8 V	2.2 V	2.0 V	1.8 V
0.59	0.11	0.09	0.09	0.41	0.30	0.28
0.55	0.13	0.11	0.11	0.40	0.32	0.31
0.5	0.15	0.15	0.15	0.39	0.33	0.32
0.45	0.19	0.19	0.17	0.37	0.33	0.31
0.4	0.19	0.16	0.14	0.36	0.32	0.30
0.3	0.22	0.18	0.14	0.32	0.32	0.30
0.2	0.25	0.21	0.18	0.29	0.26	0.25
0.1	0.18	0.18	0.11	0.15	0.14	0.14

Table 2. Hydrogen production rate at unit volume electrolysis cell against the total electrical energy applied to the cell at various voltages at 0.2M (0.3M) sodium zincate (potassium zincate) with graphite electrodes

Voltage (V)	Hydrogen production rate at unit volume electrolysis cell against the total electrical energy applied to the cell (m ³ m ⁻³ h ⁻¹ kWh ⁻¹)	
	Na ₂ Zn(OH) ₄	K ₂ Zn(OH) ₄
1.8	19.5	4.07
2	13.5	3.64
2.2	6.3	3.40

Table 3. Current density at different voltages shown for all concentrations (sodium zincate / zinc electrode)

Concentration (M)	Average current density (A/cm ²)		
	2.2 V	2.0 V	1.8 V
0.59	0.13	0.12	0.11
0.55	0.15	0.13	0.12
0.5	0.16	0.15	0.12
0.45	0.18	0.16	0.13
0.4	0.19	0.18	0.15
0.3	0.20	0.19	0.16
0.2	0.23	0.20	0.18
0.1	0.18	0.16	0.13

Table 4. Hydrogen production rate at unit volume electrolysis cell against the total electrical energy applied to the cell at various voltages at 0.2M sodium zincate with zinc electrodes

Voltage (V)	Hydrogen production rate at unit volume electrolysis cell against the total electrical energy applied to the cell (m ³ m ⁻³ h ⁻¹ kWh ⁻¹)
1.8	13.8
2	9.3
2.2	5.2

Figures

Figure 1. a) Schematic diagram of the experimental set-up for an electrolysis cell, b) photo of the experimental set-up

Figure 2. a) Change in hydrogen production rate and current over time, b) schematic diagram showing how the zinc layer on the anode changes with time.

Figure 3. Resistance and conductivity of the electrolyte as a function of concentration at room temperature.

Figure 4. Hydrogen production rate against sodium/potassium zincate concentrations for different voltages.

Figure 5. Current density against time for a) sodium zincate and b) potassium zincate electrolyte at 0.2M for different voltages.

Figure 6. Effect of electrode surface on hydrogen production (for sodium zincate solution).

Figure 7. Hydrogen production rate against sodium zincate concentrations vs voltage for zinc electrodes.

Figure 8. Current density vs time for sodium zincate electrolyte (0.59M) at 1.8 V, 2.0 V and 2.2 V for an electrolysis system using zinc electrodes.