

A Study of Molar Solubility of Sodium Chloride, Sodium Carbonate, Sodium Nitrate and Calcium Sulphate in Aqueous Media at Various Temperatures

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Abstract

Water being a universal solvent and very easy to get, most chemical reactions are carried out in aqueous media. However, there are other chemical reactions which take place in non-aqueous solvent media like benzene, chloroform and liquid ammonia. In view of this therefore, the molar solubilities per dm^3 in aqueous solutions of the following inorganic salts were studied at different temperatures. i.e. sodium chloride (NaCl), sodium carbonate (Na_2CO_3), sodium nitrate (NaNO_3) and calcium sulphate (CaSO_4). The solubilities per dm^3 of these salts were initially determined from their various saturated solutions. These values were converted to molar solubilities per dm^3 for accuracy and proper analysis: the impact of temperature on molar solubility of each salt was observed and recorded. While temperature increase has no effect on the molar solubility of sodium chloride, the molar solubility of calcium sulphate decreases with increase in temperature. There was a drastic increase in the molar solubility of sodium carbonate from 30°C to 50°C from 5.0mol dm^{-3} to 18.9mol dm^{-3} . The value then decrease gradually with increase in temperature and recorded a minimum value of 5.7mol dm^{-3} at 70°C . However, there was a gradual increase of the molar solubility of sodium nitrate from 30°C to 60°C i.e. from 8.2mol dm^{-3} to 12.0mol dm^{-3} and the value slightly dropped down to 11.4mol dm^{-3} at 70°C .

Keywords: Study, Solubilities, In-organic salts, Temperature.

INTRODUCTION

Solubility is the maximum amount of solute in grams or moles that will dissolve in one (1) dm^3 of a solvent at a particular temperature. (Butler and James, 1998) It is usually express in mol dm^{-3} or g/dm^3 .

Through solubility, one can determine the extent to which different solutes can dissolve in a solvent at a particular temperature. The solubility of solid substances is generally affected by the following factors i.e. temperature, nature of solute and solvent as well as the surface area of the solute. (Yalkowsky and Samuel, 2003). In determining the solubilities of substances, their saturated

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solutions in aqueous or non aqueous solvents were initially prepared. Depending upon the nature of the solute and the solvent, the solution could be unsaturated, saturated or super saturated one.

Unsaturated solution is a solution that can dissolve more solute at a particular temperature. A saturated solution is a solution that contains as much solute as it can dissolve at a particular temperature in the presence of un-dissolved solute. It is interesting to note that a saturated solution is not affected by the presence of excess un-dissolved solute. An Equilibrium is therefore established between the dissolved solutes of a saturated solution and the un-dissolved solutes. If the temperature of the solution is increased, there will be an increase in the solubility. In the same vein, if the temperature of the saturated solution is reduced by cooling, there will be a decrease in the solubility of the solute and the excess solute will crystallize out (Lin Yibin, 2008).

Super saturated solution is a solution which contains more solute than it can normally dissolve or more solute than is required to form a saturated solution at a particular temperature.

Chemists are always interested in counting the number of particles like atoms, molecules and ions in substances. It is practically impossible to count the numbers as they are incredibly small. Consequently, they (chemist) now resorted to using a unit called mole (Ababio, 2002).

Mole means the amount of substance. Virtually all the calculations in chemistry are related to mole. The short form for mole is mol. One mole of substance is the amount of that substance that contains as many elementary particles (atoms, molecule, ions, protons, electrons etc) as the number of atoms in exactly 12 grams of carbon -12. One mole of an element is equal to the relative atomic mass of that element. e.g. one mole of magnesium atom is 24. Similarly, one mole of a particle (atoms, molecules, ions etc) contains the Avogadro's number of particle represented as N_A or L with a value of 6.02×10^{23} . One mole of a gas occupies a volume of 22.4 dm^3 at standard temperature and pressure (STP) (Raymond and Ken, 2014).

When an ionic crystal dissolves in water, the ions become separated and energy is required to break up the lattice. In solution, cations and anions are solvated and energy is released in the process. i.e (Afanasev, 2011).

