www.rjptonline.org

RJPT
Research Journal of Pharmacy and Technology

RESEARCH ARTICLE

Hydroquinone Solubility in Pure and Binary Solvent Mixture at Various Temperatures with FTIR

Chandrakant S. Aher

Department of Chemistry, M.S.G. College Malegaon Camp, Malegaon, Dist. Nashik, Pin 423105. India.
*Corresponding Author E-mail: **chandsaher1980@rediffmail.com**

ABSTRACT:

Gravimetric method is used to measured hydroquinone solubility in water, ethanol and in water+ethanol binary mixtures at temperatures (293.15, 295.15, 298.15, 300.15, 303.15, 305.15, 308.15, 310.15 and 313.15)K. Mole fractions solubility of hydroquinone are correlated with temperature by using the Apelblat equation. The combined nearly ideal binary solvent (NIBS)-Redlich-Kister equation is used to fit experimental hydroquinone solubility data in mixed solvents at constant temperature. ΔH^0_{soln}, ΔG^0_{soln}, and ΔS^0_{soln} are thermodynamic functions of hydroquinone in different solvents, obtained from the modified van't Hoff equation. FTIR study is done for some hydroquinone solution.

KEYWORDS: Hydroquinone, Solubility, Density, Apelblat equation and FTIR.

INTRODUCTION:

Dyes, paper, pesticides, polymeric material, pharmaceutical and petrochemical product etc are produced by using hydroquinone as the major benzene metabolite. It is used as a developing agent in photography, dye intermediate, stabilizer in paints, varnishes oils and motor fuels. In addition, hydroquinone has been used as an antioxidant in the rubber and food industry. From 1950 to 2001 hydroquinone is applied in the commercially available cosmetic skin lightening formulations in European Union countries and since 1960 hydroquinone is commercially available as a medical product, cosmetic formulations of products for coating finger nails and hair dyes[1,2].

The antimicrobial properties of arbutin as the main compound and hydroquinone as the active metabolite was determined and compared with the antimicrobial properties of *A. unedo leaf* extracts so as to test the extent to which arbutin is responsible for antimicrobial activity[3]. Wide used of such substances increase phenolic compounds in industrial wastewater, these are toxic to aquatic life and human bring[4].

Solubility data is required for selection of proper solvent and design an optimized crystallization process, in this paper the systematic study of solubility and densities of hydroquinone in water, ethanol and water + ethanol binary solvents over the entire composition range from zero to one mole fraction at temperatures (293.15 to 313.15) K is reported. The thermodynamic functions for saturated hydroquinone solution are calculated using modified van't Hoff equation.

MATERIALS AND METHOD:

Material:-

Triple distilled water is used in all experiments. Other chemicals is supplied by

Chemical Name	Supplier Name	Percentage purity	Standard
Hydroquinone	Sigma-Aldrich co.	99%	Reagent Grade
Ethanol	Merck, Darmstadt, Germany.	≥99.8%	G.R.

Selected solvent ethanol is very common and mainly used for many industrial processes. Temperature selected is closer to room temperature.

Apparatus and Procedure:-

Many methods are available to determined the solubility[5, 6]. In this work the solubility of hydroquinone is measured using an apparatus similar to that described

in the literature[7, 8]. In this work an excess amount of hydroquinone is added to the binary solvents mixtures prepared by weight (Shimadzu, Auxzzo Ltd.) with an uncertainty of ±0.1 mg, in a specially designed 100 ml double jacketed glass flask. Water is circulated at constant temperature in jacket between the outer and inner walls of the flask. The temperature of the circulating water is controlled by auto temperature control thermostat within (±0.1)K. The solution is continuously stirred using a magnetic stirrer to assured equilibrium and no further solute dissolved. The temperature of solution is same as that of circulating water. The stirring is stop and the solution is allowed to stand to get the supernatant liquid, which was withdrawn from the flask in a weighing bottle with the help of pipette which is hotter than the solution. The weight of this sample was taken and kept in an oven at 343 K until the whole solvent was evaporated and the residue was completely dry. This is confirmed by weighing two or three times until a constant weight was obtained after keeping the sample in an oven for another 30 min every time. The solubility has been calculated using weight of solute and weight of solution. Each experimental value of solubility is an average of at least three different measurements and the standard uncertainty of the experimental mole fraction solubility(x_B), value is ±

0.003. The mole fraction solubility(x_B), initial the mole fraction of ethanol (x_C^0), were calculated using usual equations[9]. The standard uncertainty for x_C^0 is 0.0002. Densities are determined using a 15 cm^3 bicapillary pycnometer as described earlier. For calibration of pycnometer triply distilled and degassed water with a density of 0.99705 $g \cdot cm^{-3}$ at 298.15 K was used. The pycnometer filled with air bubble free experimental liquids was kept in a transparent wall thermostat (maintained at constant temperature ± 0.1 K) for (10 to 15) min to attain thermal equilibrium. The heights of the liquid levels in the two arms were measured with the help of a travelling microscope, which could read to 0.01 *mm*. The estimated standard uncertainty of the density measurements of the solvent and binary mixtures was 10 $kg \cdot m^{-3}$. [10-13]

RESULTS AND DISCUSSION:
Solubility: Table 1 show the experimental and calculated (using Apelblat equation) values of solubility (x_B) of hydroquinone at 293.15 to 313.15 K in water, ethanol and water + ethanol respectively. The density of each saturated solution is also reported. Variation of solubility with x_C^0 is visually shown in Figure 1.

