THE SOLUBILITY OF GASES IN LIQUIDS

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CONTENTS

I. Introduction

The solubility of gases in liquids was one of the physical properties studied by the early chemists. Many well-known names appear in the literature, as Henry in 1803, Berthelot, Bunsen, Carius, and Roscoe in 1855, Winkler in 1889, and other later workers. In general, the solubility of gases in liquids has been studied by investigators who have wished to have these particular bits of data in the investigation of the gas or, more infrequently, of the liquid or the solution. Few were interested in the phenomenon itself and in studying general behavior in the solubility of gases in liquids. The literature cited shows the names of a large number of workers,—men who studied the solubility of all gases in a wide variety of liquids, men who used all types of apparatus and experimental conditions, and men who exercised various degrees of experimental technique in producing their data. The results vary from those of high precision to those little more than qualitative. Many investigators neglected some of the important factors in solubility or failed to record all the data. Most of the work has been done at random pressures near atmospheric, and the values corrected to 760 mm. by means of Henry's law; this is usually permissible. Many workers have failed to indicate whether the pressure was total or partial, thus introducing a large uncertainty. The data usually are calculated to either the Bunsen or the Ostwald coefficient, and if the worker does not state which is used, the values diverge as the temperature increases from 0°C.

Some attempts have been made to correlate gas solubility with the properties of the liquid or the solution. Most of these are entirely empirical and are based on a few data secured by one investigator.

It is the purpose of this paper to review the work done on the solubility of gases in liquids and to discuss the various factors of importance in this field of work. Solubilities at high pressures have not been included, because of the special apparatus and technique used. The merits and limitations of the various experimental methods are described. The equations proposed for correlating gas solubility with other variables have been collected and rewritten in a uniform system. All existing data on the solubility of gases in liquids have been tabulated, so that the user may know the range of temperature and pressure of the experiments and the relative precision of the data. It has not been possible in this review to collate the data, but an indication is given of the probable precision of the results on the basis of method used, the completeness of the data, and the consistency of the results among themselves.

II. Methods and Apparatus

Many methods of procedure and kinds of apparatus have been used in the measurement of gas solubility. Most of these can be classed definitely as either chemical or physical. Chemical methods depend on specific chemical properties of the gas, and thus can be used with only a limited number of gases. Physical methods usually depend on no such properties and are thus more general. When suitable chemical methods are available, however, they are frequently more accurate and usually much quicker.

A. PHYSICAL METHODS

The physical methods used are quite varied. Most of them are saturation methods, in which the measurement is that of the quantity of the gas necessary to saturate a quantity of initially gas-free solvent. Some are extraction methods, in which the measurement is that of the volume of gas that can be extracted from a quantity of saturated solution.

1. *Removal of gases from solvent*

In saturation methods, the liquid must be gas-free at the start. In extraction methods, the gas is all to be extracted. Hence, in either case, the complete removal of gas from a liquid is important. This is not an easy matter. In saturation methods the problem is not as difficult as in extraction methods, since loss of gas is permissible, loss of solvent usually is, and usually the solvent as obtained contains only atmospheric gases. The presence of traces of the atmospheric gases probably has little effect on the solubility of others, though there is no authority in the literature on this point. Buchanan (36) made quantitative measurements of the extraction of carbon dioxide from water and from aqueous salt solutions by boiling. He distilled solutions saturated with the gas, and tested portions of the distillate for carbon dioxide with barium hydroxide. He found that when the solvent was distilled water, the first eighth of the distillate contained nearly all the gas, the second eighth a trace, and the rest none. When the solution contained sulfates, he found it necessary to boil nearly to dryness to remove all of the carbon dioxide. If the sulfate were removed by the addition of barium chloride, the resulting solution gave up its gas about as readily as did distilled water. Leduc (176) found that even after boiling distilled water a long time, it gave up gas bubbles on freezing. Successive freezing in a vacuum did not free the water completely from gas. Metschl (207) made measurements of the gas liberated when a solution saturated at several atmospheres pressure was shaken at 1 atmosphere. Water and organic liquids were the solvents, and the gases included hydrogen, oxygen, nitrogen, and carbon dioxide. All of these except the latter were readily liberated, in amounts predicted by the solubility figures in the literature; hence equilibrium was evidently reached. In the case of carbon dioxide, however, the results indicated either that gas was lost before measurement or that equilibrium was not reached by shaking. Porter (246) made statements about the difficulty of removing gas from water, but evidently these were due largely to misunderstanding of the solubility curves (see Sillitto (286)). Seyler (281) raised the temperature of water samples containing oxygen, to see if the oxygen were lost. He concluded that, when the solution was not shaken, the gas stayed in solution in going 5 or 6 degrees above the equilibrium temperature, but that shaking established equilibrium.

The usual method of preparing gas-free liquid for solubility measurements has been boiling, followed by cooling in a vaccum. Bunsen (37, 38, 39, 40) used this method, and most others have followed him. Hibben (120) has applied vacuum sublimation to prepare gas-free liquid.

Paunov (234) found that, under the influence of ultrasonic frequencies, the amounts of gas absorbed decrease by about 50 per cent.

2. Saturation methods

The general principle most frequently employed is the measurement of a volume of gas before it is brought in contact with a quantity of gas-free solvent, and its measurement again after equilibrium is established. The volume dissolved is found by difference. There are many different arrangements of the essential parts of the apparatus to achieve this end. Henry (118) used this method in 1803. The impure gases that he used and the limitations in material necessarily caused results of a very low degree of accuracy. Bunsen (37, 38, 39, 40) used an apparatus that he designed, which employed the same principle. His apparatus is shown in figure 1.

The calibrated absorption tube e is fastened at the bottom to a small iron band b; this screws into the small iron stand a. By this arrangement the open end of the tube can be screwed tight against a plate of rubber covering the lower surface of the stand. Thus the tube can be completely sealed. On each side of the stand are two steel springs c, which fit into two upright grooves in the wooden base, f, of the apparatus. When the tube and stand are in place, it is easy to open or close the absorption tube by giving it a turn to the right or left. The water jacket g is held firmly in place by the screws ii. The tube r is for the purpose of pouring in mercury and removing it, so that any desired pressure in the absorption bulb can be obtained by adjusting the mercury level in the water jacket. The temperature of the water can be read on the thermometer d. The water jacket is closed on top by a hinged lid. The piece of rubber s serves to hold the tube in place during the shaking necessary in the process of absorption.

The experiment is conducted in the following manner: A volume of the gas to be examined is first collected in the tube over mercury, and its volume, temperature, and pressure are read. A measured volume of air-free water is introduced under the mercury into the tube, which is then sealed by being screwed tightly against the rubber plate. The tube is then placed in the water jacket, which contains some mercury at the bottom. The pressures within and without are equalized by turning the tube slightly. The tube is then sealed again by turning, and vigorously shaken. This agitation, with opening and closing of the tube, is repeated many times, until no further change of volume is perceptible. The ob-

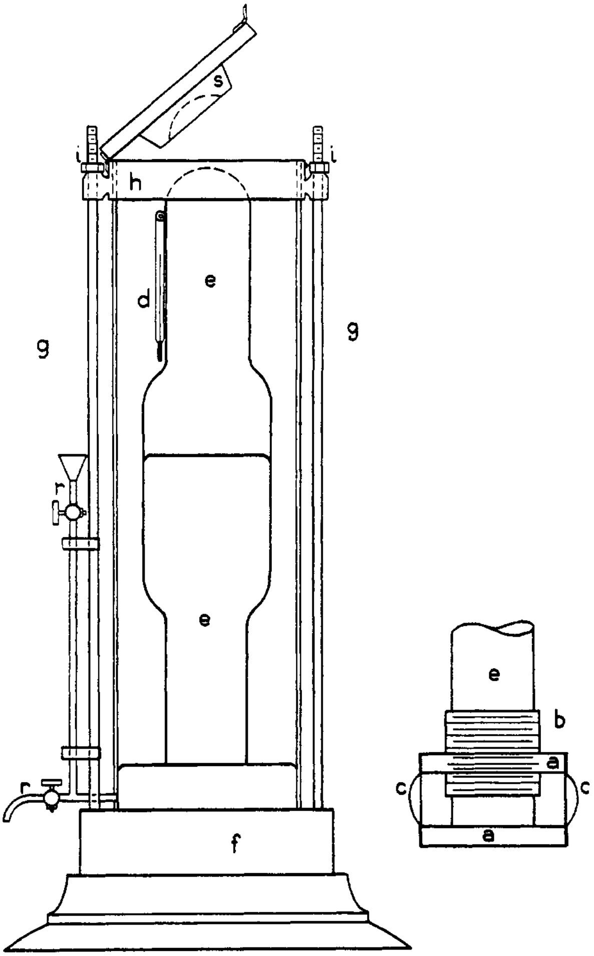

FIG. 1. Bunsen's apparatus for the determination of the solubility of gases

servations necessary for the measurement and reduction of the residual gas are then made.

Bunsen applied this method with considerable success to the measurement of the solubilities of the common gases, including hydrogen, oxygen, nitrogen, air, methane, carbon dioxide, ethylene, and ethane, in water. Carius (44) used Bunsen's apparatus for the same gases in alcohol; Schickendantz (267) used it for ethane in water; and Than (306) used it for propylene in water. Khanikoff and Luginin (149) used an apparatus similar to Bunsen's for the carbon dioxide–water system at pressures up to several atmospheres. Maclaurin (192) modified Bunsen's method to avoid contact of the solution with mercury, for the measurement of the solubility of oxygen in potassium cyanide solutions. Ramsay and coworkers (250, 251) used Bunsen's method to determine the solubility of argon and of helium in water. Tower (314) used it with nitric oxide in sulfuric acid. Sander (266) used a modification of Bunsen's method for the measurement of the solubility of carbon dioxide in water and in organic solvents at high pressures, up to 140 atmospheres. For the measurement of the solubility of a number of gases in cyclohexanol Cauquil (46) used a method which, from his description, appears to be similar to Bunsen's.

Ostwald (231a) introduced a method which proved to be much better than Bunsen's and within a few years almost entirely displaced it. The fundamental difference was that the gas is measured in a buret connected to the absorption vessel, rather than in the absorption vessel itself. The buret and leveling tube are similar to those used in the Hempel gas analysis equipment. From one arm of a three-way stopcock a flexible capillary tube connects to an absorption bulb, similar to a gas-sampling bulb with three-way cocks at the ends. Ostwald usually used lead for the flexible capillary, though silver and platinum are mentioned. This assembly of equipment is capable of almost infinite variation and refinement. McDaniel (204) used the apparatus in substantially its original form; his apparatus is shown in figure 2.

The essential parts are a gas buret, A, connected by a capillary tube, M, to an absorption pipet B, so that the entire apparatus is of glass. The buret and pipet are inclosed within water jackets, the temperature of each being regulated by electrically heated coils in the water. The whole apparatus is clamped solidly on a rigid frame, so that it can be taken in the hands and shaken to bring the gas into intimate contact with the liquid.

In operation, the pipet B is filled completely with gas-free solvent. The source of gas is connected to the apparatus at T, and the gas is passed through a saturator H filled with the solvent. First, the stopcocks C and D are turned so that the capillary is filled with the gas, C is then closed and D is opened to the buret A, which is filled with the gas; measurement is made by adjusting the mercury leveling bulb F. The source of gas then is disconnected, D and C are adjusted to allow gas to flow into pipet B, stopcock G is opened, and a measured volume of solvent is withdrawn to give a gas volume in the pipet B. The mercury level in the buret is adjusted, stopcocks C and D are closed, and the entire apparatus is shaken to dissolve gas in the solvent. At intervals, additional gas from the buret is added to the pipet to maintain the pressure as the gas dissolves. The original volume of the liquid in the pipet minus that withdrawn gives the volume of solvent in which gas is absorbed. To the volume of gas remaining in the buret at equilibrium is added the volume of liquid withdrawn which is the new vapor space above the solution. From the volume of gas absorbed and the volume of solvent used, the solubility of the gas is calculated.

