Reviews

Solubility of Total Reduced Sulfurs (Hydrogen Sulfide, Methyl Mercaptan, Dimethyl Sulfide, and Dimethyl Disulfide) in Liquids

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The solubility of the total reduced sulfurs, which include hydrogen sulfide (H_2S), methyl mercaptan (methanethiol, CH_3SH), dimethyl sulfide (CH_3SCH_3), and dimethyl disulfide ($CH_3S_2CH_3$), is of great interest in various applications in the chemical, oil, and gas industries and in environmental protection as well. They can occur naturally in the environment and can also be present in numerous industrial gaseous streams (petroleum, natural gas, some chemical industries like the pulp and paper industry). The aim of this review is to update different aspects concerning the solubility data of these compounds in various liquids, which are essential for the design and operation of absorption scrubbing equipment and/or of interest in many technical areas (e.g., the petroleum and natural gas industry). The review deals with the compound's characterization in direct connection with their source and a survey of relevant existing experimental data including (i) all data concerning the solubility of methyl mercaptan, dimethyl sulfide, and dimethyl disulfide published generally before January 2006 as well as (ii) data concerning the solubility of hydrogen sulfide published generally after January 1987. The liquids include water, aqueous electrolyte solutions, nonaqueous solvents, and alkanolamines. Both pure compounds and mixtures are considered.

Introduction

The quartet of total reduced sulfurs (TRS) includes hydrogen sulfide (H_2S), methyl mercaptan (methanethiol, CH_3SH), dimethyl sulfide (CH_3SCH_3), and dimethyl disulfide ($CH_3S_2CH_3$). These volatile compounds are notoriously known for causing malodorous air pollution. Their solubility is of great interest in various applications in the chemical industry, in the oil and gas industry, and in environmental protection as well. On the one hand, hydrogen sulfide is the most common sulfur-containing compound encountered in natural gas and in light and middle distillate oil fractions sweetening, in heavy oil hydrotreating, and in various non-petroleum applications. On the other hand, TRS are the major components known to occur in the Kraft pulp mill emissions and are responsible for their distinctive odor problems at very low concentrations in the neighboring agglomerations. The strong and persistent odor is due to the use of the white liquor (alkaline solutions containing sodium sulfide and sodium hydroxide) for the digestion of wood and the conversion into pulp. During operation, sodium sulfide combines with organic compounds to form organosulfur compounds (mostly organic sulfides and mercaptans).

Solubilities of sulfur compounds in liquids are essential for the design and operation of absorption scrubbing equipment in many technical applications. First, this review deals with the solubility of these compounds in pure and mixed physical and chemical solvents related to processes concerning their removal from gaseous streams by scrubbing/regeneration. Second, because of the importance of reliable solubility data of pollutants for environmental processes as well as the impact of these sulfur compounds on the earth's climate and radiation balance, data concerning their solubility in aqueous electrolyte solutions (for example, in saline solutions) have also been included. Third, due to the presence of sulfur compounds (namely, hydrogen sulfide and methyl mercaptan) in natural gas and in light and heavy hydrocarbon fractions, their solubility has also been covered in the review.

The main directions explored in this review are as follows: (i) characterization of the sulfur compounds in direct connection with their sources and (ii) a survey of relevant existing experimental data including (i) all data concerning the solubility of methyl mercaptan, dimethyl sulfide, and dimethyl disulfide published generally before January 2006 and (ii) data concerning the solubility of hydrogen sulfide published generally after 1987, taking into account that a complete IUPAC collection already covered the data published before January 1987.[1] In addition, a review of experimental data for temperatures between (273 and 363) K and pressures up to 1 MPa was published by Carroll and Mather.[2] Recently, Chapoy et al.[3] have also included in their paper a short review of hydrogen sulfide solubility in water.

The liquids include water, aqueous electrolyte solutions, nonaqueous solvents, and alkanolamines. Both pure compounds and mixtures are considered.

Physical and Chemical Properties of Total Reduced Sulfurs

Hydrogen sulfide (H_2S) and methyl mercaptan (MM) are gases at ambient temperature, while dimethyl sulfide (DMS) and dimethyl disulfide (DMDS) are low boiling-point volatile liquids. Only H_2S and MM can dissociate in aqueous solution:[4]

$$H_2S \underset{}{\overset{K_1}{\rightleftharpoons}} HS^- + H^+ \overset{K_2}{\rightleftharpoons} S^{2-} + 2H^+$$

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$$CH_3SH \overset{K}{\rightleftharpoons} CH_3S^- + H^+$$

TRS components are highly corrosive and toxic gases.[4,5] As an example, hydrogen sulfide irritates eyes and respiratory tract at a concentration of only 20 $\mu L \cdot L^{-1}$; at 500 $\mu L \cdot L^{-1}$, 30 min of exposure causes severe sickness whereas 1000 $\mu L \cdot L^{-1}$ and 30 min of exposure cause death.[6] Methyl mercaptan and dimethyl disulfide appear to be somewhat less toxic than hydrogen sulfide but can produce similar effects only at slightly higher concentrations. Dimethyl sulfide is considerably less toxic. At low concentration, hydrogen sulfide is responsible for an odor of rotten eggs, while at high lethal concentrations, it is odorless. Hydrogen sulfide has a very wide explosion limit of (4.3 to 45.5) % (volume fraction) in air as compared to the other three sulfur compounds.

Industrial Implication for Sulfur Compounds

Hydrogen sulfide and organic sulfides may be present in gaseous or liquid hydrocarbon streams. A large fraction of the available natural gas contains a significant amount of acid gases (mainly hydrogen sulfide and carbon dioxide as well as traces of COS and mercaptans), whose removal is important for avoiding offensive odor, upgrading its heating value, and making the gas suitable for pipeline transmission and utility use. Hydrogen sulfide is the most common sulfur-containing gas component in the natural or synthetic gas and has both corrosive and toxic properties. Removal of hydrogen sulfide by the so-called sweetening process can allow the production and recovery of elemental sulfur. In the natural gas industry, almost complete removal of hydrogen sulfide is required, while carbon dioxide is allowed to some extent in the sales gas.[7] When only a little carbon dioxide is present in the raw gas, it is advantageous to remove the hydrogen sulfide only, thus avoiding the carbon dioxide removal cost. Moreover, carbon dioxide is not a waste stream in some cases but a useful product, as in the case of enhanced oil recovery, which accounts for the largest industrial use of CO_2. In this process, the selective desulfurization is followed by carbon dioxide removal, allowing its utilization.

The conventional technique for sweetening natural gas uses a liquid solvent for removing the objectionable gases through absorption. Low partial pressures of these compounds in the natural gas are generally achieved by the use of chemical solvents, especially aqueous alkanolamine solutions. Alternatively to this energy-intensive process, the use of mixtures of physical (polar organic) and chemical (alkanolamines) solvents can be considered as a more economical option. The present review will consider only the elimination of sulfur compounds.

Hydrogen sulfide and mercaptans are also constituents of many petroleum reservoirs. Their presence should therefore be taken into account for accurate multi-component phase equilibria modeling. Hydrogen sulfide should also be eliminated by hydrotreating petroleum fractions to prevent catalyst deactivation during subsequent catalytic cracking. Knowledge of the solubility data of these sulfur compounds in various hydrocarbons, mainly alkanes, is necessary for calculation and prediction of the phase behavior and other thermodynamic properties of such systems.

TRS referred to as noncondensible gases (NCG) are also part of a well-known environmental problem afflicting pulp mills exploiting the Kraft pulp mill process. They are emitted from digesters, turpentine recovery systems, evaporators, brownstock washer hoods and seal tanks, knotter hoods, mud filters, causticizers, and liquor and brownstock storage tanks.[6] They are formed in the Kraft pulping process, and because they come in contact with different waters and liquors, they contribute to the odor problems of Kraft mills. Among the TRS, H_2S is the most abundant in the effluents. TRS are highly corrosive to carbon steel. Moreover, they are highly toxic, being responsible for deaths and injuries in the pulp and paper industry. Since the early 1990s, several Canadian provincial governments and the United States promulgated a number of regulations upon the Kraft pulp manufacturers to collect and treat their TRS emitting vents. The olfactory threshold of TRS for human beings is 4 orders of magnitude below the regulated emission level, which is approximately 5–10 $\mu L \cdot L^{-1}$ in Canada and the United State. This has ultimately given rise to strict regulations in order to reduce the emissions from specific sulfate pulp process equipments such as kilns, evaporators, washers, etc. Considering the progressive nature of legislations, it is anticipated that increasingly tighter regulations will be applicable in the near future especially in North America where about 15 % of the world's Kraft mills are in operation. Because of their toxic and corrosive characters, they must be removed down to very low concentration levels. Various approaches have been advocated over the years to reduce the concentration of these pollutants in the effluents below environmentally acceptable limits.[8,9]

Methylated Sulfur Compounds (MM, DMS, and DMDS) in the Natural Environmental Processes

The organosulfur compounds in general have an important impact on the environmental processes. Methylated sulfur follows a complex cycle in aquatic ecosystems, involving dimethyl sulfide (DMS) and dimethyl sulfoxide (DMSO), in addition to various volatile and nonvolatile species like methyl mercaptan (MM), DMDS (dimethyl disulfide), and dimethyl sulfoniopropionate (DMSP).[10] DMS is the most abundant volatile sulfur compound in the seawater and a climatically active trace gas in the atmosphere. It is produced by the decomposition of DMSP, which is synthesized at the surface of ocean by marine phytoplankton.[11] It is supersaturated in most natural waters, and there is therefore a net flux of this component into the atmosphere. Dacey et al.[12] mentioned that the modeling of the global sulfur cycle indicates a significant flux of sulfur from the oceans to the atmosphere. In the air, DMS is oxidized to sulfate and sulfonate aerosols that play a key role in the atmospheric acidity and the formation of cloud condensation nuclei over the oceans.[10] Its impact on the earth's radiation balance and climate was also discussed by Bates et al.[13] Knowledge about MM, DMS, and DMDS solubility and their emission in the atmosphere is therefore important for climate studies and global change predictions.

Gas Solubility Data

Tables 1 to 4 present all experimental data published in the open literature concerning the solubility of hydrogen sulfide (from 1987 to 2006, generally before January 2006) and MM, DMS, and DMDS (generally before January 2006). Hydrogen sulfide solubility data reported by Lee et al.[14] were also considered here because they have not been included in the solubility data compilation by Fogg and Young.[1] Temperature and total pressure—gas partial pressure whenever specified—ranges as well as the estimated errors given by the authors (when available) are included.

Hydrogen Sulfide Solubility. As the measurements for the water–hydrogen sulfide system attracted many researchers and the available data are quite abundant for comparison, this system is more extensively discussed here. In Figure 1, some experimental Henry's constants for hydrogen sulfide in water are