Table 1: Experimental $x_{B(exp.)}$ and Calculated $x_{B(cal.)}$ Values of Mole Fraction Solubility and Density (ρ) of Hydroquinone for Various Initial Mole Fractions, (x_C^0), of Ethanol at Temperatures (293.15 to 313.15) K and Pressure 101.32 kPa^a.

x_C^0	$x_{B(exp.)}$	$x_{B(cal.)}$	RD	$\rho \cdot 10^{-3}/$ $kg \cdot m^{-3}$	x_C^0	$x_{B(exp.)}$	$x_{B(cal.)}$	RD	$\rho \cdot 10^{-3}/$ $kg \cdot m^{-3}$
293.15 K					305.15 K				
0.0000	0.0102	0.0103	0.0079	1.0090	0.0000	0.0161	0.0162	0.0090	1.0119
0.1008	0.0354	0.0354	0.0012	0.9992	0.1008	0.0536	0.0537	0.0013	1.0081
0.2000	0.0703	0.0702	0.0015	0.9980	0.2000	0.0933	0.0932	0.0010	1.0063
0.3001	0.1003	0.1001	0.0022	0.9941	0.3001	0.1239	0.1236	0.0022	1.0004
0.3999	0.1250	0.1244	0.0044	0.9875	0.3999	0.1473	0.1468	0.0036	0.9921
0.5000	0.1434	0.1433	0.0011	0.9795	0.5000	0.1664	0.1658	0.0037	0.9827
0.5999	0.1575	0.1575	0.0002	0.9698	0.5999	0.1830	0.1817	0.0073	0.9728
0.7002	0.1697	0.1695	0.0008	0.9598	0.7002	0.1928	0.1908	0.0108	0.9619
0.8000	0.1788	0.1790	0.0011	0.9489	0.8000	0.1993	0.1989	0.0022	0.9508
0.8999	0.1834	0.1834	0.0000	0.9379	0.8999	0.2030	0.2027	0.0014	0.9396
1.0000	0.1877	0.1883	0.0030	0.9259	1.0000	0.2063	0.2072	0.0042	0.9280
295.15 K					308.15 K				
0.0000	0.0112	0.0112	0.0000	1.0097	0.0000	0.0180	0.0180	0.0009	1.0127
0.1008	0.0387	0.0386	0.0024	1.0009	0.1008	0.0580	0.0581	0.0019	1.0099
0.2000	0.0745	0.0744	0.0012	1.0001	0.2000	0.0991	0.0984	0.0072	1.0083
0.3001	0.1045	0.1045	0.0000	0.9950	0.3001	0.1292	0.1288	0.0031	1.0016
0.3999	0.1281	0.1285	0.0035	0.9882	0.3999	0.1520	0.1519	0.0006	0.9934
0.5000	0.1475	0.1476	0.0008	0.9795	0.5000	0.1698	0.1705	0.0042	0.9838
0.5999	0.1625	0.1621	0.0028	0.9702	0.5999	0.1885	0.1870	0.0082	0.9732
0.7002	0.1737	0.1737	0.0000	0.9602	0.7002	0.1938	0.1950	0.0066	0.9624
0.8000	0.1829	0.1824	0.0023	0.9491	0.8000	0.2036	0.2036	0.0003	0.9514
0.8999	0.1871	0.1869	0.0007	0.9381	0.8999	0.2070	0.2071	0.0002	0.9398
1.0000	0.1920	0.1918	0.0010	0.9268	1.0000	0.2109	0.2113	0.0019	0.9286
298.15 K					310.15 K				
0.0000	0.0122	0.0122	0.0037	1.0102	0.0000	0.0199	0.0201	0.0071	1.0135
0.1008	0.0419	0.0420	0.0034	1.0024	0.1008	0.0627	0.0628	0.0015	1.0123
0.2000	0.0786	0.0788	0.0027	1.0014	0.2000	0.1032	0.1038	0.0061	1.0098
0.3001	0.1085	0.1090	0.0046	0.9961	0.3001	0.1339	0.1341	0.0012	1.0029
0.3999	0.1322	0.1328	0.0045	0.9892	0.3999	0.1569	0.1572	0.0021	0.9943