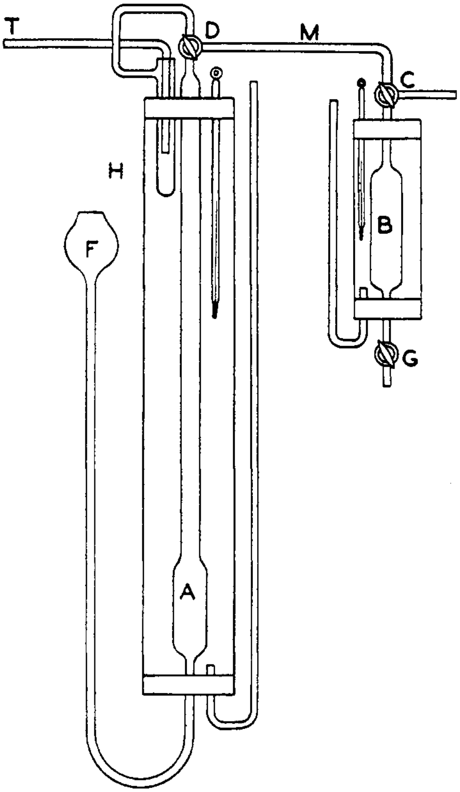

FIG. 2. Ostwald type of apparatus for the determination of the solubility of gases

While Bunsen's apparatus was ordinarily used in its original form, Ostwald's seems rarely to have been. The chief variations have been the provision for agitation of the liquid and provision for a gas–liquid interface. Usually the gas buret has been stationary and the absorption flask has been shaken. This procedure necessitates a flexible joint between the two. McDaniel shook the whole apparatus (see figure 2) as a unit. Stern (296) used a glass-capillary spiral to provide the flexible joint. Maxted and Moon (202), using an apparatus designed from that of Just (147), also used a glass spiral. In order to get more freedom from the joint, Steiner (295) used a platinum capillary, as did Timofeev (312). Secenov (279) and Estreicher (77) also used metal spirals. Curry and Hazleton (61) used a copper capillary.

If an iron bob inclosed in glass is placed in the absorption vessel and moved by a magnet the connection between this vessel and the buret can be rigid. Åkerlöf (3), Antropoff (5), Cady, Elsey, and Berger (41), Cassuto (45), and Wright and Maass (344) have made use of this idea. Showalter and Ferguson (284) used the ground-glass joint of a stopcock to provide the flexible joint.

Lunge (186) fastened the buret and absorption vessel together with a short stub of rubber tubing, thus providing a flexible joint. Others have used the same idea. Lannung (172) assembled both parts in one rigid piece and shook the whole assembly to agitate the liquid. His apparatus is shown in figure 3. Absorption bulb A is made to contain various volumes in different pieces of apparatus. The gas buret and manometer are combined in B C, behind which is a measuring scale read with a telescope. The buret is calibrated from mark *a* to mark *b*. The entire apparatus is attached to an aluminum frame so that it can be shaken. The apparatus is evacuated at *s* and mercury admitted at 1 until the apparatus is filled as far as *s*, the manometer tube C and the movable reservoir g. To the ground joint *s* is attached an L-tube, the other end of which dips into the pure solvent; the solvent is drawn into A by letting mercury run out at 1 until A is about half full. The L-tube is detached, and the solvent is deaerated by suction at *s* and confined in such a way as not to be in contact with stopcocks. The entire apparatus is placed in an air thermostat and allowed to come to equilibrium. Buret B is filled with the gas through stopcock 4 and the volume measured when saturated with solvent vapor. The solvent surface is lowered from *f* by letting mercury out at 1. The entire apparatus is agitated until equilibrium is established, the volume in B is measured, and the solubility is calculated from the decrease in volume.

In McDaniel's apparatus, the absorption vessel was filled with liquid at the start, and, to provide a suitable gas–liquid interface, liquid was drained from the bottom. The amount so drained was weighed, and thus the corresponding volume of gas that replaced it was found. Steiner

(295), Timofeev (312), and Secenov (279) provided the gas–liquid interface in the same way. As an alternative, mercury may be placed in the bottom of the vessel, to be drained out and weighed in the same manner. Cady, Elsey, and Berger (41) and Lannung (172) used this idea. Lunge (186) connected the bottom of the vessel to a leveling bulb of mercury, which could be lowered to let mercury out of the vessel. After equilibrium was reached, the bulb could be raised, filling the vessel again, and no correction was necessary for gas that replaced the liquid drained. Christoff (49) brought the liquid and gas in contact before measuring the gas

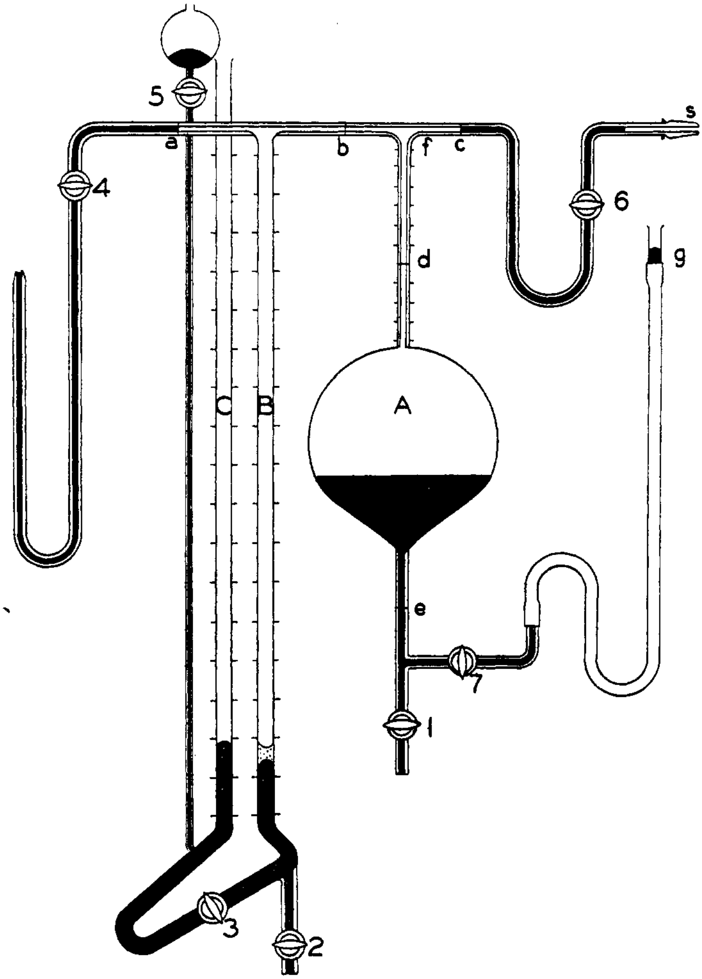

FIG. 3. Ostwald type of apparatus, as used by Lannung

volume. He found that the rate of solution was such that no appreciable volume of gas was dissolved during the measurement. Estreicher (77) and Drucker and Moles (71) left a vacuum above the liquid when the vessel was filled, which volume was filled with gas from the buret after the gas was measured. Others have done the same. Manchot (197) used an absorption vessel with two compartments, one filled initially with gas, and the other, a smaller one placed above it, with liquid. After measurement of the gas volume, the liquid could be drained into the lower compartment, filling it only part full and leaving considerable gas above it. The upper compartment was filled with gas from the lower one during the draining. Markham and Kobe (201) used a similar apparatus, in which the liquid–gas interface could be made larger relative to the volume of solvent. Usher (320) filled the absorption vessel with gas before measurement, and let in a known amount of liquid later. In his measurements with solutions, he put the solid solute in the absorption vessel before filling it with gas, then introduced the gas and made the buret reading, and later introduced a known amount of solvent, effecting solution of the solid solute in the vessel itself. Such a procedure involved the assumption that the solid did not absorb the gas. Homfray (128), working on *p*-azoxyphenetole, put the solid crystals in contact with the gas, then observed the change in the volume of the gas when the crystals were melted and later heated to the anisotropic state.

Some investigators have saturated the gas with liquid vapor before filling the buret, while others have kept the gas in the buret dry. Horiuchi (133) has discussed the relative merits of both methods. If the gas in the buret is saturated, the vapor pressure of the solvent is of little consequence. If the gas is dry, however, the vapor pressure must be known accurately, since all gas coming into the free space above the liquid in the absorption vessel picks up vapor, increasing its volume to an extent determined by the vapor pressure. On the other hand, if the gas in the buret is saturated, any part of the apparatus that is not in the thermostat may collect condensed solvent if the thermostat is above room temperature. The capillary between the buret and the absorption vessel is usually out of the thermostat. Drops of liquid in this capillary would make the pressure adjustment in the buret uncertain. If the gas in the buret is dry, the temperature of the whole apparatus can be changed and thus a range of temperature can be covered with one filling.

As ordinarily used, Ostwald's apparatus has involved at least one mercury surface in contact with the gas, and sometimes in contact with the solvent as well. This feature is a serious handicap when dealing with systems that react with mercury. To avoid this difficulty, Wright and Maass (344), working with hydrogen sulfide, used a modification in which the volume of gas remained constant, while the pressure varied and was measured with a manometer having a glass diaphragm. Bancroft and Belden (10) used a similar arrangement.

Cady, Elsey, and Berger (41) stated that the violent shaking frequently used to effect equilibrium could cause pressures at the surfaces of the absorption vessel far in excess of that measured by the manometer. Thus the solvent would be supersaturated with respect to the pressure read on the manometer, and the solubility results would be too high. To test this point, Morgan and Pyne (216) used an apparatus in which the gas was bubbled through the liquid repeatedly. They felt that no such supersaturation as Cady, Elsey, and Berger mention could result in their apparatus, and that, if they could check the solubility values found by others who had used the shaking method, the question of this supersaturation would be answered. The values obtained by Morgan and Pyne for the system carbon dioxide–water checked exactly the values found by the shaking method. Thus, in many instances at least, this method has caused no appreciable error. Hainsworth and Titus (107) also used a method in which the gas was bubbled repeatedly through the liquid. They approached equilibrium from both sides, getting the same values in each case. Thus they were certain that equilibrium had been established. Bancroft and Belden (10) found that 30 sec. of shaking established equilibrium in the hydrogen sulfide–aniline system, and that identical values were found when equilibrium was approached from either side. Rakestraw and Emmel (249) found that when sea water was shaken with air and the solution was allowed to stand till bubbles were no longer visible, the nitrogen content was 2 per cent higher than the equilibrium value.

In most methods of determining gas solubility, the average solubility throughout a volume of liquid is determined. The solubility may change with depth as a result of the hydrostatic head. Few experimenters have considered this point. Morgan and Richardson (218) determined the effect of hydrostatic head on the solubility of oxygen in water and found it to have the same effect as any other pressure.

The change in the volume of the solution as a gas dissolves necessarily introduces a certain error. The error so introduced is probably less than other experimental errors in the case of the gases of small solubility. Markham and Kobe (201) showed that the solubility of carbon dioxide in aqueous solutions might be in error up to 0.1 per cent for this reason.

Instead of measuring the gas volumetrically, gravimetric means may be used. The gas-free solution is weighed, then saturated with gas by bubbling, and then weighed again, to give the solubility. The solvent carried away by the escaping gas can be caught and weighed, and correction applied therefor. This method is limited to the more soluble gases. It has been used by Raoult (254) on the ammonia-water system, including aqueous salt solutions; by Prytz (247) on hydrogen sulfide and carbon dioxide in water; by Naumann (223) on cyanogen in water; by Baskerville and Cohen (11) on phosgene in organic solvents; and by others.

3. Extraction methods

These methods involve the extraction by some means of the gas contained in a quantity of saturated solution, and the measurement of the volume of gas so extracted. Thus the procedure of the saturation method is reversed. As mentioned earlier, the difficulty of complete extraction is of special interest here.