Table 1. Hydrogen Sulfide Solubility Data from Literature

T/K	P/kPa	estd uncertainty	ref	T/K	P/kPa	estd uncertainty	ref
Water							
296.65−367.65	101	not given	25	298.15, 313.15	101	$\delta T = \pm 0.1$ K	110
						$\delta P = \pm 0.14$ kPa	
299.85	492.43−1659.66	$\delta T = \pm 0.04$ K	27	298.15, 313.15, 333.15	not given	$\delta T = \pm 0.1$ K	61
		$\delta P = \pm 1.4$ kPa				$\delta P = \pm 0.15$ kPa	
293.95−594.15	222−13861	$\delta T = \pm 0.2$ K for	18	293.15, 303.15,	$P_{H_2S} = 30−80$	$\delta T = \pm 0.01$ K	75
		$T < 473$ K		313.15, 323.15		$\delta P/P = \pm 0.0025$	
		$\delta P/P = \pm 0.0001$				$\delta x_{H_2S} = \pm 1$ %	
		$\delta \log (K_H) = \pm 0.01$					
313.15	470.4−2489.5	$\delta T = \pm 0.1$ K	26	298.16	503−797	$\delta T \pm 0.002$ K	3
		$\delta P/P = \pm 0.001$		308.2	483−2483	$\delta P \pm 1$ kPa	
				318.21	507−3094	$\delta n_{H_2S}/n_{H_2S} = \pm 0.02^a$	
				328.28	497−3475		
				338.34	509−3962		
Aqueous (1−5) mol·kg^{-1} NaCl				Aqueous (0.2−2.5) mol·kg^{-1} NaCl			
296.15−369.15	101	not given	25	428.4	1196−1405	$\delta \log(K_H) = \pm 0.01$	18
				489.55	2743−2787		
				593.65	13721−13842		
Aqueous ≈(4 and 6) mol·kg^{-1} NaCl				Aqueous ≈(2 and 4) mol·kg^{-1} $(NH_4)_2SO_4$			
313	299.2−2853	$\delta T = \pm 0.1$ K	31	313	194.6−2829	$\delta P = \pm 1$ kPa for	31
333	429.8−4299	$\delta P = \pm 1$ kPa for		333	192.7−4266	$P < 0.5$ MPa	
353	333.1−6220	$P < 0.5$ MPa		353	98.9−6145	$\delta P = \pm 4$ kPa for	
393	248.3−9700	$\delta P = \pm 4$ kPa for		393	495.9−9542	$P > 0.5$ MPa	
		$P > 0.5$ MPa					
Aqueous ≈(0.5 and 1) mol·kg^{-1} Na_2SO_4				Aqueous ≈6 mol·kg^{-1} NH_4Cl			
313	331.2−2854	$\delta P = \pm 1$ kPa for	31	313	117.4−2855	$\delta P = \pm 1$ kPa for	31
333	457.6−4272	$P < 0.5$ MPa		353	647−6239	$P < 0.5$ MPa	
353	569.1−6154	$\delta P = \pm 4$ kPa for		393	1224−9659	$\delta P = \pm 4$ kPa for	
393	843−9784	$P > 0.5$ MPa				$P > 0.5$ MPa	
Aqueous ≈(3 and 6) mol·kg^{-1} $NaNO_3$				Aqueous ≈6 mol·kg^{-1} NH_4NO_3			
313	160.7−2829	$\delta P = \pm 1$ kPa for	30	313	319.8−2846	$\delta P = \pm 1$ kPa for	30
333	105.3−4297	$P < 0.5$ MPa		353	477−6236	$P < 0.5$ MPa	
353	307.7−6186	$\delta P = \pm 4$ kPa for		393	823−8325	$\delta P = \pm 4$ kPa for	
393	444.7−9393	$P > 0.5$ MPa				$P > 0.5$ MPa	
Aqueous ≈(1 and 2) mol·kg^{-1} NaOH				Aqueous ≈(4 and 5.8) mol·kg^{-1} CH_3COONa			
313	15.1−2855	$\delta P = \pm 1$ kPa for	30	313	183.4−2831	$\delta P = \pm 1$ kPa for	29
333	19.7−4308	$P < 0.5$ MPa		333	162−4268	$P < 0.5$ MPa	
353	47.3−6245	$\delta P = \pm 4$ kPa for		353	128.1−6185	$\delta P = \pm 4$ kPa for	
393	190.5−9319	$P > 0.5$ MPa		393	662−9708	$P > 0.5$ MPa	
Aqueous ≈6 mol·kg^{-1} CH_3COONH_4							
313	200.3−2850	$\delta P = \pm 1$ kPa for	29				
353	260−6171	$P < 0.5$ MPa					
393	382−8824	$\delta P = \pm 4$ kPa for					
		$P > 0.5$ MPa					
Methanol							
298.15	86.5−2022	$\delta T = \pm 0.1$ K	32	298.15	18.3−414	$\delta T = \pm 0.02$ K	33
348.15	176−5800	$\delta P/P = \pm 0.001$				$\delta P = \pm 0.2$ kPa	
398.15	924−10100	$\delta x_{H_2S} = \pm 0.02$ %				$\delta x_{H_2S} = \pm 0.0001$	
448.15	2450−11200						
Isopentane				Neopentane			
323.15	210−3620	$\delta T = \pm 0.1$ K	35	323.15	359−3620	$\delta T = \pm 0.1$ K	35
353.15	470−6447	$\delta P/P =$		353.15	734−6447	$\delta P/P =$	
383.15	912−8377	$\pm (0.001−0.002)$		383.15	1358−7453	$\pm (0.001−0.002)$	
413.15	1608−7315	$\delta x_{H_2S} = \pm 0.001$		413.15	2308−5454	$\delta x_{H_2S} = \pm 0.001$	
Hexane				Cyclohexane			
322.95	430−3200	$\delta T = \pm (0.1−0.2)$ K	38	323.05	400−3035	$\delta T = \pm (0.1−0.2)$ K	38
372.95	985−7390	$\delta P = \pm 5$ kPa		372.55	1150−7190	$\delta P = \pm 5$ kPa	
422.65	910−7545	$\delta x_{H_2S} =$		422.65	1085−9495	$\delta x_{H_2S} =$	
		$\pm (0.0008−0.007)$				$\pm (0.001−0.007)$	
Isooctane				*n*-Decane			
323.15	239−1043	$\delta T = \pm 0.1$ K	43	323.15	233−995	$\delta T = \pm 0.02$ K	43
373.15	367−1136	$\delta P = \pm 1$ kPa		373.15	227−994	$\delta P = \pm 1$ kPa	
423.15	557−1305			423.15	260−1036		
473.15	1242−1658			473.15	405−1179		
				523.15	688−1290		
n-Dodecane				*n*-Tridecane			
313.2−434.4	524−5675	$\delta T = \pm 0.1$ K	49	323.15	237−788	$\delta T = \pm 0.02$ K	43
		$\delta P/P = \pm 0.001$		473.15	277−981	$\delta P = \pm 1$ kPa	
		$\delta x_{H_2S} = \pm 0.002$		523.15	357−1123		

Table 1 (Continued)

T/K	P/kPa	estd uncertainty	ref	T/K	P/kPa	estd uncertainty	ref
Pentadecane							
422.6	1130−11210	$\delta T = \pm (0.1-0.2)$ K $\delta P = \pm 5$ kPa $\delta x_{H_2S} = \pm (0.005-0.01)$	38				
n-Hexadecane							
323.15	221−987	$\delta T = \pm 0.02$ K	43	323.3	529−3044	$\delta T = \pm 0.1$ K	46
373.15	207−1002	$\delta P = \pm 1$ kPa		373.3	549−7063	$\delta P/P = \pm 0.001$	
423.15	233−982			423.2	584−7414	$\delta x_{H_2S} = \pm 0.002$	
473.15	214−978						
523.15	258−1028						
n-Eicosane				Squalane			
323.0	404−1602	$\delta x_{H_2S} = \pm 0.002$	51	323.15	213−1041	$\delta T = \pm 0.02$ K	
361.3	544−3044			373.15	265−992	$\delta P = \pm 1$ kPa	
423.3	458−3064			423.15	192−1023		
				473.15	240−1024		
				523.15	282−1048		
Benzene							
323.15	455−3230	$\delta T = \pm (0.1-0.2)$ K	38	304.30	102.4−1026	$\delta T = \pm 0.02$ K	33
372.65	930−7155	$\delta P = \pm 5$ kPa		323.50	132.2−1180.5	$\delta P = \pm 0.2$ kPa	
422.65	1100−9800	$\delta x_{H_2S} = \pm (0.001-0.007)$				$\delta x_{H_2S} = \pm 0.0001$	
n-Propylbenzene				Hexane + Pentadecane			
313.2	395−2570	$\delta T = \pm 0.2$ K	54	424.5	1215−7520	$\delta T = \pm 0.3$ K	38
393.4	1040−9290	$\delta P = \pm (5-10)$ kPa				$\delta P = \pm 10$ kPa	
473.5	1510−12970					$\delta x_{H_2S} = \pm (0.003-0.009)$	
n-Hexadecane + *n*-Eicosane				EG			
323.2	1000, 2000, 3000	$\delta T = \pm 0.1$ K	52	298.15	3.24−2030	$\delta T = \pm 0.5$ K	55
		$\delta P/P = \pm 0.001$		323.15	3.2−3520	$\delta P/P = \pm 0.001$	
		$\delta x_{H_2S} = \pm 0.002$		348.15	4.9−5660		
				373.15	3.64−6750		
				398.15	6.46−6460		
Aqueous w(EG) = (10−50) %				DEG			
298.15	not given	$\delta T = \pm 0.1$ K	61	298.15	1.93−1980	$\delta T = \pm 0.1$ K	59
		$\delta P = \pm 0.15$ kPa		323.15	3.42−3500	$\delta P/P = \pm 0.001$	
				348.15	4.91−5610	$\delta x_{H_2S} = \pm (2-3)$ %	
				373.15	6.07−7480		
				398.15	7.4−6880		
Aqueous $w(\text{PEG}_{400})$ = (10−50) %				TEG			
298.15	not given	$\delta T = \pm 0.1$ K	61	298.15	3.73−1958	$\delta T = \pm 0.5$ K	56
		$\delta P = \pm 0.15$ kPa		323.15	6.57−3452	$\delta P/P = \pm 0.001$	
				348.15	5.33−5480		
				398.15	6.64−6540		
Poly(glycol ethers)^b				PC			
288.15−373.15^c	$P_{H_2S} = 3-100$	$\delta P/P = \pm 0.02$	63	298.15	81.7−1292.4	$\delta T = \pm 0.02$ K for $T < 343$ K	66
				323.15	96.5−1446.1	$\delta T = \pm 0.5$ K at 373 K	
				373.15	233.1−1595.6	$\delta P = \pm 3.5$ kPa	
						$\delta x_{H_2S} = \pm 2\%$	
NMP				NOP			
298.15	211.3−1186.6	$\delta T = \pm 0.02$ K for $T < 343$ K	66	306.70	106.2−1109.9	$\delta T = \pm 0.02$ K	33
323.15	183.6−1384.8	$\delta T = \pm 0.5$ K at 373 K		323.60	127.1−1044.4	$\delta P = \pm 0.2$ kPa	
373.15	175.2−1558.6	$\delta P = \pm 3.5$ kPa				$\delta x_{H_2S} \pm 0.0001$	
		$\delta x_{H_2S} = \pm 2\%$					
TMS							
303.15	55.2−1379.1	$\delta T = \pm 0.02$ K for $T < 343$ K	66	298.15	6.56−1997	$\delta T = \pm 0.5$ K	73
323.15	76−1375.7	$\delta T = \pm 0.5$ K at 373 K		313.15	1.61−2855	$\delta P/P = \pm 0.001$	
373.15	241.4−1654.6	$\delta P = \pm 3.5$ kPa		343.15	1.46−5130		
		$\delta x_{H_2S} = \pm 2\%$		373.15	1.98−5890		
				403.15	2.05−5820		
313.15	$P_{H_2S} = 123-2090$	$\delta x_{H_2S} = \pm (2-3)\%$	72				
373.15	$P_{H_2S} = 348-2350$						
Aqueous ≈(2 and 4) mol·kg^{-1} PIPH_2							
313.14	189.2−2862	$\delta T = \pm 0.1$ K	104				
333.16	158.6−4330	$\delta P/P = \pm 0.001$					
353.18	136.3−6248						
373.14	154.3−8748						
393.16	250.2−8721						

Table 1 (Continued)