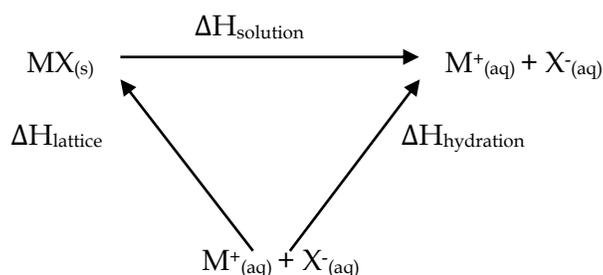


Fig. I: Relationship between enthalpy change of solution, hydration enthalpy and lattice enthalpy.

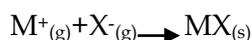
The energy change involved when one mole of a substance is dissolved completely in a solution is called the enthalpy change of solution ($\Delta H_{\text{solution}}$). Infact enthalpy change of solution can be used to indicate the relative solubility of ionic compounds. The enthalpy change of solution can be related to hydration enthalpy and lattice enthalpy i.e $\Delta H_{\text{solution}} = \Delta H_{\text{hydration}} - \Delta H_{\text{lattice}}$

Hydration enthalpy ($\Delta H_{\text{hydration}}$) is the amount of energy released when one mole of aqueous ions is formed from its gaseous ions. i.e.



Hydration enthalpy must have a negative value as it is the amount of energy released from attraction between ions and water molecules. The amount of energy released is therefore dependant on the charge/radius ratio of the ion. The greater the charge on the ion, the greater is the attraction between the ions and the water molecules, and the larger is the amount of energy given out and vice versa (Shoemaker *et al.*, 2009).

Lattice enthalpy ($\Delta H_{\text{lattice}}$) is the enthalpy change when one mole of an ionic crystal is formed from its constituents ions in the gaseous state. i.e.



The processes could be best explained by the above diagram Fig. I

If the magnitude of $\Delta H_{\text{hydration}}$ is greater than that of $\Delta H_{\text{lattice}}$ then $\Delta H_{\text{solution}}$ is negative (exothermic) and the compound will generally be soluble in water. If however, the value of $\Delta H_{\text{hydration}}$ is positive and small, the compound will still be soluble in water (Donald and Duane, 2001).

Group I metal ions are comparatively large in size and carry small charges, so, their effective nuclear charges are not very high and therefore their lattice enthalpies are usually more or less the same as their hydration enthalpies. Hence, their enthalpy changes of solution become negative or slightly positive in value (Donald and Duane, 2001).

Group II metal ions have higher charges and smaller sizes, so $\Delta H_{\text{hydration}}$ of their compounds are larger in magnitude than those of group I compounds. Consequently their $\Delta H_{\text{solution}}$ are less negative or more positive. Since the size of anion is a crucial factor affecting the solubility, the magnitude of hydration enthalpy is dependent on the charge /radius ratio of the cations. The larger the cations, the smaller the hydration enthalpy (Onuchukwu and Waziri, 1987). On the other hand, lattice enthalpy is dependent on the sum of the ionic radii i.e. $r^+ + r^-$. The greater the values of $r^+ + r^-$, the smaller will be lattice enthalpy (Wong *et al.*, 2002).

BACKGROUND OF THE RESEARCH

The choice of these salts for the purpose of this research work is based on their easy accessibility and safety. They were all purchased from local chemical vendor here in Azare. These Salts are safer in handling, cheap and highly qualitative. Their industrial and domestic applications in our daily life cannot be over emphasized. Specifically, sodium chloride is used in paper and plastic industries. Domestically, beside being used as food ingredient, it is also used in food preservation and water treatment, Calcium Sulphate is used in chemical industries as desiccant and coagulating agent (American Chemical Society, 2006). Domestically, Calcium Sulphate is used in the manufactured of dietary calcium supplements, which are often prescribed to people deficient in calcium. Calcium deficiency causes weakness of hair, bone and teeth brittleness (Nguyen, 2010). Sodium carbonate occurs as extensive beds of sodium minerals and as sodium rich water (Brines). In laboratory, the suspended sodium hydrogen carbonate (NaHCO_3) was removed from the

carbonating tower and heated to 300°C producing sodium carbonate as per the following equation (Whitten *et al.*, 2009)



Industrially, sodium carbonate is used in the manufacture of paper and glass. Domestically, Sodium bicarbonate is used as a water softening agent.