These methods are, in general, useful for the analysis of naturally occurring solutions, such as sea water. They were first used in this way. Later they were applied by many to artificially saturated solutions.

Reichardt (255) described an apparatus for boiling the gas out of water and collecting it for measurement. Tornöe (313) also described such an apparatus. Numerous others,—Dittmar (65), Hamberg (111), Petterson and Sonden (242), Clowes (53), Winkler (341), Weigert (331), James (142), and Ruppin (264),—have described apparatus designed for the same purpose. Buchanan (36) boiled aqueous carbon dioxide solutions, catching the distillate in barium hydroxide for analysis. The same idea has been applied to other systems (see the work of Calingaert and Huggins (42) on ammonia and water).

Other investigators have extracted the gas from solution by evacuation; e.g., Bohr (25) pumped carbon dioxide from its solution in water. The apparatus of Van Slyke (324, 241a) is the best known for this type of measurement (figure 4). It is used principally to determine the amount of gases dissolved in blood and blood fluids. The short pipet A contains 50 cc. and has several graduations on it; *a* corresponds to 2 cc. The pipet is connected to the manometer and to the mercury leveling bulb. The sample of gas solution is introduced through stopcock b by a special pipet in such a way that the solution does not come in contact with the air. Then by lowering the leveling bulb the gas solution is evacuated, and the pipet is shaken for 2 or 3 min. to assist in liberating the gas. The liberated gas is compressed into the volume *a* and the pressure read on the manometer. An empirical correction is made for the gases redissolved during the compression. The gases collected can be analyzed for carbon dioxide or oxygen by introducing the appropriate absorbent solution through stopcock b and determining the pressure after the particular component has been removed. Objections can be raised to this method because of the question concerning the extraction of dissolved gases and the corrections applied to the results. However, the data presented by the authors indicate that satisfactory results were obtained. The method of saturating blood was adopted from Austin, Cullen, and Hastings (8). Conant and Scott (55), Kubie (170), Hawkins and Shilling (114, 115) and others have used the Van Slyke method. Orcutt and coworkers (230), in using it, applied a correction for the gas not extracted. Results have been obtained by this method that check very well those obtained by saturation methods. Scotti-Foglieni (277) described an apparatus for the saturation of liquid with gas, to be followed by analysis.

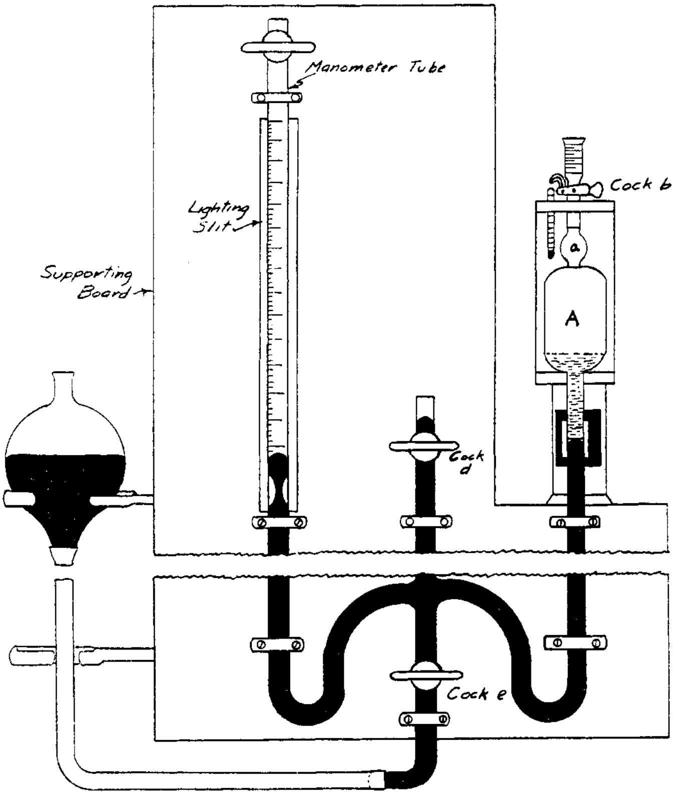

FIG. 4. Apparatus of Van Slyke

4. Miscellaneous methods

The foregoing methods have been used far more than other physical methods. A few others have been used to a slight extent.

The solubility of the radioactive gases has been measured by electrical means, determining the concentration in the liquid phase and in an inert gas phase, usually air. The concentration of the gas under investigation is necessarily very small. Traubenberg (315), Hofmann (127), Kofler (160, 161) and others have applied this method to radium emanation, Klaus (155) and Boyle (28) have applied it to thorium emanation, and Hevesy (119) has applied it to actinium emanation.

The solubility of a gas has sometimes been determined by the freezing-point lowering produced by the addition of the gas to saturation. This method is necessarily limited to one temperature for each solvent, and is uncertain because of the possibility of association or dissociation of the gas in the liquid phase. It is further limited to those systems which give a freezing-point lowering of sufficient magnitude. Prytz (247) applied this method to hydrogen sulfide and carbon dioxide in water. Garelli and Falciola (97) applied it to carbon dioxide, nitrous oxide, acetylene, and hydrogen sulfide in water and in several organic liquids. Their results checked values found in other ways. Garelli (96) and Garelli and Monath (98) also used this method. Falciola (79) reported that, if oxygen, methane, or several other gases were present, the freezing point might be raised by the addition of the above gases.

Schwab and Berninger (275) used a special method for solubility measurements, based on the change in pressure in a bubble rising in a column of liquid. The results are not very accurate.

B. CHEMICAL METHODS

In the case of the less soluble gases, the solution has usually been saturated by bubbling the gas through at atmospheric pressure, after which the content of gas has been determined by suitable means. The total pressure was that of the atmosphere, and the solvent was assumed to have the same vapor pressure as in the pure state. Obviously, the very soluble gases like ammonia and the hydrogen halides affect the vapor pressure of the solvent and must be treated differently. Their vapor pressures have usually been measured by bubbling an inert gas through the solution, and absorbing the gas under investigation from the gas stream in some medium where its quantity can be determined. Very low vapor pressures can be found in this way. The introduction of inert gas probably affects the solubility of the very soluble gas but little.

Schutzenberger (274), Mohr (213), Winkler (336), and Levy and Marboutin (178) developed chemical methods for the determination of oxygen in solution. That of Winkler, involving the oxidation by the oxygen of manganous hydroxide, the reduction by iodide of the manganic hydroxide so produced, and the titration of the liberated iodine, has proved to be

$$O_2 + 4Mn^{++} + 4H^+ \rightarrow 4Mn^{+++} + 2H_2O$$

$$2Mn^{+++} + 2I^- \rightarrow 2Mn^{++} + I_2$$

$$I_2 + 2Na_2S_2O_3 \rightarrow Na_2S_4O_6 + 2NaI$$

much the best. This method has been of value both in the analysis of natural waters and in solubility measurements. Konig and Mutschler (164), Tiemann and Preusse (311), Roscoe and Lunt (261), Clowes and Biggs (54), and Naylor (224) report the use of these methods. Coste and Andrews (59) showed that Winkler's method is not accurate in the presence of ammonium salts in quantity.

The more soluble gases are usually acid or alkaline in nature, and frequently have an oxidizing or reducing character as well. Thus several methods for analysis are usually available.

The oxidizing nature of chlorine has usually been the basis for its analysis. Roscoe (259), however, precipitated it as silver chloride.

Lewis and Keyes (179) determined hydrogen cyanide by precipitating it as the silver salt.

The alkaline character of ammonia and of the amines has been the basis for their analysis. Doijer (66), Perman (239, 240, 241), Locke and Forssall (183) and others have used this property in their vapor pressure or solubility measurements. Hydrogen chloride is naturally analyzed as an acid (see Shunke (273) and numerous others). Carbon dioxide has been determined in solution as an acid by Bohr (26) and by Kosakevich (168). Phosgene was determined by Atkinson and coworkers (7) by treatment with alkali in excess, and back-titration. Kremann and Honel (169) determined acetylene by absorption in silver nitrate solution, followed by titration of the acid liberated.

Sulfur dioxide has usually been determined by its oxidation with iodine, as was done by Fox (92). This is also true of hydrogen sulfide (Goldschmidt and Larson (102)). Briner and Perrottet (31) found the solubility of ozone in water by shaking the water with air containing ozone, and later analyzing each phase for ozone, using potassium iodide and thiosulfate.

When inert gas is used to sweep the soluble gas from solution, in order to measure its partial pressure, there is some choice of the inert gas. Air was used by Doijer (66) and by Lofman (184). Gahl (95) used the gas produced by electrolysis of water for this purpose. The ease of regulating the flow of gas and the accuracy of its measurement have led to the frequent use of this method. Dolezalek (67), McLauchlan (205), and Gaus

(99) found the partial pressures of hydrogen chloride, hydrogen sulfide, and ammonia, respectively, using electrolytic gas. Stegmuller (294) used nitrogen for the same purpose in determining the partial pressure of hydrogen iodide.

Frequent measurements have been made of the distribution of a volatile substance between two liquids. If the solubility of the gas in one of the liquids is known, the distribution data provide means of finding approximately its solubility in the other liquid. Bell and Feild (16) determined the distribution ratio of ammonia between water and chloroform, and Smith (290) worked with amines distributed between xylene and water; other systems have been investigated.

III. SOLUBILITY RELATIONSHIPS

As may be expected in any field where many investigators have published, there exists a variety of ways in which these results have been expressed. Some are in common usage; others, which are seldom used, should be the ones in common usage. The investigators also have attempted to find some correlation between the properties of the liquid and the properties of the gas that will make it possible to predict the solubility of a gas in a given liquid from known properties or from a few solubility data.

A. METHODS OF EXPRESSING SOLUBILITY

1. Bunsen absorption coefficient, α

This coefficient was proposed by Bunsen (39) and was defined by him as follows: "The volume of gas, reduced to 0° and 760 mm. pressure of mercury, which is absorbed by the unit volume of liquid under the pressure of 760 mm. is called the absorption-coefficient, or coefficient of absorption." Although not so stated in the definition, his calculations show that the pressure meant is always the partial pressure of the gas. Bunsen used the ideal gas laws to reduce the gas volume to standard conditions. As these laws are not exact, the coefficients found by different methods, i.e., physical and chemical, can be expected to differ. Thus Markham and Kobe (201) showed that in the case of carbon dioxide at 0°C., deviations of 0.7 per cent were to be expected between the two methods. This deviation will increase as the behavior of the gas departs from the ideal gas laws.

Many of the past workers have not controlled the total pressure carefully, so that the partial pressure of the gas has remained at 760 mm. Frequently the total pressure has been maintained at 760 mm. and the vapor pressure of the solvent has been neglected. Other workers have not used a partial gas pressure of 760 mm. but have corrected their actual results to this pressure by the use of Henry's law.

If the solubility is calculated according to the Bunsen coefficient, except that the amount of solvent is 1 g., the result is known as the Kuenen coefficient. For solutions this has been extended by Markham and Kobe (201) to mean the volume of gas (in cubic centimeters) at a partial pressure of 760 mm., reduced to standard conditions, dissolved by the quantity of solution containing 1 g. of solvent; thus it is proportional to gas molality. It is designated by S.

If the solubility is calculated as grams of gas dissolved per 100 cc. of solvent at the temperature of the experiment and a partial gas pressure of 760 mm., the result is known as the Raoult absorption coefficient.

2. Ostwald coefficient of solubility, L

This coefficient was defined by Ostwald (231a) as "the ratio of the volume of the absorbed gas to that of the absorbing liquid. If these are V_l and V_s, respectively, the solubility is $L = V_l/V_s$." For the reaction

$$\underset{C_l}{\text{Gas (in liquid phase)}} \rightleftharpoons \underset{C_g}{\text{Gas (in gas phase)}}$$

the Ostwald coefficient of solubility can be written as

$$L = \frac{C_l}{C_g} = \frac{V_l}{V_s} \tag{1}$$

which represents the ratio of the concentrations of gas in the liquid phase and in the gaseous phase. This is in reality an equilibrium constant, and the Ostwald coefficient is independent of the *partial* pressure of the gas as long as ideality may be assumed. However, the temperature and total pressure must be designated to fix the value of the coefficient. If the total pressure is maintained at 760 mm., the volume of gas absorbed, reduced to 0°C. and 760 mm. by the ideal gas laws, per unit volume of liquid is frequently designated as β, an absorption coefficient.