T/K	P/kPa	estd uncertainty	ref	T/K	P/kPa	estd uncertainty	ref
Aqueous (0.038−0.28) mol·L^{-1} CDTA[d]				Aqueous (0.038−0.28) mol·L^{-1} CDTA-Fe(II)[e]			
293.15	$P_{H_2S} = 32.86-76.01$	$\delta T = \pm 0.01$ K	75	293.15	$P_{H_2S} = 35-77.3$	$\delta T = \pm 0.01$ K	75
303.15	$P_{H_2S} = 38.76-76.09$	$\delta P/P = \pm 0.0025$		303.15	$P_{H_2S} = 35.1-81.8$	$\delta P/P = \pm 0.0025$	
313.15	$P_{H_2S} = 38.82-79.99$	$\delta x_{H_2S} = \pm 1\%$		313.15	$P_{H_2S} = 35.2-82.8$	$\delta x_{H_2S} = \pm 1\ \%$	
323.15	$P_{H_2S} = 39.95-84.13$			323.15	$P_{H_2S} = 38.6-90.5$		
Aqueous (2.5 and 4.95) mol·L^{-1} MEA (w(MEA) = (15 and 30) %)				Aqueous (2.5 and 5) mol·L^{-1} MEA (w(MEA) = (15 and 30) %)			
313.15	$P_{H_2S} = 0.974-435.3$	$\delta\alpha = \pm (2-3)\ \%$	86	298.15	$P_{H_2S} = 0.152-1822$	$\delta T = \pm 0.5$ K	14
333.15	$P_{H_2S} = 0.963-409.7$			313.15	$P_{H_2S} = 0.651-1936$	$\delta\alpha = \pm (3-5)\ \%$	
353.15	$P_{H_2S} = 3.124-415.9$			333.15	$P_{H_2S} = 0.654-2133$		
373.15	$P_{H_2S} = 1.159-367.9$			353.15	$P_{H_2S} = 1.565-2259$		
				373.15	$P_{H_2S} = 4.89-2164$		
				393.15	$P_{H_2S} = 9.65-2317$		
Aqueous 2.5 mol·L^{-1} MEA (w(MEA) = 15 %)							
298.15	$P_{H_2S} = 6.62-1391.87$	$\delta\alpha = \pm (1.5-2.3)\ \%$	85	353.15	$= 2.18-147.8$	$\delta T = \pm 0.1$ K	76
						$\delta\alpha = \pm (4-5)\ \%$	
298.15	$P_{H_2S} = 0.01-7$	not given	87[f]	298.15, 313.15	$P_{H_2S} \sim 0.001-2$	$\delta T = \pm 0.2$ K	88[f]
Aqueous w(DEA) = (20, 35, 50) %							
299.85	$P_{H_2S} = 0.1544-41.21$	$\delta\alpha = \pm (0.9-1.4)\ \%$	91	299.85	$P_{H_2S} = 298.12-1869.14$	$\delta T = \pm 0.04$ K	27[g]
338.75	$P_{H_2S} = 0.8322-61.69$			338.75	$P_{H_2S} = 360.98-1810.68$	$\delta P = \pm 1.4$ kPa	
388.75	$P_{H_2S} = 11.69-75.29$			388.75	$P_{H_2S} = 239.11-1810.89$	$\delta\alpha = \pm (4-7)\ \%$	
Aqueous 2 mol·L^{-1} DEA (w(DEA) = 20 %)							
298.15, 313.15	$P_{H_2S} \sim 0.001-10$	$\delta T = \pm 0.2$ K	88[f]	313	$P_{H_2S} = 0.04-0.50$	not given	94
				323	$P_{H_2S} = 0.03-0.33$		
Aqueous w(DEA) = (10, 20, 30, 40, 50) %[g]				Aqueous w(DEA) = 41.78 %			
298.15, 313.15,	not given	$\delta T = \pm 0.1$ K	61	313.17	6.06−1337.6	$\delta T = \pm 0.03$ K	90
333.15		$\delta P = \pm 0.15$ kPa		373.01	89.06−1008.2	$\delta P = \pm 0.1$ kPa	
						± 0.0001 P	
						$\delta x_{H_2S} = \pm 0.0005$	
Aqueous (2 and 4) mol·L^{-1} TEA (w(TEA) = (30 and 55) %)				Aqueous w(MDEA) = 20 %			
313	$P_{H_2S} = 0.09-6.32$	not given	94	310.9	$P_{H_2S} = 14.33-1355.48$	$\delta\alpha = \pm (1.5-2.3)\%$	85
				338.7	$P_{H_2S} = 13.23-1536.6$		
				388.7	$P_{H_2S} = 34.6-1267.26$		
Aqueous 1 mol·L^{-1} MDEA (w(MDEA) = 11.8 %)				Aqueous 2 mol·L^{-1} MDEA (w(MDEA) = 23.4 %)			
298.15	$P_{H_2S} = 13.65-1278.98$	$\delta\alpha = \pm (1.5-2.3)\ \%$	85	313.15	$P_{H_2S} = 0.52-1600$	$\delta T = \pm 0.5$ K	99
						$\delta\alpha = \pm 5\ \%$	
Aqueous w(MDEA) = (35 and 50) %				Aqueous 8 mol·kg^{-1} (w(MDEA) = 48.8 %))			
313.15	$P_{H_2S} = 0.00183-313$	$\delta T = \pm 0.5$ K	100	313	147.9−2159	$\delta T = \pm 0.1$ K	101
373.15	$P_{H_2S} = 0.551-301.7$	$\delta P/P = \pm 0.001$		353	344.2−2783	$\delta P/P = \pm 0.001$	
		$\delta\alpha = \pm (2-3)\ \%$		393	351.5−2678		
Aqueous 2.57 mol·L^{-1} MDEA (w(MDEA) = 30 %)				Aqueous ≈(2−4) mol·kg^{-1} MDEA (w(MDEA) = (19.3−32.3) %)			
313.15	$P_{H_2S} = 3.331-445.7$	$\delta\alpha = \pm (2-3)\ \%$	86	313	165.2−2304.4	$\delta T = \pm 0.1$ K	26
333.15	$P_{H_2S} = 3.767-332.3$			333	183.4−2929.2	$\delta P/P = \pm 0.01$	
353.15	$P_{H_2S} = 2.161-426.5$			373	287.7−4085.4		
373.15	$P_{H_2S} = 1.498-348$			393	410.5−4895.9		
				413	904.2−4253.1		
Aqueous w(MDEA) = (11.83, 23.63) %				Aqueous w(MDEA) = 46.78%			
298	$P_{H_2S} = 0.023-1.468$	$\delta\alpha < 0.05$	102	313.16	6.21−1040	$\delta T = \pm 0.03$ K	90
313	$P_{H_2S} = 0.04-1.611$			373.01	90.34−865.41	$\delta P/P = \pm 0.1$ kPa ± 0.0001 P	
						$\delta x_{H_2S} = \pm 0.0005$	
Aqueous w(MDEA) = (10, 20, 30, 40, 50) %[g]				Aqueous w(MDEA) = 30 %[f]			
298.15, 313.15,	not given	$\delta T = \pm 0.1$ K	61	313.15	$P_{H_2S} \sim 0.3-65$	$\delta T = \pm 0.1$ K	103
333.15		$\delta P = \pm 0.15$ kPa				$\delta P = 10$ Pa	
Aqueous w(DGA) = 40 %				Aqueous w(DGA) = 60 %			
313.15	$P_{H_2S} = 1.31-1752.1$	$\delta\alpha = \pm (1.5-2.3)\ \%$	85	323.15	$P_{H_2S} = 11.1-1701.79$	$\delta\alpha = \pm (1.5-2.3)\ \%$	85
333.15	$P_{H_2S} = 15.51-1374.02$			353.15	$P_{H_2S} = 13.16-1642.18$		
Aqueous w(TBAE) = (20−50) %							
298.15, 306.15,	101	$\delta T = \pm 0.1$ K	110				
313.15		$\delta P = \pm 0.14$ kPa					
Aqueous 2 mol·L^{-1} AMP (w(AMP) = 18%))							
313.15	$P_{H_2S} = 2.69-2160$	$\delta T = \pm 0.5$ K	111	343.15	$P_{H_2S} = 4.9-1874$	$\delta T = \pm 0.5$ K	113
373.15	$P_{H_2S} = 2.26-2010$	$\delta\alpha = \pm (2-3)\ \%$				$\delta\alpha = \pm (2-3)\ \%$	
313.15	$P_{H_2S} = 2.69-178$	$\delta T = \pm 0.1$ K	76				
		$\delta\alpha = \pm (4-5)\ \%$					
Aqueous w(AMP) = 30 %				Aqueous 3.43 mol·L^{-1} AMP (w(AMP) = 32 %)			
313.15	$P_{H_2S} = 1.22-133.5$	$\delta\alpha = \pm (2-3)\ \%$	114	323.15	$P_{H_2S} = 4.98-1858$	$\delta T = \pm 0.5$ K	112
333.15	$P_{H_2S} = 2.01-116.2$					$\delta\alpha = \pm (2-3)\ \%$	
353.15	$P_{H_2S} = 2.86-149.2$						
373.15	$P_{H_2S} = 4.88-116.1$						

Table 1 (Continued)

T/K	P/kPa	estd uncertainty	ref	T/K	P/kPa	estd uncertainty	ref
Mixed MEA + TMS							
303.15	$P_{H_2S} = 14.5-1178.2$	$\delta T = \pm 0.02$ K	117[i]	303.15	$P_{H_2S} = 9.3-1246.2$	$\delta T = \pm 0.02$ K	118[j]
		$\delta P = \pm 3.5$ kPa		323.15	$P_{H_2S} = 16.3-1316.1$	$\delta P = \pm 1$ kPa for $P < 20$ kPa	
				373.15	$P_{H_2S} = 59.1-1390.6$	$\delta P = \pm 3.5$ kPa for $P > 20$ kPa	
Mixed Aqueous 2-PE + TMS[h]				Mixed MEA + NMP[k]			
313.15	$P_{H_2S} = 0.0284-2410$	$\delta T = \pm 0.1$ K	115	298.15	$P_{H_2S} = 28.1-1085.5$	$\delta T = \pm 0.02$ K	116
373.15	$P_{H_2S} = 0.254-5550$	$\delta\alpha = \pm 4\%$		323.15	$P_{H_2S} = 20.1-1301.3$	$\delta P = \pm 3.5$ kPa	
				373.15	$P_{H_2S} = 46.5-993.9$		
Mixed DEA + TMS							
303.15	$P_{H_2S} = 26.2-1214.5$	$\delta T = \pm 0.02$ K	117[l]	303.15	$P_{H_2S} = 14.3-1225.8$	$\delta T = \pm 0.02$ K	119[m]
		$\delta P = \pm 3.5$ kPa		323.15	$P_{H_2S} = 24.6-1374.8$	$\delta P = \pm 3.5$ kPa	
				373.15	$P_{H_2S} = 53.3-1439.7$	$\delta x_{H_2S} = \pm 2\ \%$	
Mixed DEA + NMP[n]				Aqueous MDEA + H_2SO_4[o]			
298.15	$P_{H_2S} = 73.2-1035.3$	$\delta T = \pm 0.02$ K	116	313	95.8−2708	$\delta T = \pm 0.1$ K	105
323.15	$P_{H_2S} = 29.5-934.5$	$\delta P = \pm 3.5$ kPa		353	252.2−3158	$\delta P/P = \pm 0.001$	
373.15	$P_{H_2S} = 81.2-1274.9$			393	356.1−3656		
Aqueous MDEA + Na_2SO_4[p]				Mixed Aqueous MDEA + $PIPH_2$[q]			
313	32.7−2806	$\delta T = \pm 0.1$ K	105	353.14	136.4−6207	$\delta T = \pm 0.1$ K	104
353	173.6−3866	$\delta P/P = \pm 0.001$				$\delta P/P = \pm 0.001$	
393	265.8−3769						
Mixed Aqueous MDEA + TMS[r]				Mixed Aqueous and Nonaqueous MDEA + EG[s,f]			
313.15	$P_{H_2S} = 1.3-1470$	$\delta T = \pm 0.5$ K	99	298.15, 313.15, 333.15, 333.15	$P_{H_2S} = 0.34-38.8$	$\delta T = \pm 0.1$ K	103
373.15	$P_{H_2S} = 1.58-3210$	$\delta\alpha = \pm 5\%$				$\delta P = \pm 10$ Pa	
Mixed TEA + PC[t,f]				Mixed Aqueous AMP + TMS[u]			
283.15, 313.15	$P_{H_2S} \sim 2.5-110$	$\delta T = \pm 0.1$ K	121	313.15	$P_{H_2S} = 2.45-1610$	$\delta T = \pm 0.5$ K	120
		$\delta P = \pm 0.01$ bar		373.15	$P_{H_2S} = 4.54-2200$	$\delta\alpha = \pm (2-3)\%$	
		$\delta x_{H_2S} = \pm 1 \times 10^{-6}$ (mass fraction)					
Mixed Aqueous MDEA + MEA[v]				Mixed Aqueous MDEA + DEA[w]			
313.15	$P_{H_2S} = 1.309-444.6$	$\delta\alpha = \pm (2-3)\%$	86	313.16	6.88−1134.7	$\delta T = \pm 0.03$ K	90
333.15	$P_{H_2S} = 1.399-442.3$			373.01	93.62−931.98	$\delta P = \pm 0.1$ kPa ± 0.0001 P	
353.15	$P_{H_2S} = 1.307-398.9$					$\delta x_{H_2S} = \pm 0.0005$	
373.15	$P_{H_2S} = 1.285-379.4$						
Mixed Aqueous MEA + AMP[x]							
313.15	$P_{H_2S} = 0.53-181.6$	$\delta\alpha = \pm (2-3)\ \%$	114				
333.15	$P_{H_2S} = 1.08-159$						
353.15	$P_{H_2S} = 1.33-138.4$						
373.15	$P_{H_2S} = 0.83-99.1$						