Sodium nitrate is synthesized industrially by reducing nitric acid with a base like sodium hydroxide i.e.



Sodium nitrate is used as anti-microbial agent in food industries (Barnum and Dennis, 2003). A fair knowledge of the solubility of these salts and their variation with temperature will be of great importance to chemists for their laboratory undertakings and research work. The same knowledge will benefit chemical vendors and chemical industries.

MATERIALS AND METHODS

The reagents used in this work include the followings; Sodium Chloride, Sodium Nitrate, Sodium Carbonate, Calcium Sulphate and Distilled Water.

The equipment and apparatus used are Weighing Balance, Evaporating dish, Bunsen burner, Beakers, Conical Flask, Funnel, Boiling Tube, Filter paper, Thermometer and Desiccator.

EXPERIMENTAL PROCEDURE

A clean dry boiling tube was half filled with distilled water and heated to 30°C. 0.5g of Sodium chloride crystals were added and stirred while the temperature was maintained at 30°C. Addition of sodium chloride continued until undissolved ones were seen at the bottom of the boiling tube. At this juncture, the sodium chloride solution was filtered with help of a filter paper and the filtrate is the unsaturated solution at 30°C. A clean dry empty evaporating dish was weighed and recorded as m_1 (45.50g). A good portion of the unsaturated solution was poured into the empty evaporating dish, reweighed and recorded as m_2 (61.10g). The saturated solution (m_2) was heated to dryness on water bath. When all the water has evaporated into the atmosphere, the residue was allowed to cool down in a dessicator for about 30 minutes after which, the content was weighed and recorded as m_3 (49.50g). In order to find the volume of solvent (Water) that has evaporated into the atmosphere, the mass m_3 (crystal residues) was subtracted from the mass of saturated solution m_2 . i.e. $m_2 - m_3$ (61.10- 49.50)g = 11.60cm³(V). In determining the mass of solute (NaCl) that dissolved in 11.60cm³ of water (which has evaporated into the atmosphere), the mass of empty evaporating dish (m_1) was subtracted from the mass of sodium chloride residue i.e. $m_3 - m_1$. (49.50- 45.50) = 4.00g. (m_4). To determine the solubility of the solute (NaCl) in gdm³ of the solution, the following formula was applied i.e.

$$\frac{\text{mass of solute}(m_4)}{\text{volume of solvent}(V\text{cm}^3)} \times 1000$$

$$= \frac{4.00g}{11.60cm^3} \times 1000.$$

$$= 345gdm^{-3}$$

The molar solubility of the same sodium chloride at this temperature (30°C) was calculated by dividing the solubility in gdm⁻³ with its (NaCl) molecular weight i.e.

$$\frac{345gdm^{-3}}{58.5}$$

$$= 5.90 \text{ moldm}^{-3}$$

The molar solubilities of sodium chloride at 40°C, 50°C, 60°C and 70°C were determined by repeating the above mentioned procedures. Similarly, the molar solubilities of calcium sulphate (CaSO₄), sodium carbonate (NaCO₃) and sodium nitrate (NaNO₃) at 30°C, 40°C, 50°C, 60°C and 70°C. were determined using the above prescribed method.

RESULTS AND DISCUSSION

Table 1 shows the variation of molar solubility in moldm⁻³ of sodium chloride with temperature in degree centigrade

Salt	Temperature (°C)	Molar solubility (moldm ⁻³)
Sodium Chloride	30	5.90
	40	5.85
	50	5.90
	60	5.90
	70	5.90

Table 2, shows the variation of molar solubility in moldm⁻³ of calcium sulphate with temperature in degree centigrade

Salt	Temperature (°C)	Molar solubility (moldm ⁻³)
Calcium Sulphate	30	7.40
	40	5.80
	50	4.00
	60	2.20
	70	0.50

Table 3. Shows the variation of molar solubility in moldm⁻³ of Sodium carbonate with temperature in degree centigrade

Salt	Temperature (°C)	Molar solubility (moldm ⁻³)
Sodium carbonate	30	5.00
	40	11.80
	50	18.90
	60	12.00
	70	5.70

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Table 4. Shows the variation of molar solubility in mol dm^{-3} Sodium nitrate with temperature in degree centigrade.