As pointed out in discussing the Bunsen coefficient, early workers frequently did not distinguish between total pressure and partial gas pressure, or did not consider the vapor pressure of the solvent. Thus in many cases results reported as α really are β. Likewise, the results found by physical and by chemical methods differ by the departure of the gas from the ideal gas laws.

3. Henry's law constant

Henry (118) stated his law as ". . . under equal circumstances of temperature water takes up the same volume of condensed gas as of gas under ordinary pressure." The modern presentation of this law for the ideal phases of a gas in equilibrium with a liquid is:

$$\begin{array}{ccc} \text{Gas (in liquid phase)} & \rightleftharpoons & \text{Gas (in gas phase)} \\ X_l \text{ or } C_l & & X_g,\ P_g, \text{ or } C_g \end{array}$$

$$X_g = K_x X_l \tag{2}$$

$$P_g = K_1 X_l \tag{3}$$

For a dilute solution of the gas:

$$P_g = K_2 C_l \tag{4}$$

$$C_g = K_c C_l \tag{5}$$

Thus it is seen that the Ostwald coefficient L, equation 1 is the reciprocal of K_c.

The Henry law constant, K_2, is a satisfactory though unwieldy method of expressing gas solubility and is the method used in the *International Critical Tables* (138a). It is noted that, the larger the value of K_1, the lower is the solubility.

Henry's law has been used by many investigators to calculate their data from an experimental pressure to a partial gas pressure of 760 mm. Over the short range usually encountered, no error is introduced. However, the worker must keep in mind that the equations given are for ideal dilute solutions and should apply any necessary corrections.

4. Interconversion of expressions for the solubility

(*a*) From the Bunsen coefficient:

$$\beta = \alpha \frac{760 - P_s}{760} \tag{6}$$

As β is the solubility coefficient measured at a total pressure of 760 mm., α is decreased from its partial pressure of 760 mm. by applying Henry's law.

$$L = \alpha \frac{T}{273} = \beta \frac{T}{273} \frac{760}{760 - P_s} \tag{7}$$

The Ostwald coefficient is calculated from α by correcting the gas volume to the temperature at which absorption was carried out.

$$S = \frac{\alpha}{\rho(1 - u)} \tag{8}$$

If the solubility is expressed per gram of solvent in a solution, the factor ρ $(1 - u)$ gives the grams of solvent per cubic centimeter of solution.

$$K_1 = \frac{17.033 \times 10^6 \rho}{\alpha M_s} + 760 \tag{9}$$

The units of K_1 are those of pressure, expressed as millimeters of mercury. Only for the very soluble gases does the constant term of 760 mm. alter appreciably the significant figures of the value for K_1 calculated in the first term of the equation, and thus it usually is neglected.

$$K_2 = \frac{17033}{\alpha} \tag{10}$$

The units of K_2 are (mm. Hg) (liters of solvent)/mole of gas (see equation 4).

(*b*) From the Ostwald coefficient, L:

$$\alpha = L \frac{273}{T} \tag{11}$$

$$\beta = L \frac{273}{T} \frac{760 - P_s}{760} \tag{12}$$

$$K_c = 1/L \tag{13}$$

(*c*) From the Henry law constant, K_1:

$$\alpha = \frac{17.033 \times 10^6 \rho}{(K_1 - 760) M_s} \tag{14}$$

The 760 mm. in the denominator may be neglected unless it is appreciable in comparison with K_1, that is, unless the number of moles of gas dissolved appreciably affects the total moles of solution.

5. *Nomenclature*

It is desirable at this point to tabulate the nomenclature recommended for the expression of gas solubilities. The equations used here have been rewritten to conform as closely as possible to this system.

A = work done in dissolving one mole of gas;
C_g = concentration, as gram-moles per liter, of gas in the gas phase;
C_l = concentration, as gram-moles per liter, of gas in the liquid phase;
C_s = concentration, as gram-moles per liter, of salt in the liquid phase;
K_1 = Henry's law constant (to fit equation 3);
L = Ostwald coefficient (defined on page 536);

M_g = molecular weight of gas;
M_s = molecular weight of solvent;
P_g = partial pressure of gas;
P_s = partial pressure of solvent;
P_t = total pressure;
R = gas constant;
S = unit gas solubility (defined on page 536);
T = temperature, degrees Kelvin;
T_c = critical temperature of gas, degrees Kelvin;
V_l = volume of the gas in the liquid phase;
V_s = volume of the solvent;
X_g = mole fraction of gas in the gas phase;
X_l = mole fraction of gas in the liquid phase.

a, b, c, k, n = arbitrary constants;
d = differential operator;
e = base of natural logarithms;
ln = natural logarithm;
log = common logarithm;
m = molality of salt;
t = temperature, degrees Centigrade;
u = decimal fraction of solute in solution.

α = Bunsen coefficient (defined on page 535);
β = solubility coefficient (defined on page 536);
γ = activity coefficient of the dissolved gas;
μ = ionic strength of the salt;
ρ = density of the solution.

Where necessary for distinguishing between the property of a solution and that of the pure solvent, the property of the solvent has been given the zero subscript.

B. VARIATION OF SOLUBILITY WITH PRESSURE

Henry (118) was the first to show the variation of gas solubility with pressure. With crude apparatus and impure gases, he performed experiments the results of which he summarized as follows: "The results of at least fifty experiments, on carbonic acid, sulfuretted hydrogen gas, nitrous oxide, oxygenous and azotic gases, with the above apparatus, establish the following general law: that, under equal circumstances of temperature, water takes up, in all cases, the same volume of condensed gas as of gas under ordinary pressure." Bunsen (39) used his method to confirm Henry's conclusions, using the carbon dioxide–water system. After this confirmation of Henry's law by Bunsen, others performed experiments on other systems to test it. His law is accepted as the normal behavior of gas–liquid systems to such an extent that frequently only the deviations from it are noted.

The less soluble gases have been found to satisfy the law well at moderate pressures. Thus Morgan and Richardson (218) found that at 25°C. the system oxygen–water satisfied Henry's law in the pressure range 175 to 760 mm. Kireev and Romanchuk (151) found that hydrogen and methane in xylene, in ethylene chloride, and in several petroleum fractions satisfied Henry's law in the temperature range -20°C. to $+40$°C. at pressures from 50 to 760 mm. Briner and Perrottet (31) experimented with ozone in water, finding Henry's law to hold. The gas phase was air containing 0.3 to 9 per cent ozone. Boyle (29) confirmed Henry's law for solutions of radon in water and in several other solvents. The concentrations were necessarily very low, and air was present. Findlay and Shen (86) found hydrogen in water to satisfy Henry's law at 25°C. over the pressure range 750 to 1400 mm.

More data are available for gases of intermediate solubility, such as carbon dioxide. Roscoe (258, 259) tested Henry's law in the case of the chlorine–water system, in which it apparently failed; however, he varied the pressure by the addition of inert gas. Perman (238) worked with several gases in water, using an extraction method which gave only the relative pressures. He found that hydrogen sulfide, carbon dioxide, and chlorine satisfied Henry's law. Findlay and coworkers (81, 82, 83, 84, 85, 86, 87) tested Henry's law as applied to carbon dioxide and nitrous oxide n water, and to carbon dioxide in aqueous solutions of alcohol and several electrolytes, at 25°C. The pressure range was 250 to 1400 mm. The results show that the law is satisfied exactly in this range. Buch (35) tested Henry's law for the carbon dioxide–water system from pressures of 1 atmosphere to 1/20,000 of an atmosphere, and found it to hold. Khanikoff and Luginin (149) worked with carbon dioxide and water to test Henry's law. At the temperature they used (15°C.) the law failed, as they stated, at 4 atmospheres pressure. Actually, deviations of 5 per cent appeared in their data at less than 2 atmospheres. Vukolov (328, 329) tested the law as applied to the solubility of carbon dioxide in chloroform and carbon disulfide. He found maximum deviations of 4.5 per cent in the pressure range 36 to 760 mm. Secenov (279) experimented with carbon dioxide and aqueous salt solutions. He concluded that at low pressures (about one-third of an atmosphere) all salt solutions follow Henry's law. Few data were presented to justify this conclusion. Stern (296) found the solubility of carbon dioxide in several organic solvents at -78° and -59°C. over the pressure range 50 to 760 mm. Using Bunsen's coefficient, he found that Henry's law was not satisfied. However, at the temperatures used the perfect gas law did not hold. When the results were expressed as Ostwald coefficients, the law held. Lewis and Keyes (179) found that the pressure of hydrogen cyanide over its aqueous solution was proportional to the molarity. Bancroft and Belden (10), working with hydrogen sulfide in aniline at 22°C. and at pressures up to 1200 mm., found Henry's law to hold.

The very soluble gases usually have not satisfied Henry's law, except at elevated temperatures or at very low concentrations. Sims (287) tested Henry's law for two systems and found that in both cases it failed at lower temperatures, but that, as the temperature increased, the deviations became less until at 50°C. the sulfur dioxide–water system satisfied the law, and at 100°C. the ammonia–water system did also. Smith and Parkhurst (291) measured the solubility of sulfur dioxide in water and in solutions of calcium and magnesium bisulfites, and found that Henry's law was satisfied at pressures up to 800 mm. in the temperature range 5° to 60°C. Quite contradictory conclusions have been reached for the ammonia–water system. Gaus (99) found that ammonia in water and in salt solutions at 25°C. satisfied Henry's law up to about 1 normal concentration with respect to ammonia (14 mm. pressure). Abegg and Reisenfeld (1) reached the same conclusions. Roscoe and Dittmar (260) found that Henry's law failed when applied to the systems ammonia–water and hydrogen chloride–water. Calingaert and Huggins (42) found that at 100°C. the ammonia–water system deviated from Henry's law even at low concentrations. They concluded that the deviations could be explained by the electrolytic dissociation of the ammonia in solution (*cf.* MacDougall (190a)). Klarmann (154), however, verified Henry's law for the same system at concentrations of 0.5 to 1/128 normal, at 0°C. Perman (238), who obtained only relative pressures, found that ammonia, hydrogen chloride, and sulfur dioxide in water did not satisfy Henry's law at room temperatures. Doijer (66) tested Henry's law for the ammonia–water system at 60°C., in the pressure range 8 to 60 mm., and found it to hold.

In cases where the gas reacts to a certain extent with salt in the solution, a modified form of Henry's law sometimes holds. Thus Hufner (137) found that nitric oxide dissolved in ferrous salt solutions in accordance with the equation

$$\alpha = a + bP_g \tag{15}$$

The pressure range 550 to 710 mm. was covered at 20°C., Neuhausen and Patrick (225, 226) proposed the use of an adsorption equation for solubility

$$\alpha = a\left(\frac{\sigma P_g}{P_0}\right)^{1/n} \tag{16}$$

in which σ is the surface tension and P_0 is the vapor pressure of the liquefied gas. It satisfied their own data for the system ammonia-water, and those of others for hydrogen chloride, carbon dioxide, and sulfur dioxide in water and in alcohols. The equation does not apply to gases above the critical temperature.

The work of Frolich and coworkers (94a) on the extension of Henry's law to high pressures is of importance in engineering calculations. The solubility of hydrogen, nitrogen, oxygen, and of methane and other hydrocarbons in water, alcohols, hydrocarbons, and heavier petroleum fractions was studied up to 200 atmospheres. They concluded that, when the gas does not form a chemical compound with the solvent, it follows Henry's law over a wide pressure range within the limits of error allowed in engineering calculations. The solubilities of these gases may be considered linear functions of the absolute pressure, the validity being dependent upon the extent to which the solute obeys the ideal gas law. However, the straight-line relationship still holds at high pressures, provided corrections are applied for deviations from the ideal gas law. A practical rule is that the solubility of a gas of the vapor type is a linear function of pressure up to one-half to two-thirds of its saturation value at that temperature.