[a] Gas mole numbers. [b] Diglyme, diethylene glycol dimethyl ether; triglyme, triethylene glycol dimethyl ether; tetraglyme, tetraethylene glycol dimethyl ether; dowanol DM, diethylene glycol methyl ether; dowanol TBH, triethylene glycol butyl ether. [c] Only data at 298.15 are tabulated. [d] *trans*-1,2-Cyclohexanediaminetetraacetic acid. [e] Fe(II) chelate complex of *trans*-1,2-cyclohexanediaminetetraacetic acid. [f] Experimental data are not tabulated. [g] Physical solubility. [h] w(2-PE) = 45 % (2-piperidineethanol), w(TMS) = 40 %. [i] w(MEA) = 15 %. [j] w(MEA) = (15 and 30) %. [k] w(MEA) = 15 %. [l] w(DEA) = 15 %. [m] w(DEA) = (15, 30, and 50) %. [n] w(DEA) = 30 %. [o] 3.992 $\mathrm{mol \cdot kg^{-1}}$ MDEA, 0.9862 $\mathrm{mol \cdot kg^{-1}}$ H_2SO_4. [p] 1.900 $\mathrm{mol \cdot kg^{-1}}$ MDEA, 0.9566 $\mathrm{mol \cdot kg^{-1}}$ Na_2SO_4. [q] 1.975 $\mathrm{mol \cdot kg^{-1}}$ MDEA, 1.966 $\mathrm{mol \cdot kg^{-1}}$ $PIPH_2$. [r] w(MDEA) = 20.9 % (2 M), w(TMS) = 30.5 %. [s] w(MDEA) = 30 %, w(EG) = 70 %; w(MDEA) = 30 %, w(EG) = 65 %; w(MDEA) = 30 %, w(EG) = 60 %. [t] w(TEA) = (0, 2, 5, and 10) %. [u] w(AMP) = 16.5 % (2 $\mathrm{mol \cdot L^{-1}}$ at 23 °C), w(TMS) = 32.2 %. [v] 0.51 $\mathrm{mol \cdot L^{-1}}$ MDEA, 3.97 $\mathrm{mol \cdot L^{-1}}$ MEA; 1.54 $\mathrm{mol \cdot L^{-1}}$ MDEA, 2 $\mathrm{mol \cdot L^{-1}}$ MEA. [w] w(MDEA) = 37.73 %, w(DEA) = 7.64 %. [x] w(MEA) = 6 %, w(AMP) = 24 %; w(MEA) = 12 %, w(AMP) = 18 %; w(MEA) = 18 %, w(AMP) = 12 %.

represented along with two correlated curves: (i) based on selected experimental data reported in three references {from (273 to 563) K}, considered as “optimum values”[2,15−17] and (ii) based on experimental measurements made at temperatures from (293 to 593) K.[18] Henry’s constants (K_H) from Chapoy et al.[3] were calculated using the following modified K_H-T correlation:[19] $\log(K_H/\mathrm{MPa}) = 84.44 + 0.0101845T(\mathrm{K}) - 3.792 \times 10^3/T(\mathrm{K}) - 29.5008 \log(T(\mathrm{K}))$. It is important to mention that despite the great interest for this system, there are only few measurements made at high temperatures.[17,18,20−22] We just note that some measurements at high temperatures were given by Drummond[23] in his Ph.D. thesis; the data have not been published in the open literature nor included in the solubility data series published by Fogg and Young.[1] Fernández-Prini et al.[24] developed a correlation for the Henry’s constants, based on four data sources that cover a wide range of temperatures from (273 to 588) K.[2,17,20,22] They analyzed the available data for this system and discussed the reasons why the experimental data reported by Suleimenov and Krupp[18] and Kozintseva[21] have not been included in their correlations. On the basis of a critical examination of the existing experimental data, the authors considered that data of high accuracy extended only up to about 450 K, and they mentioned therefore that their correlated values above this temperature should be considered tentative until they are confirmed in new experimental works. The corresponding correlation curve included in Figure 1, based on parameters that covered the temperature range between (273.15 and 533.09) K, was however extrapolated beyond the range of the fitted data up to 600 K.[24] Taking into account that measurements at so high temperatures are generally very rare, data by Kozintseva[21] were also represented in Figure 1 for comparison.

Table 2. Methyl Mercaptan Solubility Data from Literature

T/K	P/kPa	estd uncertainty	ref
Water			
273.15, 293.15, 313.15, 333.15[a]	not given	$\delta T = \pm 0.5$ K	123
333.15	101	$\delta K_H = \pm 8$ %	124
303.15	$P_{MM} = 101$	not given	125
298.15, 308.15, 318.15, 343.15[b]	101	$\delta T = \pm 0.1$ K	126
310.9	207−13790	$\delta T = \pm 0.01$ K	127
366.5	689−13790	$\delta P/P = \pm 0.003$	
422	1724−13790	$\delta x_{MM} = \pm 1$ %	
469.9	3447−20680		
533.2	6895−20680		
588.7	13790−20680		
273.15, 298.15, 323.15	101	$\delta K_H = \pm 3$ %	128
323	77.299−232.467	$\delta T = \pm 0.003$ K	132
353	147.85−384.138	$\delta P/P = \pm 0.001$	
373	220.872−501.727		
298.15	$P_{MM} = 108$	$\delta P/P = \pm 0.001$	130
323.15	$P_{MM} = 101.3$		
353.15	$P_{MM} = 7.56$		
373.15	$P_{MM} = 7.02$		
298.15, 303.15, 308.15, 313.15, 318.15, 323.15, 328.15, 333.15	101	$\delta T = \pm 0.01$ K $\delta K_H = \pm 2$ %	129
Aqueous 1 mol·L^{-1} NaCl, 1 mol·L^{-1} Na_2SO_4, 0.1 mol·L^{-1} $NaSCH_3$, (0.05 and 1) mol·L^{-1} H_2SO_4			
303 K	$P_{MM} = 101$	not given	125
Seawater[b,c]			
298.15, 308.15, 318.15, 343.15	101	$\delta T = \pm 0.1$ K	126
Hexane			
323	90.741−157.507	$\delta T = \pm 0.003$ K	132
353	191.678−312.754	$\delta P/P = \pm 0.001$	
373	305.197−466.625		
Toluene			
323	29.024−723.60	$\delta T = \pm 0.003$ K	132
353	69.879−151.411	$\delta P/P = \pm 0.001$	
373	116.115−232.874		
Aqueous (0.038−0.28) mol·L^{-1} CDTA[d]			
298.15, 303.15, 308.15, 313.15, 318.15, 323.15, 328.15, 333.15	101	$\delta T = \pm 0.01$ K $\delta K_H = \pm 2$ %	129
Aqueous (0.038−0.28) mol·L^{-1} CDTA-Fe(II)[e]			
298.15, 303.15, 308.15, 313.15, 318.15, 323.15, 328.15, 333.15	101	$\delta T = \pm 0.01$ K $\delta K_H = \pm 2$ %	129
Aqueous w(MDEA) = 50 %			
313.15	$P_{MM} = 0.208−327$	$\delta P/P = \pm 0.001$	130
343.15	$P_{MM} = 0.268−425$		
Aqueous w(DEA) = 35 %			
313.15	$P_{MM} = 0.105−6.12$	$\delta T = \pm 0.1$ %	134
343.15	$P_{MM} = 0.307−12.5$	$\delta P/P = \pm 0.001$ $\delta x_{MM} = \pm (5−7)$ %	

[a] Experimental data are not tabulated. [b] Calculated on the basis of the parameters of the linear regression curve ($\log K_H = a + b/T$) using experimental data on a temperature range (298.15 to 343.15) K (primary experimental data are not given). [c] Ionic strength of an aqueous sodium sulfate solution of (0.7 to 4) (the units are not given). [d] *trans*-1,2-Cyclohexanediaminetetraacetic acid. [e] Fe(II) chelate complex of *trans*-1,2-cyclohexanediaminetetraacetic acid.

Table 3. Dimethyl Sulfide Solubility Data from Literature

T/K	P/kPa	estd uncertainty	ref
Water			
298.15	101	not given	135
273.15, 283.15, 293.15, 313.15, 333.15, 353.15[a]	not given	$\delta T = \pm 0.5$ K	123
293.15	101	$\delta T = \pm 0.1$ K	136
298.15, 308.15, 318.15, 343.15[b]	101	$\delta T = \pm 0.1$ K	126
272.35−305.55	101	$\delta T = \pm 0.5$ K $\delta K_H = \pm 2$ %	12
273.15, 298.15, 323.15	101	$\delta K_H = \pm 3$ %	128
291.15	101	$\delta K_H = \pm 5.1$ %	139
288.15, 293.15, 298.15, 303.15, 308.15	101	$\delta T = \pm 0.01$ K $\delta K_H = \pm 2$ %	142
292.6−332.8	101.3	$\delta T = \pm 0.1$ K $\delta K_H = \pm 5$ %	143
Seawater			
not given	(101)	not given	138
Seawater[c]			
298.15, 308.15, 318.15, 343.15[b]	101	$\delta T = \pm 0.1$ K	126
Seawater[d]			
300.15	(101)	not given	137
Seawater[e]			
273−302.25	101	$\delta T = \pm 0.5$ K $\delta K_H = \pm 2$ %	12
Seawater			
291.15, 298.15, 308.15, 317.15	101	$\delta K_H = \pm 3.6$ %	139
Aqueous w(NaCl) = (2−32) %			
291.15	101	$\delta K_H = \pm 5$ %	139
Aqueous (0.038−0.30) mol·L^{-1} CDTA[f]			
288.15, 293.15, 298.15, 303.15, 308.15	101	$\delta T = \pm 0.01$ K $\delta K_H = \pm 2$ %	142
Aqueous (0.038−0.10) mol·L^{-1} CDTA-Fe(II)[g]			
288.15, 293.15, 298.15, 303.15, 308.15	101	$\delta T = \pm 0.01$ K $\delta K_H = \pm 2$ %	142
Aqueous w(MDEA) = 50 %			
293.2−342.7	101.3	$\delta T = \pm 0.1$ K $\delta K_H = \pm 5$ %	143

[a] Experimental data are not tabulated. [b] Calculated on the basis of the parameters of the linear regression curve ($\log K_H = a + b/T$) using experimental data on a temperature range (298.15 to 343.15) K (primary experimental data are not given). [c] Ionic strength of an aqueous sodium sulfate solution of (0.7 to 4) (the units are not given). [d] 19×10^{-6} μg/g Cl^- (solubility data are given in this paper, but the original manuscript containing the measurements was not published). [e] (0 to 19.1) $\times 10^{-6}$ μg/g Cl^-. [f] *trans*-1,2-Cyclohexanediaminetetraacetic acid. [g] Fe(II) chelate complex of *trans*-1,2-cyclohexanediaminetetraacetic acid.

The agreement between the experimental data presented in Figure 1 is rather good at temperatures up to about 360 K. For a better comparison, the hydrogen sulfide solubility expressed by the Henry's law constants for several temperatures is given in Table 5 for pressures near and below atmospheric where the system obeys the strict Henry's law.[2] The experimental values can be compared with the correlated ones given by Carroll and Mather,[2] based on selected reliable experimental data from (273 to 363) K and pressures up to 1 MPa.

Data by Suleimenov and Krupp[18] agree well with those given by Kozintseva[21] up to about 500 K. The authors used a titanium pressure vessel to measure the equilibrium vapor pressures, and the H_2S concentration was determined by iodometric titration. Between (360 and 500) K, only the values at 367 K given by Barrett et al.[25] largely deviate from the general trend. Barrett et al.[25] used a static analytical method for measuring data at atmospheric pressure. At larger temperatures, however, the

Table 4. Dimethyl Disulfide Solubility Data from Literature

T/K	P/kPa	estd uncertainty	ref
	Water		
273.15, 283.15, 293.15, 313.15, 333.15[a]	not given	$\delta T = \pm\ 0.5$ K	123
293.15	101	$\delta T = \pm\ 0.1$ K	136
298.15, 308.15, 318.15, 343.15[b]	101	$\delta T = \pm\ 0.1$ K	126
273.15, 298.15, 323.15	101	$\delta K_H = \pm\ 3$ %	128
298.15, 303.15, 308.15, 313.15, 318.15, 323.15, 328.15, 333.15, 338.15	101	$\delta T = \pm\ 0.01$ K $\delta K_H = \pm\ 2$ %	144
	Seawater[c]		
298.15, 308.15, 318.15, 343.15[b]	101	$\delta T = \pm 0.1$ K	126
	Aqueous (0.038−0.30) mol·L^{-1} CDTA[d]		
298.15, 303.15, 308.15, 313.15, 318.15, 323.15, 328.15, 333.15	101	$\delta T = \pm\ 0.01$ K $\delta K_H = \pm\ 2$ %	144
	Aqueous (0.038−0.30) mol·L^{-1} CDTA-Fe(III)[e]		
298.15, 303.15, 308.15, 313.15, 318.15, 323.15, 328.15, 333.15	101	$\delta T = \pm\ 0.01$ K $\delta K_H = \pm\ 2$ %	144

[a] Experimental data are not tabulated. [b] Calculated on the basis of the parameters of the linear regression curve (log $K_H = a + b/T$) using experimental data on a temperature range (298.15 to 343.15) K (primary experimental data are not given). [c] Ionic strength of an aqueous sodium sulfate solution of (0.7 to 4) (the units are not given). [d] *trans*-1,2-Cyclohexanediaminetetraacetic acid. [e] Fe(III) chelate complex of *trans*-1,2-cyclohexanediaminetetraacetic acid.