0.5000	0.1518	0.1521	0.0021	0.9805	0.5000	0.1751	0.1752	0.0009	0.9845
0.5999	0.1671	0.1668	0.0019	0.9707	0.5999	0.1918	0.1924	0.0031	0.9743
0.7002	0.1775	0.1780	0.0025	0.9606	0.7002	0.1984	0.1993	0.0047	0.9632
0.8000	0.1854	0.1862	0.0039	0.9494	0.8000	0.2072	0.2087	0.0072	0.9525
0.8999	0.1903	0.1906	0.0016	0.9384	0.8999	0.2109	0.2116	0.0034	0.9408
1.0000	0.1964	0.1955	0.0046	0.9271	1.0000	0.2156	0.2156	0.0003	0.9295
300.15 K					313.15 K				
0.0000	0.0134	0.0134	0.0036	1.0108	0.0000	0.0226	0.0224	0.0080	1.0146
0.1008	0.0458	0.0457	0.0029	1.0041	0.1008	0.0679	0.0678	0.0020	1.0148
0.2000	0.0829	0.0834	0.0057	1.0024	0.2000	0.1095	0.1094	0.0003	1.0119
0.3001	0.1135	0.1137	0.0022	0.9974	0.3001	0.1394	0.1396	0.0015	1.0045
0.3999	0.1370	0.1373	0.0020	0.9900	0.3999	0.1627	0.1628	0.0006	0.9957
0.5000	0.1565	0.1566	0.0003	0.9815	0.5000	0.1803	0.1801	0.0012	0.9856
0.5999	0.1701	0.1716	0.0087	0.9713	0.5999	0.1974	0.1979	0.0026	0.9754
0.7002	0.1821	0.1822	0.0009	0.9608	0.7002	0.2044	0.2037	0.0035	0.9641
0.8000	0.1909	0.1901	0.0038	0.9501	0.8000	0.2150	0.2140	0.0046	0.9532
0.8999	0.1944	0.1945	0.0007	0.9384	0.8999	0.2167	0.2163	0.0019	0.9419
1.0000	0.1993	0.1992	0.0003	0.9271	1.0000	0.2206	0.2200	0.0023	0.9302
303.15 K									
0.0000	0.0147	0.0147	0.0025	1.0113	[a]Standard uncertainties in u are				
0.1008	0.0496	0.0495	0.0019	1.0065	$u(T) = 0.1$ K, $u(x_C^0) = 0.0002$,				
0.2000	0.0885	0.0882	0.0033	1.0046	$u(x_B) = 0.003$, and $u(\rho) = 10$ $kg \cdot m^{-3}$.				
0.3001	0.1188	0.1186	0.0016	0.9990	The relative uncertainty in pressure $u_r(p) = 0.05$.				
0.3999	0.1425	0.1419	0.0040	0.9910					
0.5000	0.1615	0.1611	0.0022	0.9819					
0.5999	0.1755	0.1766	0.0061	0.9720					
0.7002	0.1864	0.1865	0.0004	0.9615					
0.8000	0.1943	0.1944	0.0004	0.9502					
0.8999	0.1989	0.1985	0.0019	0.9391					
1.0000	0.2034	0.2031	0.0011	0.9279					

[a]Standard uncertainties in u are u(T) = 0.1 K, u(x_C^0) = 0.0002, u(x_B) = 0.003, and u(ρ) = 10 kg·m^{-3}. The relative uncertainty in pressure ur(p) = 0.05.
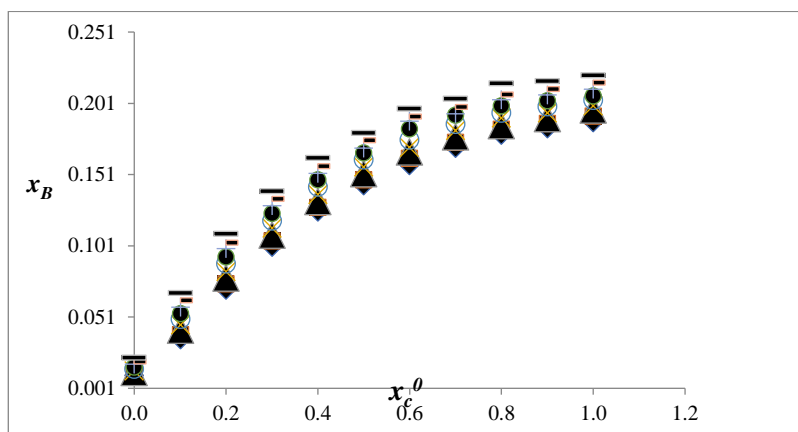

Fig. 1-Mole Fraction Solubility of Hydroquinone (x_B) Variation with Initial Mole Fraction (x_C^0) of Ethanol at Temperatures (♦T=293.15 K, ■T=295.65 K; ▲T=298.15 K; ×T=300.65 K; ○T=303.15 K; ●T=305.65 K; +T=308.15 K; -T=310.65 K & ▬ T=313.15 K).

The solubility of hydroquinone in all solvents increases with temperature. At the same temperature, the solubility trend in solvent is ethanol > water + ethanol >water. This trend implies that solubility of hydroquinone increases with increasing with mole fraction of ethanol, it is prefer to dissolve more in ethanol than water. The solubility of hydroquinone in water-ethanol mixture with x_C^0 increases with increases in x_B up x_C^0=1. This implies that there is strong dipole-dipole interaction between solute and solvent molecules. As temperature increases density goes on decreasing.