C. VARIATION OF SOLUBILITY WITH TEMPERATURE

Bunsen (37, 38, 39, 40) applied a purely empirical equation to the data that he found for the solubilities of a number of gases in water and alcohol. The equation had the form

$$\alpha = a + bt + ct^2 \tag{17}$$

Numerous others applied similar formulae. Thus Carius (44), Than (306), and Timofeev (312) applied the same equation to their data. The constants are found by substituting experimental values at three temperatures into the equation. In most cases the solubility decreased with increase in temperature, although the solubility of hydrogen in water was constant. Winkler (336) added a term in t^3 to his equation. Henrich (117) used Bunsen's data to recalculate the constants, using the method of least squares. Wiedeman (333) showed that, while the values of a, b, and c for the different gases in water were very different, the ratios b/a and c/a were nearly the same for all gases. The same was found for alcohol. Fox (93) and Whipple and Whipple (332) applied the same type of equation to atmospheric gases in distilled water and sea water. The latter also added similar terms to include the chlorinity of sea water. These equations were all purely empirical and usually were applied over a rather limited range of temperature.

Bohr (24) proposed that at constant partial gas pressure the osmotic pressure of a dissolved gas is constant. Thus αT is constant. When he substituted values, however, he found that such was not the case, but that

$$\alpha (T - a) = k \tag{18}$$

He found a to be a constant which, for five diatomic gases (hydrogen, nitrogen; oxygen, carbon monoxide, nitric oxide) in water, was a linear function of the molecular weight. These relations seem to have been at first purely empirical. In a later article (25) he made an effort to establish a theoretical basis for his first equation. He equated the rate of solution of a gas at equilibrium to its rate of escape, and made measurements of each. He established an empirical relation for the rate of escape, which proved to give the equation that he sought. The empirical nature of his result, however, remained. Later Kofler (161) showed that the same equation fitted his results for the solubility of radon in water, over the range 0° to 75°C.

Kofler (162) stated that there is a connection between the critical temperature of a gas and its solubility in a given solvent. He plotted T/T_c for a number of gases against α in water, and found that they fell on a smooth curve.

Meyer (208) applied the equation

$$S = b + e^{-a\theta} \tag{19}$$

to the solubility of various gases in various solvents. S can be replaced by either the Bunsen or the Ostwald coefficient, with corresponding values for the constants. However, the equation fits better if it refers to the amount of gas dissolved in a unit weight of solvent. θ is a measure of the temperature, on a scale such that for a given solvent there are 100 degrees between the melting point and the boiling point. In the case of water it is the Centigrade scale. Meyer found a to be nearly the same for all gases and all solvents. For the system radon–water, the equation fits the results very well.

Jager (139), from kinetic considerations, derived the equation

$$L = e^{-\frac{A}{RT}} \tag{20}$$

Empirically he found that

$$A = a\,(1 + bt\,(1 - ct)^2) \tag{21}$$

Using values from the literature for several gases in water, he calculated the values of the constants, and found that the equation was satisfied within 2 or 3 per cent. The constant c proved to be nearly constant for all gases, and equal to the temperature coefficient of the capillarity constant of water. Szeparowicz (302) applied Jager's formula as well as that of Meyer (208) to his data for the solubility of radon in water, which he had carried up to 100°C., and found that both formulae were satisfied. Jager's equation also satisfied his data for radon in benzene.

The Clapeyron equation has been used as a basis for the derivation of several equations relating gas solubility and temperature. If the heat of solution of the gas in the liquid is constant, and is not a function of temperature over the interval used, the Clapeyron equation gives:

$$\ln \frac{L_1}{L_2} = -\frac{A}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right) \tag{22}$$

Gas solubility also may be expressed as K_1, α or S in this equation. Graphically log L is a linear function of $1/T$. This equation more frequently has been expressed in the exponential form:

$$L = ae^{-\frac{A}{RT}} \tag{23}$$

Both Tammann (304) and Lannung (172) have used it in this form, which may be compared with the equation of Jager (equation 20). Lannung, using his data for the rare gases, found that log L was a linear function of $1/T$ for the organic solvents used but not for water. There was an approximately linear relationship between A and log a. Using the data of Markham and Kobe (201) for the solubility of carbon dioxide in aqueous salt solutions from 0 to 40°C., a maximum deviation of 3 per cent existed for water at 25°C. The salt solutions all showed smaller deviations, ranging down to 0.85 per cent for 3 molal magnesium nitrate solution.

If the heat of solution of the gas in the liquid is a function of the temperature, then an equation of the form of the reaction isochor results, which may be shortened to the form used by Valentiner (322):

$$\log L = \frac{a}{T} + b \log T + c \tag{24}$$

This equation fits the solubility data for the inert gases in water.

The general rule is that the solubility of a gas in water decreases with increasing temperature. However, the Bunsen absorption coefficient for hydrogen reaches a minimum at 60°C. with no further change to 100°C.; for nitrogen the minimum is at 90°C., and for helium it is at 30°C. with a marked increase up to 100°C. If, instead of the Bunsen coefficient, the Ostwald coefficient of solubility is used, the minima in the curves come at much lower temperatures and other gases show increasing values of L with rise in temperature. With helium the minimum is below 0°C. At pressures up to 1000 atmospheres this minimum in the solubility isobar becomes quite apparent; it has been studied by Wiebe, Gaddy, and coworkers (332a). Above 200 atmospheres partial pressure, carbon dioxide shows an increase in the absorption coefficient in water. In non-aqueous solvents an increase in solubility with rise in temperature is a common phenomenon. Lannung, Horiuchi, and others have shown the increased solubility with rise in temperature of the relatively insoluble gases. The effect is usually small at atmospheric pressure but may become quite large at higher pressures, as indicated by the sevenfold increase in the solubility of hydrogen from 0° to 100°C. in liquid ammonia at 1000 atmospheres (332a).

With the Ostwald coefficient for a relatively slightly soluble gas, the concentrations of the gas in the liquid and the gaseous phases should approach the same value as the solvent approaches the critical temperature, or L approaches 1. Thus, all such gases should show a minimum in the solubility isobar at some definite temperature, from below 0°C. for helium to higher temperatures for other gases. Horiuchi took the critical temperature of the solvent (T_K) into consideration and found, by plotting log L against T_K/T that the solubility lines for a particular gas in a number of solvents fall closer together. However, it may be concluded that the minimum in the solubility isobar is not a peculiar property of the gas or of the solvent, but is a phenomenon of the mixture and may be predictable from known properties of mixtures.

D. VARIATION OF SOLUBILITY WITH CONCENTRATION

The effect of the addition of another solute on gas solubility has frequently been investigated, and several formulae have been proposed to express this effect. In the following discussion "solute" will refer to the soluble substance whose concentration is the independent variable.

Raoult (254) found that, when ammonia dissolved in water and aqueous salt solutions, the gas solubility was a linear function of the solute concentration. This is the simplest relation that could be desired. Hufner (136), who worked with hydrogen and nitrogen in aqueous solutions of organic compounds, also found a linear relationship. He further found that, in comparing some compounds with each other, the solubility lowering was proportional to the molar concentration, the proportionality factor being the same for different solutes. In the case of other compounds, however, he found that the lowering produced was proportional to the weight of solute per volume of solution. His experimental results were not very good, but seem to confirm the above relation fairly well. Hudson (135) found that the solubility of sulfur dioxide in aqueous potassium chloride solutions bore a linear relationship to the salt concentration. Such was not the case when sodium sulfate was the solute, however. Konovalov

(165) found that the pressure of ammonia dissolved in aqueous solutions of copper and silver salts satisfied the formula

$$P_g = P_{g_0}(C_g - aC_s) \tag{25}$$

in which C is the ammonia concentration and a is 2 for silver and 4 for copper. For other salt solutions, he found (166) that $(P_0 - P_{g_0})/C_s$ increased as C_s increased. Abegg and Reisenfeld (1) found that the effect of salt on the pressure of ammonia was linear with the salt concentration, but that the solubility was not linear with salt concentration. Their results thus did not agree with those of Konovalov.

Secenov (278, 279) introduced an equation that has been used frequently. He stated the hypothesis: "If equal quantities of the same salt are added to equal volumes of different aqueous solutions, the percentage reduction in solubility will be the same in both." This hypothesis was poorly supported by one experiment. If it is accepted, however, there results his equation:

$$\alpha = \alpha_0 \epsilon^{-kC_s} \tag{26}$$

Another equation frequently found in the literature is referred to as Jahn's equation (140), but appears to have been published first by Gordon (104). It is

$$\alpha_0 - \alpha = kC_s^{\frac{2}{3}} \tag{27}$$

The two-thirds power brings in some surface relation. Gordon's data were not very good and did not satisfy the equation very well.

From thermodynamic considerations Roth (262) derived the relation that the molecular concentrations of gas in pure water and in dilute solutions of inert substances, at the same temperature and partial pressure, are the same. Nitrous oxide in aqueous solutions of urea, oxalic acid, and glycerol satisfied his theory fairly well. Solutions of sodium chloride and phosphoric acid did not, however, but did satisfy Jahn's equation. The glycerol solution did not satisfy Jahn's equation, since the reduction in α was nearly proportional to the solute concentration.

Steiner (295) used his data on hydrogen in aqueous salt solutions to test Secenov's equation, and found that it did not hold well. Rothmund (263) showed that, from Secenov's equation

$$\log \frac{\alpha_0}{\alpha} = \log \frac{L_0}{L} kC_s \tag{28}$$

and in dilute solution,

$$\frac{L_0}{L} = kC_s \tag{29}$$

Tammann (304), from relations that he found empirically, arrived at Secenov's formula. Åkerlöf (3) found the values of k in Secenov's formula for helium and argon in aqueous salt solutions, but had insufficient data to make any verification. Kiss, Lajtai, and Thury (152), who worked with carbon dioxide and hydrogen sulfide in aqueous solutions of organic substances, found that neither Secenov's nor Jahn's equation was satisfactory. Calvet (43) made an effort to justify Secenov's formula from experimental data on the mobility of molecules in a solvent.

The data of Markham and Kobe (201) on carbon dioxide and nitrous oxide in water and in numerous aqueous salt solutions, over a wide range of concentration, satisfied none of these equations. They proposed the equation

$$\frac{S}{S_0} = am + \frac{1}{1 + bm} \tag{30}$$

This equation fits their data within the experimental error of 0.2 per cent. They showed further (201a), from data on sulfuric acid and perchloric acid solutions, that the equation held through a minimum in the solubility curve. However, the agreement does not extend to the maximum in the curve. This equation is that of a hyperbola, with a vertical asymptote at $m = -1/b$, the other asymptote having a slope of S_0a (figure 5). Most solubility curves are in the region of this equation in which the slope is negative, before the minimum is reached.

Braun (30) determined the solubility of nitrogen and of hydrogen in aqueous solutions. He was satisfied with Roth's formula for solutions of urea and propionic acid, although in 10 per cent urea solutions there were deviations from the formula of as much as 10 per cent. For sodium and barium chlorides he used Jahn's equation, in which the maximum deviation was about 2 per cent. Levi (177) experimented with solutions of potassium iodide and urea in methanol, and reported that Jahn's formula held, while Roth's also held for the urea solution. Locke and Forssall (183) used Jahn's formula in their determination of the amount of ammonia in the copper ammonia complex in solution. Knopp (157) used his data on hydrogen and nitrous oxide in aqueous salt solutions to test Jahn's formula as well as Roth's. Neither was satisfactory. His values, however, satisfied another formula derived by Jahn (140), namely:

$$\log \frac{C_l}{C_{l_0}} = C_s(1 - f)(a + fb) \tag{31}$$

in which f is the degree of dissociation as found by conductivity, and C_l and C_{l_0} are the molecular concentrations of gas in solution and in pure water, respectively. If a is small and f is nearly constant, then the above formula reduces to

$$\log \frac{C_l}{C_{l_0}} = C_s k \tag{32}$$

This formula was not satisfactory.