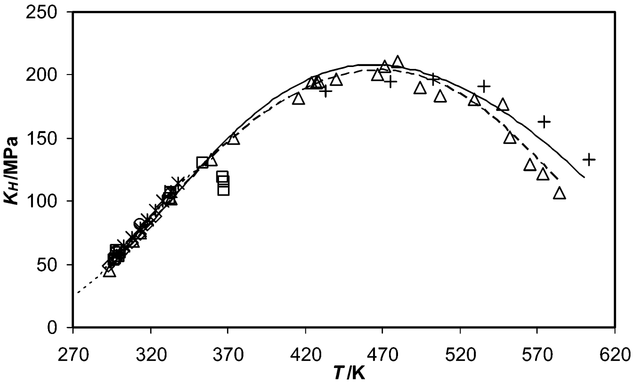

Figure 1. Henry's constant for H_2S in water: +, ref 21; △, ref 18; □, ref 25; ×, ref 110; ○, ref 61; ◇, ref 75; ∗, ref 3; - - -, corr (ref 2); — —, corr (ref 18); —, corr (ref 24).

Table 5. Henry's Law Coefficient K_H for Hydrogen Sulfide in Water at near and below Atmospheric Pressures

T/K	ref 25[a]	ref 18[b]	ref 110	ref 61	ref 75	corr (ref 2)
			K_H/MPa			
298.15	60.49	52.10	59.99	59.49	54.18	54.71
303.15		58.92			60.67	61.51
308.15		65.77			67.55	68.59
313.15		72.64	77.16	81.64	73.96	75.86
318.15		79.51			81.71	83.22
323.15		86.34			88.11	90.56
333.15	106.76	99.86		106.08		104.77
343.15		113.03				117.54
353.15		125.73				127.98

[a] Calculated from H_2S molality. [b] Estimated by interpolation from the experimental data.

Henry's constants reported by Kozintseva[21] become much larger than those given by Suleimenov and Krupp.[18] The latter authors claim the correctness of their data based especially on some inconsistencies of data by Kozintseva[21] as well as on experimental difficulties assumed encountered during her measurements. Moreover, Henry's constants given by Drummond[23] are even larger than those published by Kozintseva[21] and Suleimenov and Krupp[18] in this high-temperature range. It is interesting to note that the correlative values given by Fernández-Prini et al.[24] in the high-temperature range lie between those by Kozintseva[21] and Suleimenov and Krupp,[18] but they are in general much closer to those reported by Kozintseva.[21] New experimental data at high temperatures would therefore be very useful to elucidate the real behavior. In addition, data by Kuranov et al.[26] obtained at a constant temperature and pressures up to about 25 MPa agree well with literature data published before 1987.[1] H_2S solubility data measured by Maddox et al.[27] at 299.85 K agree very well with those reported by Lee and Mather.[17]

Several workers reported data concerning the H_2S solubility in different salt solutions. Because there is no similitude among data presented in Table 1, no comparison is possible. However, when applicable, the reliability of these data can be judged in accordance with the corresponding measurements in pure water.

The solubility data of hydrogen sulfide in sodium chloride solutions given by Barrett et al.[25] can be considered reliable up to about 358 K. The values of the solubility of hydrogen sulfide in a NaCl solution of concentration (2 and 3) mol·kg^{-1} at 298 K agree well with those given by Gamsjäger and Schindler[28] that are considered more reliable with respect to other data.[1] However, based on the results of the gas solubility in water, at temperatures larger than 360 K the solubility values might be questionable. Based on the discussion concerning the H_2S solubility in pure water, the data given by Suleimenov and Krupp[18] seem to be the only reliable measurements of gas solubility in the presence of NaCl made at temperatures higher than 400 K. New experimental data would, however, be very welcome. Xia et al.[29−31] used a static synthetic method for determining the hydrogen sulfide solubility in electrolyte solutions containing NaCl, $(NH_4)_2SO_4$, Na_2SO_4, NH_4Cl, $NaNO_3$, NH_4NO_3, NaOH, CH_3COONa, and CH_3COONH_4 over a very large pressure range (up to 10 MPa).

Leu et al.[32] studied the behavior of methanol + hydrogen sulfide, a system of interest in the petroleum and natural gas industry, using a direct static analytical method. Methanol can be generally used as solvent for removing hydrogen sulfide in various processes involving the sour natural gas or industrial synthesis gas. Data by Leu et al.[32] were measured between (298 and 448) K and pressures between the methanol vapor pressure to pressures in the critical region along each isotherm. No previous work reported the vapor compositions. Their data agree well with those measured by Fischer et al.[33] at 298 K using a static synthetic method. However, they found an inconsistency with the lower temperature, (248 to 273) K, data of Yorizane et al.[34]

The binary and multi-component systems containing hydrocarbons and hydrogen sulfide are of interest to the petroleum and natural gas industry. Therefore, VLE data for these systems are especially important in phase equilibria modeling.

Using the direct static analytical method, Leu and Robinson[35] studied the high-pressure VLE of isopentane (or neopentane)

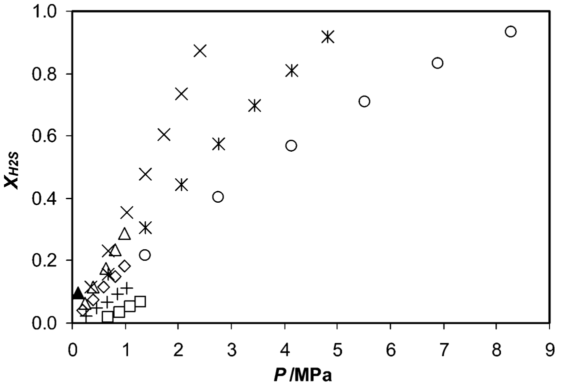

Figure 2. Mole fraction solubility of H_2S in decane: ▲, 267 K (ref 44); ×, 311 K (ref 45); △, 323 K (ref 43); ∗, 344 K (ref 45); ◇, 373 K (ref 43); ○, 377 K (ref 45); +, 423 K (ref 43); □, 523 K (ref 43).

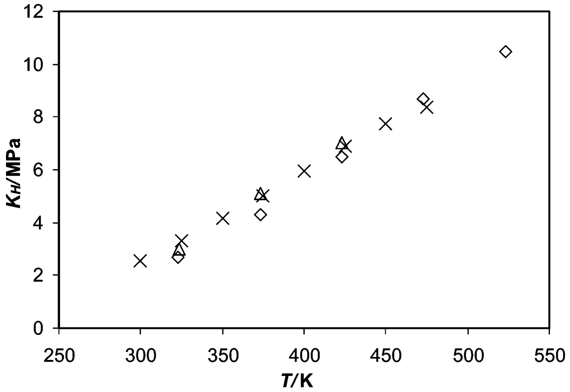

Figure 3. Henry's law constants for H_2S in *n*-hexadecane: ×, ref 48; ◇, ref 46; △, ref 43.

+ hydrogen sulfide over a large temperature range and pressures from the vapor pressure of the pentanes to the critical pressure for the binary system at each temperature. No similar data are available in the literature for this system. The only data available for the binary C_5-alkane + H_2S system concern the system containing pentane and were published by Reamer et al.[36] and Makranczy et al.[37]

Laugier and Richon[38] measured the VLE for the system hexane + hydrogen sulfide at temperatures between (323 and 423) K using a direct static analytic method with sampling of both phases. Previous measurements are given at atmospheric pressure and did not exceed 303 K.[37,39−41]

Phase equilibria between cyclohexane and hydrogen sulfide at high pressures have also been investigated by Laugier and Richon[38] between (323 and 423) K. Previous measurements were carried out at low pressures (up to about 1 bar) and temperatures from (283 to 313) K.[39,42] The corresponding data at 293 K given in these two references are significantly different.

The isooctane + hydrogen sulfide system was investigated by Yokoyama et al.[43] using a static-type apparatus with sampling of liquid phase. On the basis of the high pressure solubility data measured at temperatures between (323 and 473) K, the Henry's law constants were determined. No similar data are available in the literature for this system.

The same authors studied the hydrogen sulfide solubility in other three heavier straight-chain alkanes (decane, tridecane, and hexadecane) and in squalane and determined the Henry's law constants. There are no other available data for the system squalane + hydrogen sulfide.

Previous measurements concerning the system with decane were performed at atmospheric pressure by Gerrard[44] at (267 and 273) K, by King and Al-Najjar[40] between (288 and 343) K, by Makranczy et al.[37] at (298 and 313) K, and by Reamer et al.[45] at very high pressures up to 12411 kPa and temperatures between (278 and 444) K. The pressure influence on the mole fraction solubility in the available temperature range of (267 to 523) K is given in Figure 2.

Previous measurements concerning the system with tridecane include only the data by Makranczy et al.[37] The authors reported mole fraction solubilities at (298 and 313) K at atmospheric pressure.

The system with hexadecane has also been investigated by Feng and Mather[46] between (323 and 423) K and pressures up to 7.4 MPa and at atmospheric pressure by Bell[39] at 293 K, by Makranczy et al.[37] at (298 and 313) K, by King and Al-Najjar[40] between (288 and 343) K, by Lenoir et al.[47] at 298 K, and by Tremper and Prausnitz[48] at different pressures between 300 and 475 K. The inconsistency between Henry's law constants given by Yokoyama et al.[43] and those by Tremper and Prausnitz[48] become larger at temperatures lower than about 350 K and larger than about 450 K (Figure 3). The same disagreement was reported by Feng and Mather;[46] the values are higher at lower temperature and vice versa. In addition, the Henry's law constants by Feng and Mather[46] are between (8 and 18) % greater than those reported by Yokoyama et al.[43] at (323, 373, and 423) K.

The dodecane + hydrogen sulfide system was studied by Feng and Mather[49] using an equilibrium cell consisting of a Jerguson liquid level gauge with vapor-phase recirculation.[50] The Henry's law constants agree well with the low-pressure data previously published for this system {i.e., by Bell[39] at 293 K, by Makranczy et al.[37] at (298 and 313) K, and by King and Al-Najjar[40] between (288 and 343) K}.

Laugier and Richon[38] investigated the binary system pentadecane + hydrogen sulfide at 423 K and the ternary pentadecane + hexane + hydrogen sulfide system at 425 K. Previous data are from Makranczy et al.,[37] who measured the solubility of hydrogen sulfide in pentadecane at (298 and 313) K and atmospheric pressure.

The solubility of hydrogen sulfide in *n*-eicosane was measured between (323 and 423) K by Feng and Mather.[51] Later, the same research group investigated the gas solubility in its mixture with *n*-hexadecane at 323 K.[52] No other data are available in the literature for these systems.

The system benzene + hydrogen sulfide was studied by Laugier and Richon[38] between (323 and 423) K. Their data are in good agreement with those given by Fischer et al.[33] at 323 K. Experimental data by Fischer et al.[33] are also consistent with previous literature data for this system,[39,44] except for one value at 298 K measured by Patyi et al.[53]

Richon et al.[54] reported the VLE data for the system *n*-propylbenzene + hydrogen sulfide at high pressures and temperatures between (313 and 473) K. Patyi et al.[53] also measured the hydrogen sulfide solubility at 298 K; however, as in the case of the benzene + hydrogen sulfide system, these data seem to be questionable.

Ethylene glycol (EG), diethylene glycol (DEG), and triethylene glycol (TEG) have been commonly used for dehydration of gas streams containing the acid gases (H_2S and CO_2) inclusively in the processing of natural gas and enhanced oil recovery. The presence of these sour gases causes the solution to be corrosive, especially at the high temperatures of the regenerator, which leads to the necessity of the study of their solubility in these solvents.