But here increase of temperature and mole fraction of ethanol density goes on increases; this is because of increase of solubility with temperature and mole fraction of ethanol.

Apelblat Model:
Among the different methods, the modified semi-empirical Apelblat model (eq 1) is a suitable way to correlate solubility data against temperature[14, 15]. The equation is based on solid-liquid equilibrium theory provide excellent agreement between experimental and

calculated values of solubility[16].

$$ln x_{B=} A + \frac{B}{T} + C ln T \qquad (1)$$

A, B, and *C* are the model parameters and T is temperature in Kelvin. *A* and *B* represent the non-idealities of the solutions in terms of the variation of activity coefficients, *C* reflects to the effect of temperature on the enthalpy of fusion[17]. *A, B,* and *C* parameters are determined using non-linear least square fitting[18]. Solubility values of hydroquinone in water, ethanol and their mixtures are calculated by eq 1. Relative deviation (*RD*) [19] is calculated using eq 2.

$$RD = \frac{x_B^{exp.} - x_B^{cal.}}{x_B^{exp.}} \qquad (2)$$

The data of experimental mole fraction solubility, calculated solubility and *RD* in monosolvent (water, ethanol) and water-ethanol mixtures are listed in Table 1. The values of parameters *A, B, C* along with co-relation coefficient (R^2) are listed in Table 2.

Table 2. Model Parameters and Correlation Coefficient of the Apelblat Equation.

Solvents	Mole fraction x_C^0	Parameters			R^2
		A	B	C	
Ethanol	0.0000	-528.615	20611.26	79.8717	0.999
	0.1008	8.448876	-3051.92	-0.2426	1.000
	0.2000	-4.96045	-1622.33	1.3797	0.999
	0.3001	-26.3022	-208.589	4.3502	1.000
	0.3999	-87.4798	2809.789	13.3454	0.999
	0.5000	-5.71715	-717.335	1.0951	0.999
	0.5999	-41.0135	879.6565	6.3662	0.995
	0.7002	8.305781	-1167.21	-1.0736	0.994
	0.8000	-121.689	4717.937	18.2855	0.996
	0.8999	-71.2879	2498.368	10.7503	0.999
	0.0000	-528.615	20611.26	79.8717	0.999

NIBS-Redlich-Kister Model:
The solubility data at constant temperature is fitted into combined NIBS-Redlich-Kister model[20-23].

$$ln x_B = x_C^0 ln x_1 + x_A^0 ln x_2 + x_C^0 x_A^0 \sum_{i=0}^{3} M_i (x_C^0 - x_A^0)^i \qquad (3)$$

Where x_A^0 is initial mole fraction of water and x_1, x_2 are solubilities of hydroquinone in pure ethanol and water respectively. M_i is curve fit parameters (four parameter equation). All values of M_i along with R^2 value are listed in Table 3. The values of R^2 are close to unity shows that NIBS-Redlich-Kister model is very well applicable for this solubility data.

Thermodynamics Functions of Dissolution:
According to the van't Hoff equation, the standard molar enthalpy change of solution ΔH^0_{soln} is generally obtained from the slope of the *ln* x_B vs *1/T* plot. Average temperature T_{mean} is introduced to obtain a single value of ΔG^0_{soln} and ΔS^0_{soln} in the temperature range studied.

$$T_{mean} \frac{n}{\sum_{i-1}^{n}\left(\frac{1}{T}\right)} \qquad (4)$$

Where *n* is the number of experimental points. In the present work, T_{mean} = 302.98 K and the temperature range is (293.15 to 313.15) K in both pure solvents and binary solvent mixtures. Heat capacity of the solution can be assumed as constant. Hence values of ΔH^0_{soln} are derived using eq 5.

$$\Delta H^0_{Sol} = -R \left(\frac{\partial ln x_B}{\partial 1/T}\right) - R \left[\frac{\partial ln x_B}{\partial \left(\frac{1}{T} - \frac{1}{T_{mean}}\right)}\right] \qquad (5)$$

The $ln x_B$ Vs 10000 (*1/T − 1/Tmean*) plot of different solutions including pure solvents and binary solvent mixtures are displayed in Figures 2. From these figures, it can be seen that a trend of increasing solubility with temperature is observed. The slope and the intercept for each solvent are listed in Table 4. Thus the modified van't Hoff equation can be thought to be fit to calculate the enthalpy change of solution. The standard molar Gibbs energy change for the solution process ΔG^0_{soln}, can be calculated in the way similar to Krug et al [24] as

$$\Delta G^0_{sol} = -RT \times intercept \qquad (6)$$

Table 3. NIBS-Redlich-Kister model parameters.