Salt	Temperature ($^{\circ}\text{C}$)	Molar solubility (mol dm^{-3})
Sodium Nitrate	30	8.20
	40	9.60
	50	10.80
	60	12.00
	70	11.40

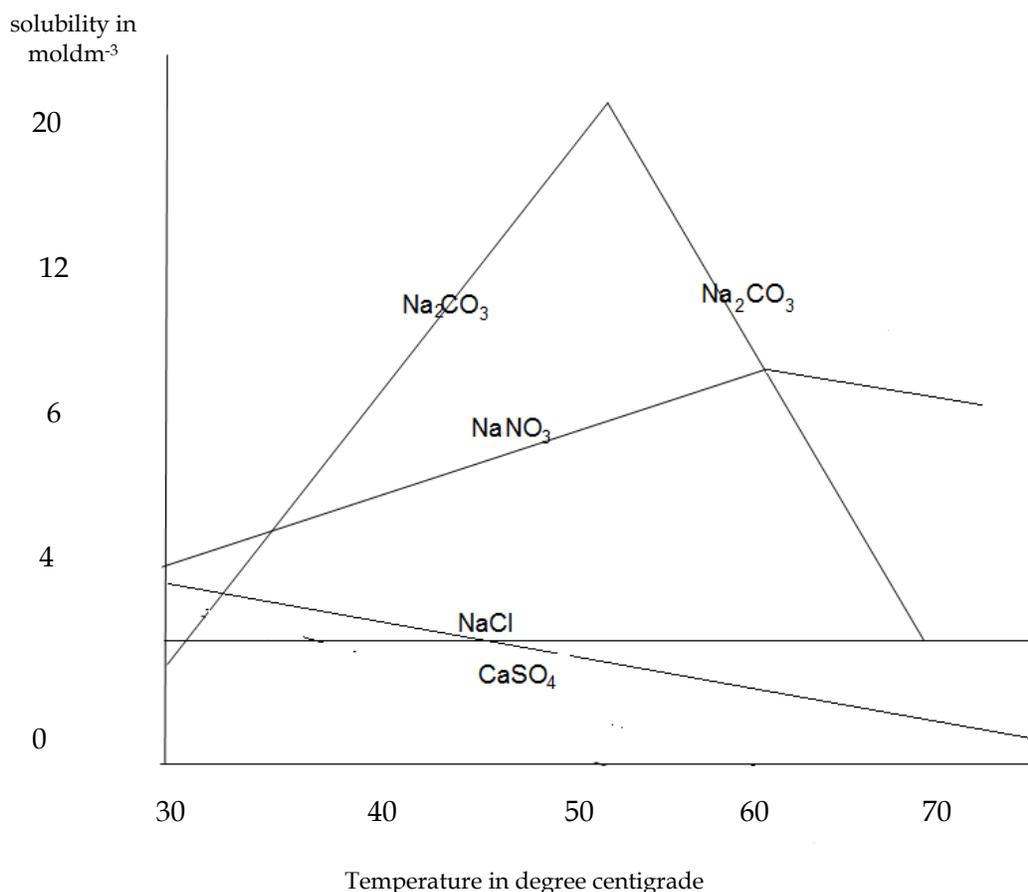


Fig. II: Variation of solubility in mol/dm^3 against temperature in Degree Centigrade