The solubility of hydrogen sulfide in ethylene glycol was studied by Jou et al.[55] between (298 and 398) K using a similar apparatus to that described by Jou et al.[56] The authors calculated the Henry's law constants and compared them with those obtained in previous works. They found a good agreement with the high-temperature values reported by Short et al.[57] that are also in good agreement with those by Gerrard.[44] However, Henry's law constants given by Lenoir et al.[47] and Byeseda et al.[58] were found to be much lower.

The solubility of hydrogen sulfide in DEG was studied by Jou et al.[59] using a windowed Jerguson cell with vapor-phase recirculation similar to that designed by Jou et al.[50] between (298 and 398) K, and the Henry's law constants were calculated. No similar data are available in the open literature for comparison.

Jou et al.[56] investigated the system triethylene glycol + hydrogen sulfide between (298 and 398) K using the same apparatus[50] and calculated the Henry's law constants. Previous measurements concerning this system include the data by Byeseda et al.[58] at 297 K and graphical data by Blake[60] between (273 and 373) K. Data reported by Jou et al.[56] agree well with those given by Blake[60] at low temperatures, up to 323 K, but are smaller at higher temperatures.

Rinker and Sandall[61] measured the solubility of hydrogen sulfide in aqueous solutions containing ethylene glycol (EG) or polyethylene glycol (PEG 400) using a modified Zipperclave reactor. The Henry's law constants are very useful for modeling the absorption or stripping of the gas in these solvents. Solubility data for polyethylene glycols including PEG 200, PEG 300, PEG 400, and PEG 1000 have also been reported by Gestrich and Reinke[62] at temperatures from (343 to 423) K and sub-atmospheric pressures.

The solubility of hydrogen sulfide in the mono- and dimethyl ethers of poly(ethylene glycol) between (288 and 373) K was investigated by Sciamanna and Lynn[63] as a part of a project concerning the development as an alternative to conventional sulfur recovery technology for removing hydrogen sulfide from gas streams and converting it to elemental sulfur. Experimental data were obtained in an equilibrium cell based on the saturation method. A fair agreement was found between the Henry's law constants in tetraglyme at temperatures up to 323 K reported by these authors and those given by Härtel.[64] Quite large inconsistencies appear at higher temperatures. Sweeney[65] also reported data at (298 and 323) K; the value at 323 K lies between those given by Härtel[64] and those given by Sciamanna and Lynn.[63] No similar data are available in the literature for the other systems studied.

Murrieta-Guevara et al.[66] investigated the solubility of hydrogen sulfide in several physical solvents {propylene carbonate (PC), *N*-methylpyrrolidone (NMP)} and sulfolane {tetramethylene sulfone (TMS)} between (298 and 373) K and high pressures, using a solubility apparatus based on a static method with evaluation of the amount of solute dissolved. The absorption process using physical solvents has the advantage of low-energy requirements for the regeneration step and is often a better choice in treating gas streams at high pressures and concentrations.

Previous measurements concerning the system with propylene carbonate include the data by Shakhova et al.[67] between (273 and 313) K, by Lenoir et al.[47] between (298 and 343) K, by Isaacs et al.[68] at (313 and 373) K, by Sweeney[65] at (298 and 323) K, and by Rivas and Prausnitz,[69] who reported only the Henry's law constants between (263 and 373) K. Good agreement was found at 373 K between the data reported by Murrieta-Guevara et al.[66] and that reported by Isaacs et al.[68] Moreover, as shown by Isaacs et al.[68] and also discussed in Fogg and Young,[1] there is a fairly good agreement between different sets of measurements.

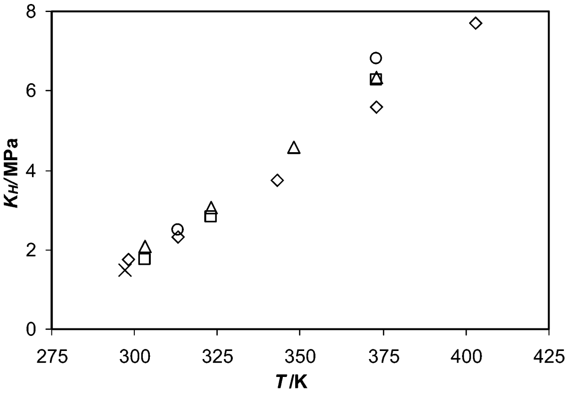

Figure 4. Henry's law constants for H_2S in sulfolane: △, ref 69; ×, ref 58; □, ref 66; ○, ref 72; ◇, ref 73.

The system *N*-methylpyrrolidone + hydrogen sulfide has also been studied by Lenoir et al.[47] at 298 K, by Rivas and Prausnitz[69] between (263 and 373) K, by Yarym-Agaev et al.[70] between (273 and 399) K, by Sweeney[65] at (298 and 323) K, and by Murrieta-Guevara and Trejo Rodriguez[71] at 298 K. The Henry's law constants reported by Murrieta-Guevara et al.[66] agree well generally with those given by Rivas and Prausnitz[69] at (298, 323, and 373) K, by Murrieta-Guevara and Trejo Rodriguez[71] at 298, and by Yarym-Agaev et al.[70] and Sweeney[65] at 323 K. However, a large discrepancy is noted at 298 K with respect to data by Lenoir et al.[47] and Sweeney,[65] who used a chromatographic method. Also, the value of the Henry's law constant at 298 K given by Yarym-Agaev et al.[70] lies between those reported by Murrieta-Guevara et al.[66] and Rivas and Prausnitz[69] and those reported by Lenoir et al.[47] and Sweeney.[65]

The system sulfolane + hydrogen sulfide has also been studied by Roberts and Mather[72] at (313 and 373) K and by Jou et al.[73] between (298 and 403) K. Previous measurements for this system were performed by Rivas and Prausnitz[69] between (303 and 373) K and by Byeseda et al.[58] at 297 K. For comparison, the Henry's law constants are given in Figure 4 at different temperatures. It should be taken into consideration that data by Byeseda et al.[58] and Rivas and Prausnitz[69] were obtained at pressures below and near 1 atm. All other measurements were performed at high and very high pressures (Table 1). The Henry's law constants given by Murrieta-Guevara et al.[66] and Jou et al.[73] were obtained by fitting the data to the Krichevsky−Kasarnovsky and the Krichevsky−Ilinskaya equations, respectively.[74] Murrieta-Guevara et al.[66] mentioned that the use of the Krichevsky−Ilinskaya equation instead led to no significant difference concerning the values of the Henry's law constants.

Fischer et al.[33] investigated the system *N*-octyl-2-pyrrolidone (NOP) + hydrogen sulfide at (306 and 323) K using a static synthetic method. As mentioned by the authors, the replacement of the methyl group by a longer alkyl group allows the reduction of the volatility by keeping a high selectivity. No similar data are available in the literature for this system.

Among the TRS compounds, the hydrogen sulfide is the most abundant component in the Kraft mill effluents. Utilization of Fe(III) chelate complex of *trans*-1,2-cyclohexanediaminetetraacetic acid (CDTA) for the oxidation of hydrogen sulfide is beneficial from the standpoint of iron sequestration and protection against precipitation in the alkaline environments characteristic of the Kraft mill streams. The physical solubility of hydrogen sulfide in CDTA and ferrous CDTA complex (which does not react with H_2S) in the temperature range of (293 to 323) K and at sub-atmospheric pressures was investigated by Iliuta et al.[75] using an apparatus based on the saturation method. No similar data for these systems are available in the literature.

The removal of sour gases, including hydrogen sulfide, from gas streams is important in many processes involved, especially in the natural gas and petroleum industry. The most commonly used solvents are the aqueous solutions of various single or mixed alkanolamines like MEA (monoethanolamine), DEA (diethanolamine), TEA (triethanolamine), MDEA (methyldiethanolamine), and AMP (2-amino-2-methyl-1-propanol). Furthermore, the addition of physical solvents to the chemical solvents can also represent a very interesting option.

MEA is widely used for the removal of acid gases, and a considerable amount of solubility data is available in the literature. Jane and Li[76] measured the solubility of hydrogen sulfide in a 15 wt % aqueous MEA at 353 K using a VLE apparatus with vapor-phase recirculation and found a good agreement between their data and those reported by Lee et al..[14] The last authors used a recirculating vapor flow equilibrium cell with a Jerguson liquid level gauge.[77] They reported solubility data in w(MEA) = (15 and 30) % aqueous solution from (298 to 393) K and showed a good agreement between their data and those measured previously,[78−80] except for those by Riegger et al.[81] Previous data by Lee et al.[82] on the same system at (313 and 373) K were also found to be in good agreement with the literature data except for a H_2S partial pressure around 100 kPa. Some evident discrepancies between data published before 1987 (except for Lee et al.[14]) are revealed in Fogg and Young.[1] Two other previous measurements on this system are also included in this collection.[83,84] Experimental data reported by Maddox et al.[85] for w(MEA) = 15 % at 298.15 K compare favorably with the data of Muhlbauer and Monaghan[79] at low partial pressures and with those given by Lee et al.[14] at high partial pressures. Li and Shen[86] also reported the hydrogen sulfide solubility in w(MEA) = 15 % at 313 K, and in w(MEA) = 30 % from (313 to 373) K. They showed that their data at 313 K (w(MEA) = 15 %) agree generally well with those measured by Lee et al.;[82] some discrepancies were observed at H_2S partial pressure around 1000 kPa. Rochelle et al.[87] and Cheng et al.[88] measured the solubility of hydrogen sulfide at very low pressures at 298 K and at (298 and 313) K, respectively, in a w(MEA) = 15 % aqueous solution using an electrode (pH−silver sulfide) method. In both cases, experimental data are not tabulated. However, the authors compared graphically their data with previous measurements. Rochelle et al.[87] showed a general fair agreement between their data and those by Leibush and Shneerson[78] and Muhlbauer and Monaghan.[79] Data from these last two references were also found to agree generally well with those reported by Jones et al.[80] and Lawson and Garst.[89] Cheng et al.[88] also found a fair agreement between their data at 298 K and those by Rochelle et al.[87] and Lee at al.[14] However, the authors claimed that “the error of their work is smaller compared with the data by Lee et al.”, but it is not clear what they exactly meant; this is valid for the other systems investigated including the DEA + hydrogen sulfide.

Rinker and Sandall[61] measured the physical solubility of hydrogen sulfide in aqueous solutions of diethanolamine (DEA) or methyldiethanolamine (MDEA) using a modified Zipperclave reactor. Because the gas reacted with these amines, the solutions were neutralized by the addition of hydrochloric acid. The authors report the Henry’s law constants without giving the values of the total or partial pressures. The physical solubility of hydrogen sulfide in protonated DEA aqueous solutions was also measured by Maddox et al.[27] and reported in terms of total and partial pressures and gas solubility. These data are very useful for modeling the gas absorption/stripping in solvents.

The solubility of hydrogen sulfide in a w(DEA) = 41.78 % aqueous solution at (313 and 373) K was recently investigated by Sidi-Boumedine et al.[90] using a computer-operated static apparatus based on the synthetic method. A table containing an overview on the previous measurements is also given. At these conditions of temperature and amine concentration, no similar data are available for comparison. Maddox and Elizondo[91] measured the solubility of hydrogen sulfide in aqueous DEA solutions {w(DEA) = (20, 35, and 50) %, between (299.85 and 388.75) K}. Their data for w(DEA) = 35 % at 299.85 K are in good agreement with those published by Lee at al.[92] at 298 K for loadings (α) up to about 0.4. At the same solution concentration, the experimental data obtained by Maddox and Elizondo[91] at 388.75 K agree with those by Lee at al.[92] obtained at 393.15 only for loadings up to about 0.1. The discrepancies between these two data sets increase significantly with the increasing of the loadings. Cheng et al.[88] measured the solubility of hydrogen sulfide at very low pressures at (298 and 313) K, in a w(DEA) = 20 % aqueous solution, using an electrode method. Unfortunately, as in the case of the MEA, experimental data are not tabulated. The authors compared graphically their data at 298 K with previous measurements performed by Rochelle et al.[87] and Lee at al.[92,93] and found a fair agreement. However, as the references concerning the works by Lee at al.[92,93] are not carefully given in the reference list, it is not very clear with what kind of data the comparison was done. Jagushte and Mahajani[94] also used an electrode method for studying the absorption of hydrogen sulfide in a w(DEA) = 20 % aqueous solution at (313 and 323) K and very low pressures. The method used by Rochelle et al.[87] and Cheng et al.[88] is based on the Kent and Eisenberg[95] model. In the method used by Jagushte and Mahajani,[94] the equilibrium H_2S pressure at these very low loadings is obtained experimentally. The authors found a fair agreement between their data at 313 K and those reported by Lal et al.[96] (as the given reference was written incorrectly, we believe this should be the correct one). However, their correlation between the H_2S partial pressure and a function of the loading capacity (a) does not seem to fit for all conditions of temperature and liquid concentration. Data by Lal et al.[96] at 313 K in a w(DEA) = 20 % aqueous solution were also found to be compatible with those by Lawson and Garst[89] given at 311 K in a w(DEA) = 25 % aqueous solution. Comparison of the results published by these last authors between (311 and 339) K and w(DEA) = 25 % with the previous measurements by Atwood et al.[83] at the same concentration and temperatures between (311 and 333) K have also shown a good agreement. Other previous measurements on this system were performed by Bottoms[97] between (298 and 328) K (w(DEA) = 50 %), who presented the data only graphically, and by Leibush and Shneerson[78] between (288 and 323) K (w(DEA) = (10 and 20) %).