T/K	Range of x_C^0	M_0	M_1	M_2	M_3	R^2
Water + Ethanol + Hydroquinone						
293.15	0.101-0.90	4.730	-3.569	3.161	-1.801	1.000
295.65	0.101-0.90	4.584	-3.397	3.271	-2.166	1.000
298.15	0.101-0.90	4.470	-3.305	3.327	-2.439	1.000
300.65	0.101-0.90	4.352	-3.159	3.518	-2.712	1.000
303.15	0.101-0.90	4.244	-3.129	3.575	-2.748	0.999
305.65	0.101-0.90	4.170	-2.939	3.630	-3.023	0.999
308.15	0.101-0.90	4.219	-2.998	4.342	-3.550	0.998
310.65	0.101-0.90	3.845	-2.750	3.579	-3.123	0.999
313.15	0.101-0.90	3.676	-2.572	3.510	-3.082	0.999

In which the intercept is used to obtained from plots of $ln x_B$ as a function of (*1/T − 1/Tmean*). The standard molar entropy change ΔS^0_{soln} is obtained from

$$\Delta S^0_{Sol} = \frac{\Delta H^0_{Sol} - \Delta H^0_{Sol}}{T_{mean}} \qquad (7)$$

Both ΔG^0_{soln} and ΔS^0_{soln} pertain to the mean temperature T_{mean} = 302.92 K.

The results are shown in Table 5, together with %ζH and %ζTS. It is worthy to note that relative contribution of enthalpy %ζH and %ζTS which are defined as

$$\%\zeta_H = \frac{\Delta H_{Sol}^0}{|\Delta H_{Sol}^0| + |T\Delta S_{Sol}^0|} X 100 \qquad \text{............ (8)}$$

$$\%\zeta_{TS} = \frac{|T\Delta S_{Sol}^0|}{|\Delta H_{Sol}^0| + |T\Delta S_{Sol}^0|} X 100 \qquad \text{............ (9)}$$

can be simply used to calculate the main contributors of enthalpy or entropy to $\Delta G^0{}_{soln}$ [25].

The values of ΔH^0 and ΔS^0 for all solutions are positive indicating the solution process as endothermic. The contribution of enthalpy to positive molar Gibbs energy is more as compared to entropy for all solutions.

Table 4. Slope(*m*) and Intercept (*c*) of the *lnx_B* vs. 10000(*1/T − 1/T_mean*) Plot along with *R²*

Water + Ethanol+Hydroquinone			
x_c^0	*m*	*c*	*R²*
0.0000	-3584	-4.201	0.995
0.1008	-2978	-3.005	0.999
0.2000	-2040	-2.427	0.999
0.3001	-1526	-2.131	0.999
0.3999	-1233	-1.949	0.997
0.5000	-1049	-1.825	0.999
0.5999	-1048	-1.732	0.994
0.7002	-841.9	-1.679	0.993
0.8000	-821.3	-1.633	0.992
0.8999	-758.2	-1.614	0.997
1.0000	-715.1	-1.591	0.996

Table 5. Thermodynamic Functions Relative to Solution Process of Hydroquinone at *Tmean* =302.928K

x_c^0	ΔH^0sol/kJ·K^{-1}·mol^{-1}	ΔG^0soln/kJ·K^{-1}·mol^{-1}	ΔS^0soln/KJ·K^{-1}·mol^{-1}	TΔS^0soln/KJ·K^{-1}·mol^{-1}	%ζH	%ζTS
Water + Ethanol						
0.0000	29.7974	10.5882	0.0634	19.2092	0.6080	0.3920
0.1008	24.7591	7.5738	0.0567	17.1853	0.5903	0.4097
0.2000	16.9606	6.1170	0.0358	10.8436	0.6100	0.3900
0.3001	12.6872	5.3709	0.0241	7.3162	0.6343	0.3657
0.3999	10.2512	4.9122	0.0176	5.3389	0.6575	0.3425
0.5000	8.7214	4.5997	0.0136	4.1217	0.6791	0.3209
0.5999	8.7131	4.3653	0.0143	4.3478	0.6671	0.3329
0.7002	6.9996	4.2317	0.0091	2.7678	0.7166	0.2834
0.8000	6.8283	4.1158	0.0089	2.7125	0.7157	0.2843
0.8999	6.3037	4.0679	0.0074	2.2358	0.7382	0.2618
1.0000	5.9453	4.0099	0.0064	1.9354	0.7544	0.2456