Usher (320) worked with carbon dioxide and aqueous solutions of nonelectrolytes. He claimed a high degree of accuracy for his data, which probably were much more of a test of the formulae than most that had been used. He felt that the formulae of Jahn and Roth were of little value, since he showed that frequently the deviation from theory is greater than the effect to be explained.

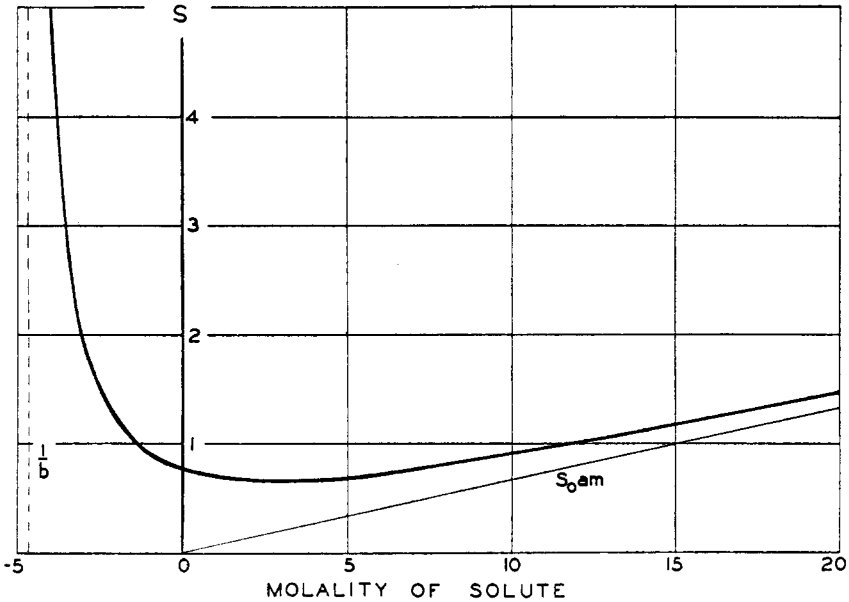

FIG. 5. Solubility isotherm for equation 30. Constants from carbon dioxide in aqueous sulfuric acid.

Philip (243) made two suggestions intended to bring into better agreement the solubility of gases in solutions: first, that all solubilities be expressed on a basis of 1000 g. of solvent, i.e., water in an aqueous solution, rather than on a volume of solution; and second, that the loss of solvent to solvate the solute accounts for the reduction in gas solubility when the weight basis is used. MacArthur (190) used his data on oxygen solubility to find the degree of hydration of a number of salts and of sucrose. The values that he found were consistent with those found by other methods. Manchot (198), from data on the solubility of nitrous oxide and acetylene, calculated the degree of hydration of a number of salts. The two gases gave similar results. Gaus (99) found an apparent connection between the atomic volume of the cation of a salt and the effect of the salt on the partial pressure of ammonia. Usher (320) abandoned the hydration hypothesis to account for the change in solubility in solutions of non-electrolytes, since in several cases the solubility was increased by the addition of solute. Perman (239, 240, 241) found that the addition of urea to an aqueous ammonia solution caused little change in the pressure of ammonia, while mannitol and several salts caused somewhat more change.

Jones, Lapworth, and Lingford (145) used the Duhem equation to express their results on the partial pressure of hydrogen chloride over water–alcohol solutions. Intermediate empirical equations enabled them to effect the integration, giving the result

$$\log P_g = a \log y + by + cy^2 + k \qquad (33)$$

in which y is the moles of hydrogen chloride per mole of alcohol. They expressed the constants as functions of the water content of the solution in a purely empirical way.

Randall and Failey (253), using examples from the literature, found that plots of $(\log \gamma)/\mu$ against $\sqrt{\mu}$ gave straight lines. In most instances these lines were horizontal. Markham and Kobe (201) confirmed this relation for their data on the solubility of carbon dioxide and nitrous oxide in aqueous salt solutions. They further showed, from differentiation of equation 30, that

$$\frac{\ln \gamma}{m} = b - a \qquad (34)$$

at low concentrations, in agreement with Randall and Failey.

E. GENERAL RELATIONSHIPS

1. *Additive effect of ions*

Steiner (295) found that, in dilute solutions of several strong electrolytes, the reduction in the solubility of hydrogen was an additive function of the ion concentrations. Van Slyke and Sendroy (326) found the same result for carbon dioxide and hydrogen in aqueous solutions of alkali chlorides, lactates, and phosphates. Markham and Kobe (201) found a similar result for the solubility of carbon dioxide in aqueous solutions of the chlorides and nitrates of sodium and potassium, in concentrations up to 1 molal.

2. Specific effect of solute

Rothmund (263), Euler (78), and McLauchlan (205) arranged numerous salts in the order of the percentage lowering of the solubility that they produced, and all found the same order. The "gases" used were phenylthiourea, ethyl acetate, and hydrogen sulfide, respectively. Reisenfeld (256) found that the equivalent solubility lowering of all salts (as they affected ammonia) was the same, barring specific chemical action. This question of chemical action has come up repeatedly. Secenov (279) stated that all salts take an active part in the absorption of carbon dioxide. Others have tried to distinguish between the chemical and the physical effects of solutes on the solubility of a gas. Rothmund (263) in this connection found that $(L_0 - L)/L_0$ was independent of temperature, and from this relation and the Clapeyron equation showed that the heat of solution of phenylthiourea in water and in salt solution was the same, indicating the absence of chemical reaction. Bell (15), from his data on the solubility of hydrogen sulfide, ammonia, and hydrogen chloride in a number of solvents, believed that the solubility depended on individual properties of the solvent molecule.

Drucker and Moles (71) plotted several properties of the solution against the composition of aqueous glycerol solutions. The properties were: heat of solution, coefficient of expansion, surface tension, specific heat, and the solubility of nitrogen and hydrogen. These properties deviated from the straight line that would result in the case of perfect solutions. The point of maximum deviation of the properties fell at the same composition, except in the case of solubilities. From the fact that these fell at different points, Drucker and Moles concluded that solubility depends on the chemical properties of the solvent, rather than on physical properties.

Skirrow (289) found the solubility of carbon monoxide in several mixtures of organic solvents. In several cases the solubility was an additive function of the solvent concentrations, but usually it was not. Some solutions showed a minimum in the surface tension–composition curve and a maximum in the solubility curve at nearly the same concentration. Christoff (49) found the same result in the solubility of carbon dioxide in several solutions.

3. Effect of surface tension

Christoff (50) measured the solubility of several gases in ether, which has an extremely low surface tension, to show that some relation existed between the two. Gases proved to be more soluble in ether than in other solvents with which comparison was made. Uhlig (318) considered the energy change in transferring a gas molecule from the gas phase into the liquid against the force of surface tension, and derived the equation

$$\ln L = \frac{-4\pi r^2 \sigma + E}{KT} \tag{35}$$

in which r is the molecular radius of the gas, E is the interaction energy, σ is the surface tension of the solution, and K is Boltzmann's constant. E and r, found from solubility data, checked the same quantities found by other means. Eley (75), in the first of a series of articles, proposed the mechanism of cavity formation, after which the gas molecule enters the cavity. At the temperature of maximum density, the energy and entropy of cavity formation are zero, but increase with increase in temperature.

Sisskind and Kasarnovskii (288) measured the solubility of argon in various organic solvents, including several homologous series. Tables were given of the solubility of the gas, the molecular volume of the liquid, the surface tension, the dipole moment, and the polarizability of the solvents. The solubility, the surface tension, and the molecular volume were in substantially the same order.

4. Effect of viscosity

Winkler (339, 342) proposed an equation relating solubility to viscosity. Elsewhere this equation is credited to Than. This equation can also be considered as a relation between solubility and temperature, since the viscosity change with temperature is the variable. The equation is

$$\frac{\alpha_1 - \alpha_2}{\alpha_1} = \frac{Z_1 - Z_2}{Z_1} \frac{\sqrt[3]{M_g}}{k} \tag{36}$$

in which Z is the viscosity and the subscripts refer to values at two temperatures. For five fixed diatomic gases in water, K proved to be nearly equal to the cube root of 54, three times the molecular weight of water. The data that Winkler gave showed remarkable agreement with the equation. Thorpe and Rodgers (309) stated that Winkler's conclusions must be changed to: "For the same gas, the decrease in solubility (not percentage decrease) is proportional to the corresponding decrease in viscosity; and further, for any gas, the factor of proportionality is greater for a greater molecular weight, but no simple relation exists."

Winkler stated that with the increase of the volume of the solvent with temperature, the coefficient should increase, but that the decrease in the viscosity should cause a tendency for the coefficient to decrease. The result should be a minimum in the temperature–solubility curve. Such a minimum has been observed in several cases, as well as a positive temperature coefficient of solubility in others.

5. Homologous compounds

Just (147) arranged a number of organic solvents in the order of their ability to dissolve each of several gases, and found the order to be nearly the same for all the gases that he used. When so arranged, the compounds of each series having a common reactive group fell in the order of their molecular weights, with the solubility decreasing as the molecular weight increased. He found the solubility ratios of two gases in the same liquid to be of the same order of magnitude for all liquids. Horiuchi (130, 131, 132, 133) found this to be true only for low-boiling gases. McDaniel (204) found that the solubilities of three gaseous hydrocarbons increased in the same order in the liquids with which he experimented. Sander (266) found that in homologs the solubility of carbon dioxide decreased with increasing molecular weight. Korosy (167) found that different gases in one solvent fit approximately a formula equivalent to

$$L = a + bT_c \tag{37}$$

a and b being constants of the solvent, and b nearly the same for all solvents. Markham and Kobe (201) used Duhring lines to express the solubility of carbon dioxide and nitrous oxide in various aqueous salt solutions. They showed that, for any salt, if the concentration at which the gas solubility was the same as for a certain concentration of a reference salt was plotted against the concentration of the reference salt, a straight line resulted (figure 6). Deviations, though greater than experimental error, were still not over several per cent.

6. Effect of the compressibility of the liquid

Ritzel (257) related gas solubility to the compressibility of the liquid. He derived the relation:

$$L = \frac{P_0 B}{\delta} \tag{38}$$

in which δ is the coefficient of dilatation, and B is the compressibility of the liquid. Accepting Ångström's (4) conclusion that the ratio of the coefficients of dilatation for two gases in one liquid is independent of the liquid, it follows that the ratio of the solubility of two gases in one liquid is nearly the same for all liquids, as Just found. Kofler (162) arranged a number of salts in the order of their ability to decrease the solubility of phenylthiourea in water, and found nearly the same ratio as that of their ability to decrease compressibility. He plotted various properties of aqueous sulfuric acid solutions against concentration,—e.g., viscosity, compressibility, volume contraction on mixing, conductivity, and the solubility of nitrogen and hydrogen. The curves of solubility and compressibility were similar. Horiuchi (134) found a relationship between partial molal volume, the solubility of a gas, and the compressibility of a liquid. It was satisfactory for hydrogen in carbon tetrachloride, but not for other systems in which comparison was made.