Tertiary amines like TEA and MDEA have also found interest in the gas-treating processes where selective removal of hydrogen sulfide in necessary. Jagushte and Mahajani[94] used an electrode method (as in the case of DEA) for studying the absorption of hydrogen sulfide in w(TEA) = (30 and 55) % aqueous solutions at 313 K and very low pressures. Their correlation between the H_2S partial pressure and a function of the loading capacity fit generally quite well for this system in the investigated conditions. Data on the same system have also been published by Atwood et al.[83] between (300 and 333) K and amine concentration of w(TEA) = (15 to 50) % and by Jou et al.[98] between (298 and 398) K and amine concentration of w(TEA) = (30 to 70) %. A good agreement is generally found between similar data from these two references. The experimental data by Jagushte and Mahajani[94] are not tabulated, and there are no similar data at 313 K. However, the solubility in a w(TEA) = 30 % solution given by Atwood et al.[83] at 311 K for a H_2S partial pressure of 0.0757 kPa seems to be consistent with that showed graphically by Jagushte and Mahajani[94] at 313 K. Because for this very low pressure a discrepancy was revealed in Fogg and Young[1] between the two sets of measurements reported by Atwood et al.[83] and Jou et al.,[98] new data for these conditions would be very useful.

A quite important amount of experimental data for the system MDEA + hydrogen sulfide appeared in the literature after 1991. Data for this system were previously published in the open literature by Jou et al.[50] between (298 and 393) K for MDEA solutions of (1, 2, and 4.28) mol·L^{-1} (w(MDEA) = (11.8, 23.4 and 48.9) %) and by Maddox et al.[85] between (298 and 388.7) K for w(MDEA) = (11.8 and 20) % solutions. These two data sets are in good agreement, except for one point at the lowest partial pressure. Sidi-Boumedine et al.[90] investigated the solubility of hydrogen sulfide in a w(MDEA) = 46.78 % aqueous solution at (313 and 373) K using a computer-operated static apparatus based on the synthetic method. The authors present in a table a review of previous data on this system; regrettably, there are many errors in the presentation of this summary, which has the inevitable consequence in the discussion of the results. MacGregor and Mather[99] measured the solubility of hydrogen sulfide in w(MDEA) = 23.4 % aqueous solution at 313 K using an apparatus consisting of a Jerguson liquid level gauge where the gas is circulated through the solvent to reach equilibrium. The authors found a good agreement between their data and those given by Jou et al.[50] Jou et al.[100] used a similar apparatus[50] for measuring the solubility of hydrogen sulfide in w(MDEA) = (35 and 50) % aqueous solution at 313 K and in w(MDEA) = 35 % aqueous solution at 373 K. For a temperature of 313 K, their data at w(MDEA) = 35 % fall generally between the values at w(MDEA) = (23.4 and 48.9) %.[50,99] The experimental pressures at w(MDEA) = 50 % are slightly higher than those given by Jou et al.[50] at w(MDEA) = 48.9 %, even if the effect of such a small difference in concentration should be much less evident. On the contrary, at the same temperature, data by Sidi-Boumedine et al.[90] at w(MDEA) = 46.78 % are systematically higher than those given by Jou et al.[100] at w(MDEA) = 50 % (for loadings lesser than about 0.2) and systematically much higher than those given by Jou et al.[50] at w(MDEA) = 48.9 % (for loadings higher than about 0.8). However, data by Sidi-Boumedine et al.[90] are generally consistent with those reported by Pérez-Salado Kamps et al.[101] for MDEA solution of w(MDEA) = 48.8 % and by Jou et al.[50] for a concentration of w(MDEA) = 48.9 %. Li and Shen[86] studied the solubility of hydrogen sulfide in w(MDEA) = 30 % aqueous solution between (313 and 373) K using vapor-recirculation equilibrium still. Their data at 313 K lie between those given by Jou et al.[100] in w(MDEA) = 35 % and those by MacGregor and Mather[99] in w(MDEA) = 23.4 % at loading lower than about 0.6. With the increase of hydrogen sulfide loading, the difference between their data and those given in these two data sets chosen for comparison become higher. As an illustration, for a loading of 0.902, the H_2S partial pressure is 445.7 kPa in Li and Shen;[86] for a loading of 0.869, the H_2S partial pressure is 103 kPa in Jou et al.;[100] for a loading of 0.895, the H_2S partial pressure is 108 kPa in MacGregor and Mather.[99] At 373 K, data by Li and Shen[86] are much lower than those given by Jou et al.[100] at low loadings (up to 0.12) and become very close at higher loadings. Kuranov et al.[26] investigated this system between (313 and 413) K and very large pressures in w(MDEA) = (19.3 and 32.3) % aqueous solution using a static analytical method. For an amine concentration of w(MDEA) = 32.3 %, their data at (313 and 373) K in w(MDEA) = 32.3 % were found to be consistent with those given by Jou et al.[100] in w(MDEA) = 35 % only at low gas solubilities {up to about (1 and 3) mol·kg^{-1} (H_2S) at (373 and 313) K, respectively}. The hydrogen sulfide solubility in MDEA solutions of w(MDEA) = (11.83 and 23.63) % at (298 and 313) K was measured at low partial pressures (gas loadings from 0.01 to 0.26) by Lemoine et al.[102] using a new apparatus based on the static synthetic method. Large discrepancies were observed between their data and similar data reported by Jou et al.[50] at very low loadings. However, the authors concluded that their data along with those by Jou et al.[50,100] and MacGregor and Mather[99] can lead to acceptable correlations in a wide range of temperature, acid loadings and MDEA concentrations. Xu et al.[103] also measured the solubility of hydrogen sulfide in w(MDEA) = 30 % aqueous solution between (313 and 373) K. Unfortunately, their data are not tabulated. A graphical comparison for a temperature of 313 K shows that their data lie between those given by Jou et al.[100] in w(MDEA) = 35 % and those by MacGregor and Mather[99] in w(MDEA) = 23.4 % at loading higher than about 0.3. At lower loadings, the Xu et al.[103] data are higher than those given in these two data sets.

MDEA can be particularly used for the selective removal of hydrogen sulfide from natural gases containing also carbon dioxide. In order to improve the absorption of hydrogen sulfide, several compounds can be added as modifiers. Data concerning the solubility of sour gases in the single and mixed solvents are therefore necessary. Xia et al.[104] studied the solubility of hydrogen sulfide in (2 and 4) mol·kg^{-1} aqueous solutions of piperazine ($PIPH_2$) between (313 and 353) K and a mixed solvent containing MDEA (2 mol·kg^{-1}) and piperazine (2 mol·kg^{-1}) at 353 K. Anoufrikov et al.[105] investigated the solubility of hydrogen sulfide in aqueous MDEA solutions in the presence of strong electrolytes (H_2SO_4 or Na_2SO_4) from (313 to 393) K. Due to the reaction of MDEA with the sulfuric acid to form MDEA sulfate, the aqueous systems studied are H_2S + MDEA (2 mol·kg^{-1}) + MDEA sulfate (1 mol·kg^{-1}) and H_2S + MDEA (2 mol·kg^{-1}) + Na_2SO_4 (1 mol·kg^{-1}).

The solubility of hydrogen sulfide in aqueous solutions of DGA was investigated by Martin et al.[106] in w(DGA) = 60 % at (323 and 373) K, by Dingman et al.[107] in w(DGA) = 65 % from (311 to 355) K, and by Maddox et al.[85] in w(DGA) = (40 and 60) % from (313 to 353) K. For an aqueous solution of w(DGA) = 60 % and 323 K, there is a reasonable agreement between data measured by Maddox et al.[85] and those by Martin et al.[106] only at loadings up to about 0.71 mol of H_2S/mol of DGA. At higher loadings, the two data sets differ substantially, except for the highest loadings (about 1.1) when the data tend to agree again. Fogg and Young[1] also mentioned that the first two sets of measurements[106,107] are consistent with each other in the pressure range over which measurements overlap, from (4 to 180) kPa (up to about 0.75 mol of H_2S/mol of DGA).

The use of aqueous solutions of sterically hindered amines has quite recently become of great interest as potential acid gas removal from natural gas.[108,109] These amines represent attractive candidates for selective removal of hydrogen sulfide from natural gases containing carbon dioxide. By neutralizing the solutions with hydrochloric acid, Munder et al.[110] measured the physical solubility of hydrogen sulfide in aqueous solutions of 2-(*tert*-butylamino)ethanol (TBAE) using a modified Zipperclave reactor. The authors report the Henry's law constants necessary for modeling the process absorption rates. The solubility of hydrogen sulfide in aqueous solutions of AMP was first studied in the research group of Mather[111−113] using a similar equilibrium cell (vapor-phase recirculation Jerguson liquid level gauge) as that used by Jou et al.[50] Data are reported between (313 and 373) K, in (2 and 3.43) $\text{mol}\cdot\text{L}^{-1}$ amine solution {about $w(\text{AMP}) = (18 \text{ and } 32)$ %}; the corresponding sets of measurements from these references complementing each other. Li and Chang[114] used a vapor-recirculation equilibrium still for measuring new hydrogen sulfide solubilities between (313 and 373) K in $w(\text{AMP}) = 30$ % aqueous solutions. The same apparatus was also used by Jane and Li,[76] whose data obtained at 313 K in $w(\text{AMP}) = 32$ % aqueous solution agree generally well (except for one point at the highest pressure) with data given by Roberts and Mather.[111]

The mixed chemical/physical solvents can also been used to remove acid gases from gas streams. They combine the advantages of chemical (usually, aqueous solutions of alkanolamines) and physical solvents (usually, organic compounds with elevated boiling points). Jou et al.[115] investigated the solubility of hydrogen sulfide in an aqueous mixed solvent containing 2-piperidineethanol ($w(\text{2-PE}) = 45$ %) and sulfolane ($w(\text{TMS}) = 40$ %) at (313 and 373) K using a similar equilibrium cell as that used by Jou et al.[50] Using an apparatus based on a synthetic method, Murrieta-Guevara and Trejo Rodriguez[71] demonstrated the absorption capacity of *N*-methylpyrrolidone (NMP) for acid gases at low pressures up to about 0.17 MPa. They further extended their measurements to high pressures. The solubility of hydrogen sulfide in mixtures of MEA ($w(\text{MEA}) = 15$ %) + NMP and DEA ($w(\text{DEA}) = 30$ %) + NMP between (298 and 373) K was determined by Murrieta-Guevara et al.[116] The solubility of hydrogen sulfide was also investigated in another mixed solvent obtained by replacing the physical solvent, NMP with another one (i.e., tetramethylene sulfone (sulfolane, TMS)). Murrieta-Guevara et al.[117] measured the solubility of hydrogen sulfide in the mixtures MEA ($w(\text{MEA}) = 15$ %) + TMS and DEA ($w(\text{DEA}) = 15$ %) + TMS, at 303 K. The measurements on the same systems were further performed at new temperatures and solution concentrations: MEA ($w(\text{MEA}) = 15$ %) + TMS at (323 and 373) K and MEA ($w(\text{MEA}) = 30$ %) + TMS between (303 and 373) K[118] as well as DEA ($w(\text{DEA}) = 15$ %) + TMS at (323 and 373) K and DEA ($w(\text{DEA}) = (30 \text{ and } 50)$ %) + TMS between (303 and 373) K.[119] The authors discuss the absorption capacity of hydrogen sulfide in various mixed solvents. MacGregor and Mather[99] investigated the solubility of hydrogen sulfide in an aqueous mixed solvent containing MDEA ($w(\text{MDEA}) = 20.9$ % (2 $\text{mol}\cdot\text{L}^{-1}$)) and TMS ($w(\text{TMS}) = 30.5$ %) at (313 and 373) K. The authors compared the solubility results for H_2S in the mixed MDEA with those for aqueous MDEA (2 $\text{mol}\cdot\text{L}^{-1}$) and another mixed solvent obtained by changing the chemical solvent {namely, AMP ($w(\text{AMP}) = 16.5$ % (2 $\text{mol}\cdot\text{L}^{-1}$)) + TMS ($w(\text{TMS}) = 32.2$ %)}.[120] Xu et al.[103] measured the solubility of hydrogen sulfide in aqueous and nonaqueous mixed solvents containing MDEA and ethylene glycol (EG) between (298 and 333) K. Unfortunately, their data are not tabulated. The authors show graphically at 313 K the influence of the EG on the solubility of hydrogen sulfide in the aqueous MDEA solution. Using a static analytical method, Pohorecki and Możeński[121] investigated another mixed solvent consisting of TEA and propylene carbonate (PC). Experimental data are not tabulated. The authors compare graphically at (283 and 313) K the solubility of hydrogen sulfide in the pure PC with that in the mixed solvent.