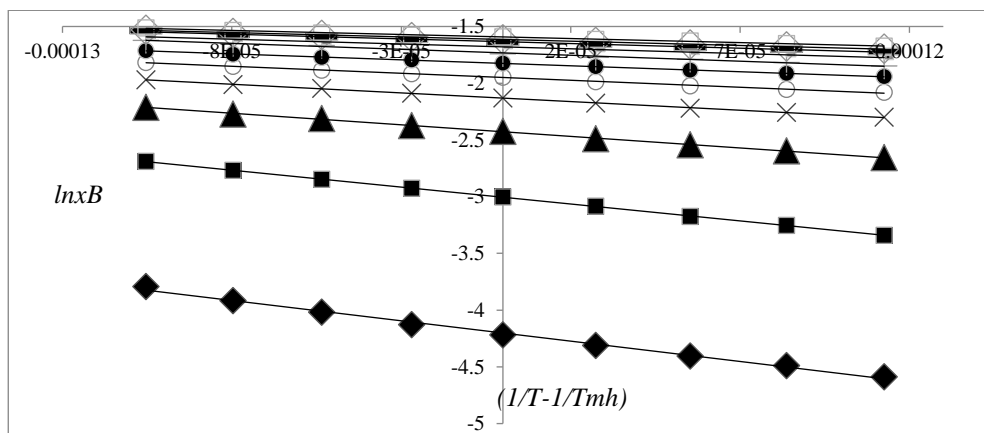**Fig. 2-Plot of *lnx_B* vs. (*1/T − 1/T_hm*) for Hydroquinone + Water + Ethanol System at various Mole fractions.** (◆x_C^0=0.0000; ■ x_C^0=0.1001; ▲ x_C^0=0.2002; × x_C^0=0.3001; ○x_C^0=0.4000; ● x_C^0=0.5000; + x_C^0=0.6020; - x_C^0=0.7002; — x_C^0=0.7998 and ◇ x_C^0=0.9000; □ x$_C^0$=1.0000).

Density values are used to calculate excess molar functions[26].

FTIR Spectra:
Hydroquinone has two identical –OH groups, showing υ-OH symmetrical stretching frequency at 3224.98 cm^{-1}. In ternary system, υ-OH decreases with increase with mole fraction of alcohol because hydroquinone interaction with binary solvent increase with increase of mole fraction of alcohol s observed in Table 6. The υ-OH decreases as amount of alcohol in mixture increases

this is because frequency is after interaction of –OH of alcohol and –OH of water. The υ-OH of alcohol is less than that of water therefore as amount of alcohol increases in mixture υ-OH decreases as shown in Fig. 3, 4, 5, 6 and 7. Solubility of hydroquinone increases with increasing amount of alcohol because hydroquinone is more soluble in alcohol than in water. This is due to solvent-solvent and solute-solvent interaction in terms of hydrogen bonding. More is the hydrogen bonding, lower the υ-OH and more is the solubility of hydroquinone.

Table 6. Experimental υ-OH for 0, 0.2, 0.5, 0.7 and 1 mole fraction of ethanol (x_C^0) in ternary solutions.