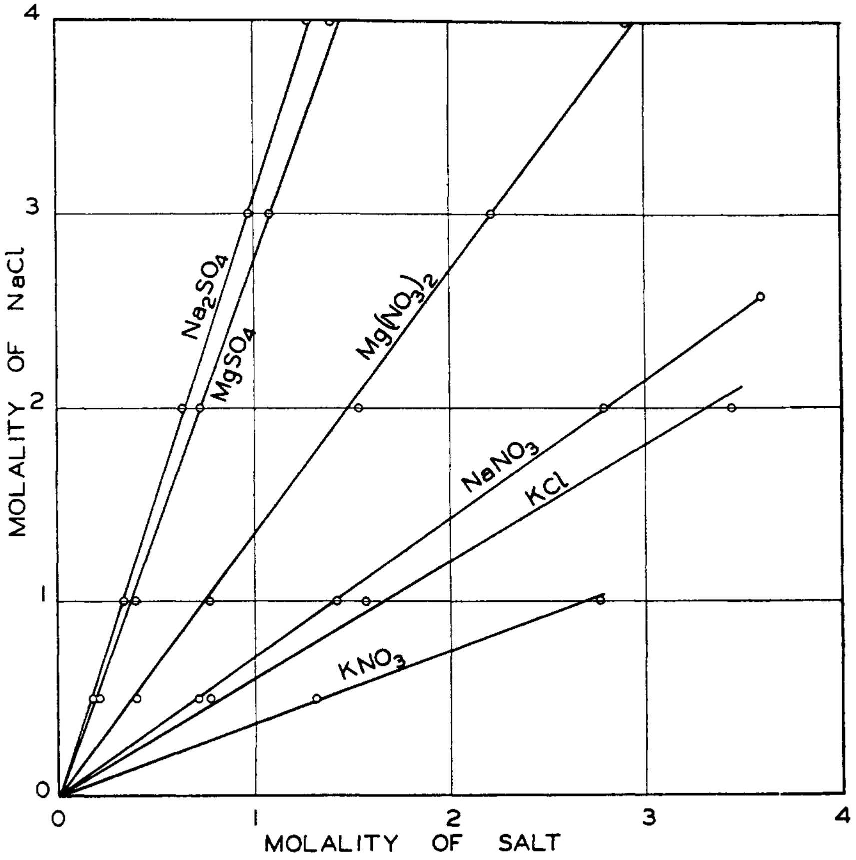

FIG. 6. Duhring lines of equal gas solubility, using sodium chloride solution as the reference solution.

7. *Relationships from Raoult's law*

Dolezalek (68) derived relationships of solubility based on Raoult's law. From his expression

$$L = \frac{X_l}{1 - X_l}\frac{1000\rho}{C_g M_g} \tag{39}$$

he calculated solubility in several organic liquids, and checked Just's data within 5 to 20 per cent. This relation, in the case of gases the solubility of which is small, readily reduces to Just's conclusion that the ratio of the solubilities of two gases in the same liquid is independent of the liquid. In the case of nitrogen and carbon monoxide, Dolezalek verified Just in this respect. This pair is the only satisfactory one that Just found for this comparison, and these gases are isoelectronic. Stern (297) used Just's data for the verification of Dolezalek's equation, and showed that, if association of the liquid were used to explain discrepancies, as Dolezalek suggested, then unreasonable degrees of association were found. Schulze (272) tested Dolezalek's theory, using values from the literature for the solubility of radon in organic liquids. The curves of solubility that he found thus were of the same general form as the experimental curves, but far off in values.

8. *Effect of the internal pressure of the liquid*

Euler (78) suggested that the decrease in gas solubility caused by a solute was due to the increase in internal pressure in the solution. He used the equivalent contraction accompanying solution as a measure of the internal pressure increase, and found that, in the case of the salts that he used, the lowering of the solubility of ethyl acetate was in the same order as the equivalent contraction.

Geffcken (100) mentioned the possible relation between gas solubility and the internal pressures of gas and liquid. Hildebrand (121) calculated the theoretical solubilities of several gases, based on Raoult's law, and compared the values with those found in the literature for a number of solvents. He stated that deviations from the value predicted by Raoult's law were large or small depending on the difference in the internal pressures of the gas and liquid, except in the case of highly polar pairs. Taylor and Hildebrand (305) used experimental data of their own on chlorine in several solvents in proceeding with the same idea. Kunerth (171) questioned the value of this theory, but, within the limitations proposed by Hildebrand himself, the data support the theory. Hildebrand (122, 123) restated his theory and its limitations as follows: "Raoult's law will be obeyed by any liquid mixture in which the internal forces of attraction and repulsion do not change with changing composition of mixture. When this condition holds the solubility of a gas may be calculated approximately from its saturation pressure, and the solubility of a solid from its melting point and heat of fusion. The above condition can exist only (a) when the components in the pure liquid phase have the same internal pressures; (b) when the different molecules are relatively symmetrical or non-polar; (c) when the tendency to form compounds is absent. Differences in either internal pressure or polarity alone produce approximately proportional positive deviations from Raoult's law and decreased solubilities. . . ."

Hamai (109, 110) found that his data on the solubility of hydrogen chloride in several organic halogen compounds did not correlate with their internal pressures or polarity, but varied in the same order as their total bond energy.

9. *Miscellaneous*

Homfray (128) found the solubility of carbon dioxide in *p*-azoxyphenetole in both the liquid crystal and the anisotropic states, and showed that the state of the solvent had considerable effect on the solubility.

Sackur (265) used gas solubility data to find the osmotic pressure of the gas in the liquid phase, and found that the results so calculated agreed with the experimental within a few per cent.

Bell (14) found that for the solubility of gases a linear relation existed between the energy and the entropy of solution of different solutes in the same solvent.

IV. SOLUBILITY DATA

In this section reference is made to all available data for a particular gas, giving the solvent employed, the range of temperature and pressure, and the reference to the original literature. An effort has been made to indicate the probable reliability of the data, on the basis of the method employed, the completeness of the data, and the consistency of the results among themselves. Comparison among the various workers in general has not attempted. Numbers ranging from 4 to 1 are found in the column headed "Value", in which a value of 4 indicates data in which considerable reliance can be placed, although comparison of these values as given by different experimenters reveals discrepancies in some cases. The smaller values indicate less reliable data; number 1 indicates data which are little more than qualitative.

TABLE 1

Solubility data

GAS	SOLVENT	PRESSURE	TEMPERATURE	VALUE	REFERENCES
			°C.		
A. Inert gases:					
1. Helium	Water	Atm.	2-30	4	(41)
		Atm.	25-75	4	(332a)
		Atm.	25	4	(3)
		Atm.	15-37	4	(172)
		1 to 6 atm.	38	4	(115)
		Atm.	0-50	3	(5)
		Atm.	0-50	2	(77)
		Atm.	18	1	(251)
			0-45	1	(321, 323)
	Methanol	Atm.	15-37	4	(172)
	Cyclohexanol	Atm.	25-37	4	(172)
	Benzene, cyclohexane	Atm.	15-37	4	(172)
	Acetone	Atm.	15-25	4	(172)
	Blood	1 to 6 atm.	38	4	(115)
	Aqueous solutions:				
	Solute: KCl, NaCl, LiCl, LiI, NaNO$_3$, HClO$_4$	Atm.	25	4	(3)
2. Neon	Water	Atm.	15-37	4	(172)
		Atm.	0-50	3	(5)
		Atm.	0-45	1	(321, 323)
	Methanol	Atm.	15-37	4	(172)
	Cyclohexanol	Atm.	25-37	4	(172)
		Atm.	25	4	(46)
	Benzene, cyclohexane	Atm.	15-37	4	(172)
	Acetone	Atm.	15-25	4	(172)
3. Argon	Water	Atm.	25	4	(3)
		Atm.	15-37	4	(172)
		Atm.	0-50	3	(5)
		Atm.	0-50	2	(77)
		Atm.	12	1	(250)
	Sea water	Atm.	2-25	2	(249)
	Methanol	Atm.	15-37	4	(172)
	Cyclohexanol	Atm.	25-37	4	(172)
		Atm.	26	4	(46)
	Benzene, cyclohexane, acetone	Atm.	15-37	4	(172)
	Chloroform		Room	2	(167)
	Paraffin oil		32		(222)
	Aqueous solutions:				
	Solute: KCl, NaCl, LiCl, NaNO$_3$, CaCl$_2$, SrCl$_2$, BaCl$_2$, MgCl$_2$, AlCl$_3$	Atm.	25	4	(3)
4. Krypton	Water	Atm.	0-50	3	(5)
		Atm.	Room	3	(180)
	Methanol	Atm.	Room	3	(180)
	Ethanol	Atm.	Room	3	(180)
			Room	2	(167)
	Amyl alcohol	Atm.	Room	3	(180)
	Glycerol	Atm.	Room	3	(180)
			Room	2	(167)
	Cyclohexanol		Room	2	(167)
	Acetic acid		Room	2	(167)

TABLE 1—*Continued*

GAS	SOLVENT	PRESSURE	TEMPERATURE	VALUE	REFERENCES
			°C.		
A. Inert gases—Continued:					
4. Krypton.....	Butyl acetate, butyl phthalate, tricresyl phosphate, acetone, tetralin, bromoform		Room	2	(167)
	Benzene	Atm.	Room	3	(180)
	Toluene, xylene		Room	2	(167)
	Petroleum fractions	Atm.		3	(180)
	Chloroform, carbon tetrachloride		Room; 0	2	(167)
	Calcium chloride solution (aqueous)		Room	2	(167)
5. Xenon........	Water	Atm.	0-50	3	(5)
B. Elementary gases:					
6. Hydrogen.....	Water	750 to 1400 mm.	25	4	(86)
		Atm.	25	4	(71)
		Atm.	0-60	4	(339)
		Atm.	0-100	4	(337)
		Atm.	0-20	3	(312)
		Atm.	20	3	(157)
		Atm.	20-25	3	(147)
		Atm.	15	3	(221)
		700 mm.	20	3	(136)
		Atm.	20	3	(49)
		Atm.	25	3	(100)
		Atm.	5-25	3	(30)
		Atm.	0-20	2	(37, 38, 39, 40)
		Atm.	20-25	2	(207)
		Atm.	25	2	(212)
		Atm.	20	2	(185)
		Atm.	15-80	2	(275)
		Atm.	11-19	2	(295)
		900 to 8200 mm.	20-25	2	(45)
	Methanol	Atm.	20-25	3	(147)
	Ethanol	Atm.	0-50	4	(202)
		Atm.	20-25	3	(147)
		Atm.	0-20	3	(312)
		Atm.	20	3	(49)
		Atm.	0-25	2	(44)
		Atm.	0-20	2	(40, 117)
	Propyl alcohol	Atm.	25	3	(147)
	Amyl alcohol	Atm.	20-25	3	(147)
	Cyclohexanol	Atm.	25	4	(46)
	Acetic acid	Atm.	20-75	4	(202)
		Atm.	20-25	3	(147)
	Methyl acetate	Atm.	−80–+40	4	(133)
	Ethyl acetate	Atm.	0-40	4	(202)
		Atm.	20-25	3	(147)
	Isobutyl acetate, amyl acetate	Atm.	20-25	3	(147)
	Benzene	Atm.	10-40	4	(202)
		Atm.	7-63	4	(133)
		Atm.	20-25	3	(147)
	Toluene, xylene	Atm.	20-25	3	(147)

TABLE 1—*Continued*

GAS	SOLVENT	PRESSURE	TEMPERATURE	VALUE	REFERENCES
			°C.		
B. Elementary gases —Continued:					
6. Hydrogen....	Xylene	50 to 760 mm.	−20–+40		(151)
	Chloroform	Atm.	0–25	4	(202)
		Atm.	20–25	3	(147)
	Carbon tetrachloride	Atm.	0–60	4	(133)
	Chlorobenzene	Atm.	−40–+80	4	(133)
	Ethylene dichloride	50 to 760 mm.	0		(151)
	Acetone	Atm.	−80–+40	4	(133)
		Atm.	20–25	3	(147)
	Diethyl ether	Atm.	0–15	4	(50)
		Atm.	−80–+20	4	(133)
	Carbon disulfide, aniline, nitrobenzene	Atm.	20–25	3	(147)
	Petroleum fractions	Atm.	20	3	(49)
		Atm.	10–20	2	(101)
		50 to 760 mm.	−20–+40		(151)
			32		(222)
					(88)
	Corn oil, lard, cottonseed oil		23–45		(327)
	Sulfuric acid	Atm.	20	3	(49)
	Aqueous solutions:				
	Solute:				
	Glycerol	Atm.	25	4	(71)
		Atm.	15	3	(221)
	Ethanol	Atm.	20	2	(185)
	Sucrose	Atm.	15	3	(221)
			11–19	2	(295)
	Dextrose	Atm.	15	3	(221)
		700 mm.	20	3	(136)
	Chloral hydrate	Atm.	15	3	(221)
		Atm.	20	3	(157)
	Urea	Atm.	5–25	3	(30)
		700 mm.	20	3	(136)
	Arabinose, alanine, acetamide, glycocoll, levulose	Atm.	20	3	(136)
	Propionic acid	Atm.	5–25	3	(30)
	Acetic, chloroacetic, hydrochloric, nitric, and sulfuric acids; NaOH, KOH	Atm.	25	3	(100)
	KNO_3, NH_4NO_3, $NaNO_3$, KCl, NaCl	Atm.	20	3	(157)
	LiCl, NaCl, KCl, $CaCl_2$, $NaNO_3$, KNO_3, $Al(NO_3)_3$, $MgSO_4$, Na_2SO_4, $ZnSO_4$, Na_2CO_3, K_2CO_3		11–19	2	(295)
	Acidified sodium sulfate solution	Atm.	25	4	(158)
7. Nitrogen......	Water[*]	Atm.	20	4	(300)
		Atm.	25	4	(230)
		Atm.	37.5	4	(105)
		1 to 6 atm.	38	4	(114)
		Atm.	23	4	(55)
		Atm.	38	4	(325)

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[*] References 58 and 2 give many results from the literature.