The blended amines combine the absorption characteristics of their constituents and can lead to considerable improvement in absorption and important savings in energy requirements.[122] Li and Shen[86] investigated the solubility hydrogen sulfide in aqueous mixtures of MEA and MDEA at various compositions between (313 and 373) K. The authors discussed the effect of amine concentration on the H_2S loading and compared the solubility of hydrogen sulfide in the aqueous MEA, MDEA, and the mixtures of these amines. The system MEA + AMP (a sterically hindered alkanolamine) $+H_2S$ (various compositions) was studied by Li and Chang[114] between (313 and 373) K. Sidi-Boumedine et al.[90] measured the solubility of hydrogen sulfide in aqueous mixtures containing DEA ($w(\text{DEA}) = 7.64$ %) and MDEA ($w(\text{MDEA}) = 37.73$ %) at (313 and 373) K. The authors compared the solubility of hydrogen sulfide in the aqueous DEA, MDEA, and amine mixtures thereof. No similar data were reported in the open literature for these blended amines.

Methyl Mercaptan Solubility. Williams and Murray[123] used the chromatographic method to analyze the solubility of Kraft mill sulfides, including methyl mercaptan, in water between (273 and 333) K. The effect of pH on the gas solubility was also investigated. Unfortunately, their data are not tabulated. Using headspace gas chromatography, Field and Gilbert[124] give the value of the distribution coefficient of methyl mercaptan at 333 K. The solubility of methyl mercaptan in water was also measured by Harkness and Kelman[125] at 303 K using the static saturation method. The authors show graphically the influence of temperature on the gas solubility, but only the value at 303 K is tabulated (expressed as a Bunsen coefficient). The solubility of methyl mercaptan was also measured in various electrolyte solutions at the same temperature. Przyjazny et al.[126] determined the distribution coefficient of various organosulfur compounds, including methyl mercaptan, between (298 and 343) K using the headspace gas chromatography. The authors did not report the raw experimental data; the values of the gas−liquid distribution coefficients (K) were therefore calculated from the coefficients of the linear correlation $\log(K)-T$. The value of the distribution coefficients at 333 K agrees well with that given by Field and Gilbert.[124] The solubility measurements have also been performed in artificial seawater (ionic strength of 0.7). Gillespie and Wilson[127] used a rocked stainless steel cylindrical cell for measuring the solubility of methyl mercaptan in water on a very large temperature range between (310.9 and 588.7) K and total pressures between (0.207 and 20.68) MPa. Both vapor and liquid phases were sampled and analyzed. The authors showed that the variation of the Henry's law constants with temperature was similar to that found for the hydrogen sulfide-water system, that is, the curve presenting a maximum in temperature. To our knowledge, no other experiments were performed for temperatures higher than 373 K. The solubility of methyl mercaptan in water was also measured by Murakami et al.[128] at (273.15, 298.15, and 323.15) K. On the basis of a static headspace method, Iliuta and Larachi[129] measured the gas−liquid partition coefficients and Henry's law constants of methyl mercaptan in water, aqueous iron-free CDTA solutions, and ferrous chelate solutions (CDTA-Fe(II) complex) between (298 and 333) K and chelate concentrations varying between (38 and 300) $\text{mol}\cdot\text{m}^{-3}$. Solubility data in water are in good agreement with those reported by Przyjazny et al.[126] on the whole temperature range and with those given by Murakami et al.[128] at 323 K. No similar data exist in the literature for the chelate systems. MM, the first member of the thiols, is a much weaker acid than hydrogen sulfide. There is little or no reaction between the mercaptans and the alkanolamines usually used for removing the acid gases from the natural gas. The solubility of methyl mercaptan in the alkanolamines was therefore investigated in the absence and the presence of acid gases. Jou et al.[130] investigated the solubility of MM in w(MDEA) = 50 % aqueous solution at (313 and 343) K and high pressures using an equilibrium cell consisting of a Jerguson liquid level gauge.[131] Their data concerning the solubility of methyl mercaptan in water between (298 and 373) K agree well with those given by Kilner et al.[132] However, on the temperature range of (273 to 373) K, the agreement between all available data in the literature concerning the solubility of MM in water is quite poor, especially at higher temperatures. A similar apparatus[133] was used to measure the solubility of MM in w(DEA) = 50 % aqueous solutions at (313 and 343) K and high pressures.[134] No previous data were found in the literature for these systems. Due to the industrial importance of mixtures containing hydrocarbons and mercaptans in the oil and gas fields processes and the necessity to model the distribution of the mercaptans between various streams containing light hydrocarbons and mercaptans, Kilner et al.[132] investigated the systems hexane + MM and toluene + MM between (323 and 373) K and high pressures using an equilibrium cell based on the static synthetic method. No similar data were found in the literature.

Dimethyl Sulfide Solubility. Unfortunately, the solubility data for dimethyl sulfide are in general quite scarce over a larger temperature range. Hine and Weimar[135] determined the solubility of dimethyl sulfide in water by UV measurements at 298 K. By using the chromatographic method, Williams and Murray[123] studied the solubility of dimethyl sulfide in pure water and at different pH between (273 and 333) K. The authors did not report the data in a tabulated form. Vitenberg et al.[136] determined the partition coefficient at 293 K using a static method based on gas chromatographic measurements. The authors found a difference of about 13 % from the data given at the same temperature by Williams and Murrey.[123] Przyjazny et al.[126] measured the distribution coefficients between (298 and 343) K using the headspace gas chromatography. The raw experimental data were not given; the values of the distribution coefficients were therefore calculated from the coefficients of the linear correlation log(K)−T. The corresponding value calculated from this equation at 293 K agrees well with that derived graphically from Williams and Murrey.[123] The solubility measurements have also been performed in artificial seawater at different ionic strengths from (0.7 to 4). Dacey et al.[12] reported the values of Henry's law constant for dimethyl sulfide in pure water and several natural waters of varying salinity between (272 and 305) K. Their results agree generally well with those given by Hine and Weimar[135] and Przyjazny et al.[126] for the system dimethyl sulfide + water and with those given by Cline and Bates[137] and Przyjazny et al.[126] for the aqueous electrolyte systems. Murakami et al.[128] reported the solubility of dimethyl sulfide in water at (273.15, 298.15, and 323.15) K. Lovelock et al.[138] also measured the dimethyl sulfide solubility in seawater, but the Henry's law constant derived from the distribution coefficient is abnormally high with respect to other data. Because the temperature and the solution concentration were not given, no comparison with other work is possible. Wong and Wang[139] used a headspace method for measuring the Henry's law constant of dimethyl sulfide in water at 291 K and in the seawater between (291 and 317) K. The authors showed the effect of NaCl on the Henry's law constants. It should also be mentioned that in two theoretical papers Brennan el al.[140] and Yaws et al.[141] use in the correlative methods the values of Henry's law constants at (293 and 298) K, respectively, taken from existing data banks without however specifying the exact data source for a particular system. Iliuta and Larachi[142] used a static headspace method for measuring the gas−liquid partition coefficients and Henry's law constants of dimethyl sulfide in water, aqueous iron-free CDTA solutions, and ferrous chelate solutions (CDTA-Fe(II) complex) between (288 and 308) K and chelate concentrations varying between 38 and 300 mol·m^{-3}. Solubility data in water were in good agreement with most previous data. The authors showed that the data from Vitenberg et al.[136] and those given in Brennan et al.[140] do not seem very plausible. No similar data exist in the literature for the chelate systems. Using an apparatus based on dynamic method, Coquelet and Richon[143] measured the Henry's law constants and infinite dilution activity coefficients of dimethyl sulfide in water between (293 and 333) K and in w(MDEA) = 50 % aqueous solution between (293 and 343) K at atmospheric pressure. The solubility data of dimethyl sulfide in water agree well with most of the previous data. No similar data exist in the literature for the amine containing system.

Dimethyl Disulfide Solubility. Experimental data for this system are even scarcer than in the case of MM and dimethyl sulfide. Besides the systems containing mercaptan and dimethyl sulfide, Williams and Murray[123] also studied the solubility of dimethyl disulfide in pure water and at different pH between (273 and 333) K. Experimental data are reported only graphically. Vitenberg et al.[136] determined the partition coefficient at 293 K using a static method based on the gas chromatographic measurements. The authors found a difference of about 9 % from the data given at the same temperature by Williams and Murrey.[123] Przyjazny et al.[126] measured the distribution coefficients between (298 and 343) K using headspace gas chromatography. Because the raw experimental data were not given, the values of the distribution coefficients were calculated from the coefficients of the linear correlation log(K)−T. The extrapolated value calculated from this equation at 293 K is higher than that given by Vitenberg et al.[136] and Williams and Murrey.[123] Przyjazny et al.[126] also reported the solubility in artificial seawater at different ionic strengths from 0.7 to 4. Murakami et al.[128] report the solubility of dimethyl disulfide in water at (273.15, 298.15, and 323.15) K. Iliuta and Larachi[144] used a static headspace method for measuring the gas−liquid partition coefficients and Henry's law constants of dimethyl disulfide in water, aqueous iron-free CDTA solutions, and ferrous chelate solutions (CDTA-Fe(II) complex) between (298 and 338) K and chelate concentrations varying between (38 and 300) mol·m^{-3}. Solubility data in water were found in fair agreement with most of the previous data. No similar data exist in the literature for the chelate systems.

Conclusions and Recommendations for Future Research

An important body of experimental work was reported in the open literature for the solubility of hydrogen sulfide (from 1987 to 2006, generally before January 2006) and MM, DMS, and DMDS (generally before January 2006), especially due to their toxic and corrosive characters for industrial and environmental processes and the increasing interest of removal of these sulfur compounds from gaseous streams. Obviously, the systems containing hydrogen sulfide were much more extensively investigated, and the available data are therefore quite abundant, especially because it is the most common sulfur-containing compound encountered in various petroleum and non-petroleum applications. The systems containing DMS and DMDS are mainly investigated in direct connection to the environmental natural processes. Even though MM can also be present in various gaseous streams along with the hydrogen sulfide, investigations concerning the systems containing this sulfur component are very limited.

As mentioned in the analysis of existing data, new experimental work for various systems would be useful for the elucidation of contradictory behaviors or for completing the existing data base, for example: (1) water + hydrogen sulfide at high temperatures, mainly above 500 K; (2) water + hydrogen sulfide + NaCl at temperatures higher than 400 K; (3) methanol + hydrogen sulfide at low temperatures, especially below 273 K; (4) cyclohexane + hydrogen sulfide at low pressures; (5) hexadecane + hydrogen sulfide; (6) triethylene glycol + hydrogen sulfide at temperatures larger than 323 K; (7) tetraglyme + hydrogen sulfide at temperatures higher than 323 K; (8) sulfolane + hydrogen sulfide at temperatures larger than about 350 K; (9) monoethanolamine (or diethanolamine, or triethanolamine) + water + hydrogen sulfide at very low partial pressures; (10) methyldiethanolamine + water + hydrogen sulfide at several conditions (T, P, loading capacity) where data are contradictory, as mentioned in the discussion section; (11) diglycolamine + water + hydrogen sulfide at loadings higher than about 0.7; (12) water + methyl mercaptan, especially above 323 K; (13) monoethanolamine + water + methyl mercaptan; (14) diethanolamine (or methyldiethanolamine) + water + methyl mercaptan at various liquid concentrations; (15) mixed solvents + methyl mercaptan. In addition, given the importance of the presence of dimethyl sulfide and dimethyl disulfide, even at very low concentrations (i.e., traces), in several gas streams, new experimental data for systems including these sulfur components would be useful for many industrial applications.

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