Ethanol (x_C^0)	Experimental υ-OH Hydroquinone+ Water+ Ethanol cm^{-1}	Solid Hydroquinone υ-OH cm^{-1}
0(Water)	3387	
0.2	3387	
0.5	3379.29	3224.98
0.7	3363.86	
1(Ethanol)	3332.99	
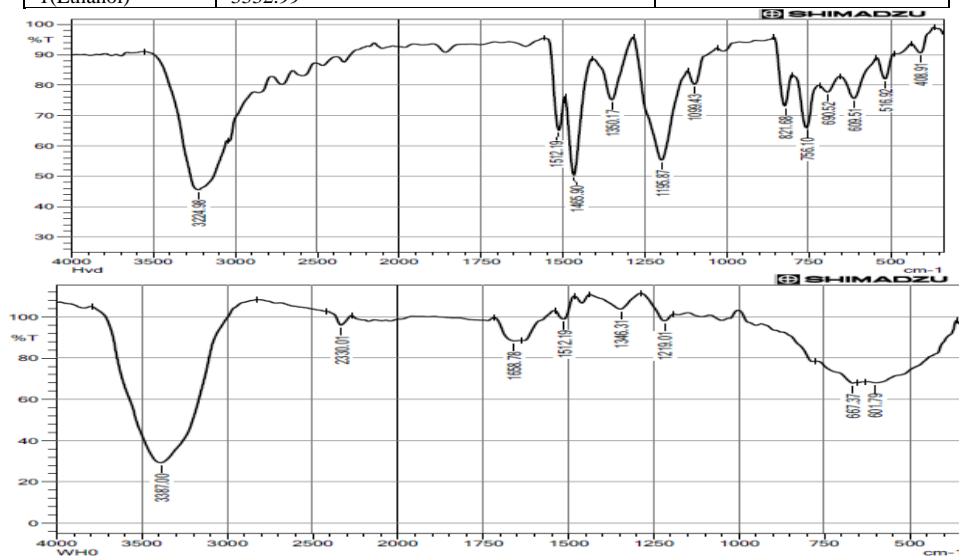
Fig. 3- FTIR spectra for Hydroquinone and Hydroquinone+Water.
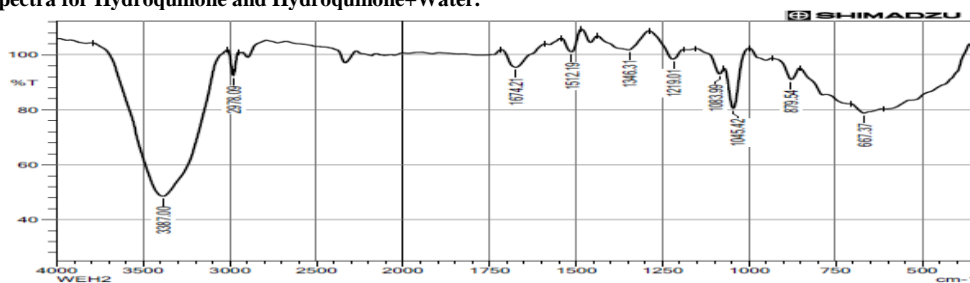
Fig. 4- FTIR spectra for solution of Hydroquinone in 0.2 mole fraction of Ethanol
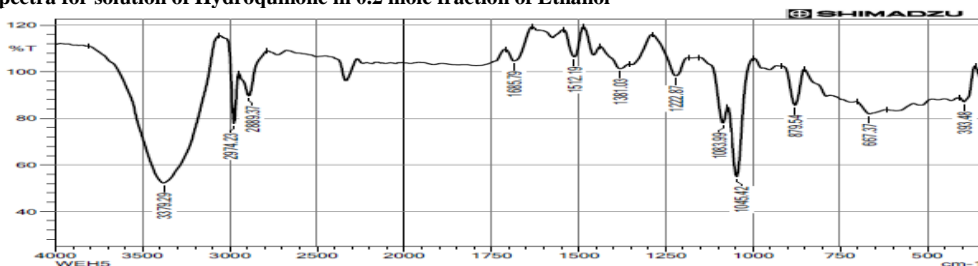
Fig. 5- FTIR spectra for solution of Hydroquinone in 0.5 mole fraction of Ethanol
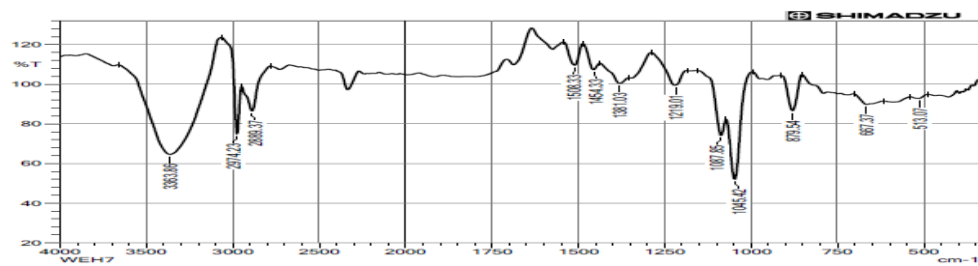
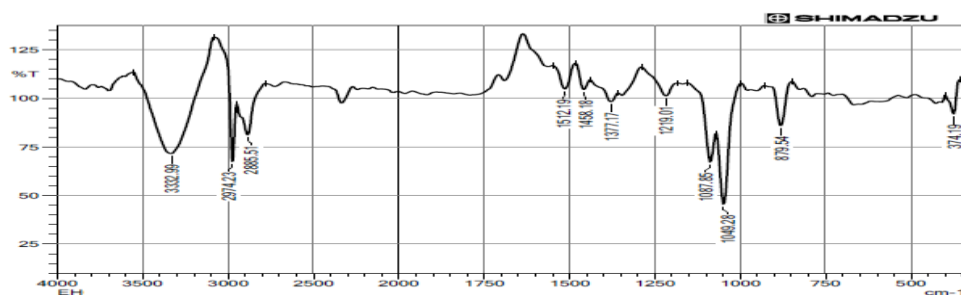**Fig. 6- FTIR spectra for solution of Hydroquinone in 0.7 mole fraction of Ethanol**

Fig. 7- FTIR spectra for Hydroquinone+Ethanol

CONCLUSIONS:

Solubility data and Thermodynamic functions including ΔH^0_{soln}, ΔG^0_{soln}, and ΔS^0_{soln} of hydroquinone in water, ethanol and their mixture are more useful in field of physical chemistry and chemical engineering calculations involving fluid flow, heat and mass transfer, pharmaceutical industry, agriculture, biology, medicine. Solubility data is required for selection of proper solvent and design an optimized crystallization process.

The solubility of hydroquinone is more in ethanol than in water and increases with increase in mole fraction of ethanol. ΔH^0_{soln} values are higher in water than water+ethanol mixture and lowest in ethanol indicates the solubility trend in various solvents. Density of solution is depends on solubility and solvent system both. Increase of density with temperature and mole fraction of ethanol indicates higher solubility. Also lower values of υ-OH shows increase of molecular interaction and hence the increases solubility of hydroquinone.