Discussion

When sodium chloride dissolves in water the ions are not hydrolyzed by water and the resulting solution is neutral. The reaction is endothermic i.e. heat is absorbed from the surrounding ΔH become +ve. Sodium chloride dissolving in water is a classic case of ionic salt. It is a coulombic interaction between a positively charged sodium ion and six water molecules as well as the corresponding interactions between a negatively charged chloride ion and six other water molecules. In the case of sodium ion, the positive charge attracts the negative end of each water molecule's dipole. The negative charge of the chloride ion attracts the positive end of the O-H

bond dipole. Both sodium and chloride ions were encased in a shell of water molecules. The shells insulate the ions from each other, allowing the oppositely charged particles to separate. The ions were not free to move and this affected their solubility. For the ions to move around, the strong charged dipole bond has to be broken down with high temperature above the boiling point of water. This is the reason why increase in temperature from 30° to 70°C has no effects on the molar solubility of sodium chloride as indicated in figure II. This finding agrees with Akinnawonu's work whereby increase in temperature has no effect on the solubility of Sodium Chloride (Akinnawonu, 2016)

In the case of calcium sulphate, its cation (Ca^{2+}) is smaller in size with the higher charges, so its $\Delta H_{\text{lattice}}$ has larger magnitude than, $\Delta H_{\text{hydration}}$. Consequently, its $\Delta H_{\text{solution}}$ is slightly positive and therefore less soluble in water. Increase in temperature from 30° to 70°C resulted in larger magnitude of $\Delta H_{\text{lattice}}$ than $\Delta H_{\text{hydration}}$. That is the more reason why the molar solubility of calcium sulphate decreases with increase in temperature as shown in figure II. This work is in conformity with that of Yuexia *et al.*, 2013 which also reported that the solubility of Calcium sulphate decreases with increase in temperature

In the case of sodium carbonate, there was a rapid increase in its molar solubility from 30° to 50°C i.e. from 5.00 mol dm^{-3} to 18.90 mol dm^{-3} . This was due to the fact that sodium ion is large in size with the smaller charge resulting into smaller magnitude of its $\Delta H_{\text{lattice}}$ than $\Delta H_{\text{hydration}}$. $\Delta H_{\text{solution}}$ therefore become more negative making its salt more soluble in water from 30°C to 50°C. Its molar solubility dropped from 18.90 mol dm^{-3} to a minimum value of 5.70 mol dm^{-3} at 70°C. This was partly due to the fact that increase in temperature increases the magnitude of $\Delta H_{\text{lattice}}$ more than $\Delta H_{\text{hydration}}$ making $\Delta H_{\text{solution}}$ less negative and therefore less soluble in water. Partly, was due to its low nuclear charge value which also increased the magnitude of $\Delta H_{\text{lattice}}$ making $\Delta H_{\text{solution}}$ less negative. This is similar to Seidell *etal* work (1919) where the solubility of Sodium carbonate increases with increase in temperature ie from 25°C to 100°C but dropped slightly from 48 to 43g/100ml of distilled water.

Considering sodium nitrate (NaNO_3), Sodium ion is larger in size with the smaller charge and therefore its hydration enthalpy ($\Delta H_{\text{hydration}}$) is larger than $\Delta H_{\text{lattice}}$. This why the molar solubility gradually increased from 30° to 60°C i.e. from 8.20 mol dm^{-3} to 12.00 mol dm^{-3} . However, the value slightly decrease from 12.00 mol dm^{-3} at 60°C to 11.40 mol dm^{-3} at 70°C. This was partly due to increased in $\Delta H_{\text{lattice}}$ with the increased in temperature and partly due to the larger size of its anion more than that of the cations. The larger the size of the anion the less soluble the salt is in water. This is similar to Akinnawonu's work in which the solubility increases from 20°C to 60°C ie from 3.0 mol dm^{-3} to 13 mol dm^{-3} at 60 °C and dropped to 2.5 mol dm^{-3} at 70°C (Akinnawonu, 2016).

CONCLUSION

Sodium chloride is a typical ionic compound and its solubility is not affected by increase in temperature. Calcium sulphate, a similar compound with the same type of bonding, its solubility in aqueous media decreases with the increase in temperature. Cations of group I metals are larger in size and carry small charges, so, their enthalpy changes of solution ($\Delta H_{\text{solution}}$) are usually negative and therefore more soluble in water compared to group II metals ions which have higher charges with the smaller sizes.

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