REFERENCES:

1. ODonoghue JL. Hydroquinone and its analogues in dermatology—a risk-benefit viewpoint. Journal of Cosmetic Dermatology. 2006; 5(3):196–203.
2. Enguita FJ. Leitão AL. Review Article: Hydroquinone: Environmental Pollution, Toxicity and Microbial Answers. Hindawi Publishing Corporation. BioMed International. 2013; Article ID 542168:14. doi.org/10.1155/2013/542168
3. Jurica K. et al. Arbutin and its metabolite hydroquinone as the main factors in the antimicrobial effect of strawberry tree (Arbutus unedo L.) leaves. Journal of Herbal Medicine. 2017; 172:S2210-8033(17)30020-9. doi.org/doi:10.1016/j.hermed.2017.03.006
4. Francisco J. Enguita. Leitão A L. Hindawi Publishing Corporation. BioMed Research International. 2013; Article ID 542168. doi.org/10.1155/2013/542168
5. Jagtap S. Magdum C. Influence of Water-soluble polymers on Epalrestat ternary complexation by kneading. Research. J. Pharm. and Tech 2019; 12(8):3602-3608.
6. Mishra R. Solubility Enhancement of Clarithromycin Using Solid Dispersion and Effervescence Assisted Fusion Technique Research J. Pharm. and Tech. 2016; 9(6):677-686.
7. Pawar RR, et al. Solubility, density and solution thermodynamics of NaI in different pure solvents and binary mixtures. J. Chem. Eng. Data. 2012; 5:3563-3572.
8. Pawar RR. et al. Solubility and density of potassium iodide in binary water-ethanol solvent mixtures at (298.15, 303.15, 308.15, and 313.15) K. J. Chem. Eng. Data. 2009; 54:935−37.
9. Pawar RR, et al. Solubility and density of potassium iodide in binary water-propan-1-ol solvent mixtures at (298.15,303.15,308.15 and 313.15) K. J. Chem. Eng. Data. 2010; 55: 1314−16.
10. Kadam UB. et al. Density, Viscosities and Ultrasonic Velocity Studies of Binary Mixtures of Chloroform with Propan-1-ol and Butan-1-ol at (303.15 and 313.150) K. J. Chem. Eng. Data. 2006; 51:60−63.
11. Hasan M. et al. Densities, Viscosities and Speed of Sound Studies of Binary Mixtures of Methylbenzene with Heptan-1-ol, Octan-1-ol and Decan-1-ol at (303.15 and 313.15) K. J. Chem. Eng. Data. 2006; 51:1922−26.
12. Marsh KN. Recommended Reference Materials for the Realisation of Physicochemical Properties. Blackwell Scientific Publications: Oxford. U.K. 1987.
13. Aher CS. Excess molar volume of aqueous 1-propanol and apparent molar volume of 2-naphthol in water, 1-propanol and their binary mixtures at various temperatures with DFT Study. J. Applicable Chem. 2018; 7(1):151-164.
14. Lourenço C. Melo C I. Bogel-Łukasik R. Bogel-Łukasik E. J. Chem. Eng. Data. 2012; 57:1525–1533.
15. Melo CI. Bogel-Łukasik R. Nunes da Ponte M. Bogel-Łukasik E. Fluid Phase Equilibria. 2013; 338:209–216.
16. Domańska U. Bogel-Łukasik R. J. Phys. Chem. B. 2005; 109:12124–12132.
17. Apelblat A. Manzurola E. J. Chem. Thermodyn. 1999; 31:85–91.
18. Gans P. et al. Investigation of equilibria in solution determination of equilibrium constants with hyperquad suite of program. Talanta. 1996; 43:1739–1753.
19. Guo H. et al. Measurement and correlation of solubility of pronamide in five organic solvents at (278.15-323.15)K. J. Chem. Thermodynamics. 2016; 99:11-15.
20. Jouyban A. et al. Solubility prediction of anthracene in mixed solvents using a minimum number of experimental data. Chem. Pharm. Bull. 2002; 50(1):25−25.
21. Joyce R. et al. Solubility of Anthracene in binary alcohol + 1,4-dioxane solvent mixtures. J. Chem. Eng. Data. 1995; 40:1124−1126.
22. Acree WE. Zvaigzne AI. Thermodynamic properties of nonelectrolyte solutions: part 4. estimation and mathematical representation of solute activity coefficients and solubilities in binary solvents using the NIBS and modified Wilson equations. Thermochim. Acta. 1991; 178:151−167.
23. Acree WE. Jr. Mathematical representation of thermodynamic properties: Part 2. Derivation of combined nearly ideal binary solvents NIBS/Redlich-Kister mathematical representation from a two body and three body interactional mixing model. Thermochim. Acta. 1992; 198:71−79.
24. Krug R. Hunter W G. Grieger R A. Enthalpy-Entropy Compensation. 2. Separation of the Chemical from the Statistical Effect. J. Phys. Chem. 1976; 80:2341−2351.
25. Cui Z. Yin Q. Gong J. WangY. Hao H. Xie C. Bao Y. Zhang M. Hau B. Wang. J. Fluid Phase Equilib. 2013; 337:354-362.
26. Mahajan MM. Raghuwanshi PB. Ultrasonic Velocity and Related Acoustical Parameters of N–(2-Hydroxybenzylidene)-3-Substituted Pyridine-2-Amine Schiff Bases in Ethanol-Water Mixture. J. Applicable Chem. 2016; 5(5):1208-1216.