TABLE 1—*Continued*

GAS	SOLVENT	PRESSURE	TEMPERATURE °C.	VALUE	REFERENCES
B. Elementary gases —Continued:					
7. Nitrogen.....	Water	600 to 720 mm.	0-50	4	(93)
		Atm.	5-25	3	(30)
		Atm.	25	3	(71)
		Atm.	0-100	3	(338)
		Atm.	0-60	3	(339)
		700 mm.	20	3	(136)
		Atm.	20-25	3	(147)
		Atm.	15	3	(221)
		Atm.	15	3	(112)
		Atm.	Room	3	(27)
		Atm.	0-25	2	(37, 38, 39, 40)
		900 to 8200 mm.	20-25	2	(45)
		Atm.	20	2	(49)
		Atm.	0-20	2	(65)
		Atm.	20-25	2	(207)
			0-14	2	(242)
	Sea water*	Atm.	2-25	4	(249)
		Atm.	2-25	4	(2)
			0-28	4	(93)
	Methanol	Atm.	20-25	3	(147, 177)
	Ethanol	Atm.	20-25	3	(147)
		Atm.	0-25	2	(40)
		Atm.	0-25	2	(44)
		Atm.	20	2	(49)
		Atm.	20-25	2	(207)
	Amyl alcohol	Atm.	20-25	3	(147)
	Benzene	Atm.	10-60	4	(133)
		Atm.	20-25	3	(147)
		Atm.	20-25	2	(207)
	Diethyl ether	Atm.	0	4	(130)
		Atm.	0-15	4	(50)
		Atm.	−80-+20	4	(133)
	Acetone	Atm.	−80-+40	4	(133)
		Atm.	20-25	3	(147)
	Carbon tetrachloride	Atm.	−20-+60	4	(133)
	Chlorobenzene	Atm.	−40-+80	4	(133)
	Methyl acetate	Atm.	−80-+40	4	(133)
	Carbon disulfide	Atm.	20-25	3	(147)
	Toluene, xylene	Atm.	20-25	3	(147)
	Chloroform, acetic acid	Atm.	20-25	3	(147)
	Ethyl acetate, isobutyl acetate, amy acetate	Atm.	20-25	3	(147)
	Aniline	Atm.	20-25	3	(147)
		Atm.	20-25	2	(207)
	Nitrobenzene	Atm.	20-25	3	(147)
		Atm.	20-25	2	(207)
	Petroleum fractions	Atm.	20	2	(49)
		Atm.	2-25	2	(170)
		Atm.	10-20	2	(101)
			32		(222)
	Cottonseed oil		23-25		(327)
	Corn oil, lard		23-25		(327)
	Blood, blood fluids	Atm.	38	4	(325)

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* References 58 and 2 give many results from the literature.

TABLE 1—*Continued*

GAS	SOLVENT	PRESSURE	TEMPERATURE	VALUE	REFERENCES
			°C.		
B. Elementary gases—Continued:					
7. Nitrogen.....	Blood, blood fluids	1 to 6 atm.	38	4	(114)
		Atm.	23	4	(55)
		Atm.	37.5	4	(105)
	Liquid oxygen		−190	1	(76)
	Liquid sulfur dioxide	100 to 700 mm.	−60–−20	2	(69)
	Sulfuric acid	Atm.	20	2	(49)
	Aqueous solutions:				
	Solute:				
	NaCl, Na$_2$CO$_3$	Atm.	38	4	(325)
	H$_2$SO$_4$	Atm.	Room	3	(27)
	Urea, propionic acid, BaCl$_2$, NaCl	Atm.	5-25	3	(30)
	Glycerol, isobutyric acid	Atm.	25	3	(71)
	Glycerol, chloral hydrate	Atm.	15	3	(112)
	Sucrose, dextrose, glycerol, chloral hydrate	Atm.	15	3	(221)
	Urea, arabinose, glycocoll, acetamide, dextrose, levulose, alanine	700 mm.	20	3	(136)
	Acidified sodium sulfate solution	Atm.	25	4	(158)
	Dyes		14.5	2	(331)
	Non-aqueous solutions:				
	Methanol solutions of urea, KI				(177)
8. Oxygen.......	Water	175 to 760 mm.	25	4	(218)
		Atm.	25	4	(216)
		Atm.	0-60	4	(339)
		Atm.	0-100	4	(338)
		550 to 800 mm.	0-50	4	(93)
		Atm.	25	3	(100)
		Atm.	20	3	(49)
		Atm.	0-20	3	(65)
		Atm.	15	3	(221)
			0-14	3	(242)
		Atm.	6-12	3	(312)
		Partial pressure in normal air	0-30	3	(332)
		Partial pressure in normal air	0-30	3	(335)
		Partial pressure in normal air	25	3	(190)
		900 to 8200 mm.	20-25	2	(45)
		Atm.	0-25	2	(37, 38, 39, 40)
		Atm.	20	2	(185)
		Atm.	18	2	(192)
		Atm.	20-25	2	(207)
		Atm.	15-80	2	(275)
		Atm.	5-25	1	(281)
			0-100		(70)
					(47)
	Sea water		0-28	3	(93)

TABLE 1—*Continued*

GAS	SOLVENT	PRESSURE	TEMPERATURE	VALUE	REFERENCES
			°C.		
B. Elementary gases —Continued:					
8. Oxygen......	Sea water	Partial pressure in normal air	0-30	3	(332)
			2-35	From literature	(2)
	Ethanol	Atm.	6-23	3	(312)
		Atm.	20	3	(49)
		Atm.	20-25	2	(207)
		Atm.	0-25	2	(40)
		Atm.	0-25	2	(44)
	Cyclohexanol	Atm.	25	4	(46)
	Methyl acetate	Atm.	−80–+40	4	(133)
	Benzene	Atm.	10-60	4	(133)
		Atm.	25	4	(217)
	Carbon tetrachloride	Atm.	0-60	4	(133)
		Atm.	20-25	2	(207)
	Acetone	Atm.	−80–+40	4	(133)
		Atm.	20-25	2	(207)
	Diethyl ether	Atm.	−80–+20	4	(133)
		Atm.	0-15	4	(50)
	Chlorobenzene	Atm.	−40–+80	4	(133)
	Petroleum fractions	Atm.	20	3	(49)
		Atm.	2-25	2	(170)
		Atm.	10-20	2	(101)
		Atm.	20		(194)
	Sulfuric acid	Atm.	20	3	(49)
	Cottonseed oil, corn oil, lard		23-45		(327)
	Blood	Partial pressure in normal air	39	1	(138)
	Liquid sulfur dioxide	100 to 700 mm.	−60–−20	2	(69)
	Aqueous solutions:				
	Solute:				
	Sucrose, dextrose, glycerol, chloral hydrate	Atm.	15	3	(221)
	Ethanol	Atm.	20	2	(117)
	H_2SO_4	Atm.	Room	2	(27)
	HCl, HNO_3, H_2SO_4, NaCl, K_2SO_4, KOH, NaOH	Atm.	15-25	3	(100)
	Sucrose, LiCl, NaCl, KCl, RbCl, CsCl, NaBr, KBr, KI, KNO_3, Na_2SO_4, K_2SO_4, $MgCl_2$, $CaCl_2$, $BaCl_2$	Partial pressure in normal air	25	3	(190)
	NH_4Cl	Partial pressure in normal air	25	1	(190)
		Atm.	25	3	(59)
	KCN	Atm.	18	2	(192)
	Acidified sodium sulfate solution	Atm.	25	4	(158)
	Gas-main condensate		0-100		(70)
9. Ozone.........	Water	About 2 to 70 mm.	0-60	3	(31)
			0-60	1	(193)
	Carbon tetrachloride	About 70 mm.	−12-0	3	(32)

TABLE 1—*Continued*

GAS	SOLVENT	PRESSURE	TEMPERATURE	VALUE	REFERENCES
			°C.		
B. Elementary gases—Continued:					
9. Ozone	Carbon tetrachloride, chloroform, acetic acid	About 50 mm.	15–18	3	(89)
	Aqueous solutions:				
	Solute:				
	NaCl	About 2 to 70 mm.	0–60	3	(31)
	H_2SO_4		0–20	2	(188)
10. Chlorine	Water	Atm.	20	4	(206)
			12	2	(20)
		Atm.	0–40	2	(38, 39)
			15	2	(144)
		190 to 380 mm.	13–38	2	(258, 259)
		Relative pressures	11	1	(215)
	Acetic acid		15	2	(144)
	Heptane	Atm.	0–40	4	(305)
	Carbon tetrachloride	Atm.	0–40	4	(305)
		Atm.	19	3	(276)
		2.5 to 19 mm.	0	3	(141)
			15	2	(144)
	Ethylene dibromide	Atm.	0–40	4	(305)
	Silicon tetrachloride	Atm.	0–40	4	(305)
	Aqueous solutions:				
	Solute:				
	Acetic acid		15	2	(144)
	HCl	Atm.	25	4	(283)
		Atm.	20	4	(206)
		Atm.	−5–+45	2	(103)
			12	2	(20)
		Atm.	20		(228)
	H_2SO_4	Atm.	25	4	(283)
	NaCl	Atm.	25	4	(283)
			10–90	2	(129)
		Atm.	15	2	(163)
		Atm.	−3–+30	2	(103)
	KCl	Atm.	25	4	(283)
		Atm.	−3–+35	2	(103)
	$CaCl_2$, $MnCl_2$, $MgCl_2$		12	2	(20)
	$BaCl_2$	Atm.	25	4	(283)
	$SrCl_2$, $CaCl_2$, $MgCl_2$, $FeCl_3$, $CoCl_2$, $MnCl_2$, $NiCl_2$, $CdCl_2$, LiCl	Atm.	−5–+35	2	(103)
	KNO_3, $NaNO_3$, $Ca(NO_3)_2$, $Mg(NO_3)_2$, K_2SO_4, Na_2SO_4, $MgSO_4$		25		(153)
11. Air	Water	Atm.	22–24	3	(248)
		Atm.	20	3	(49)
		Atm.	0–100	3	(340)
			0–20	3	(65)
		Atm.	20–25	2	(207)
		Atm.	0–25	2	(37, 38, 39, 40)
	Sea water		−5–+35	3	(65)
	Blood fluids	Atm.	20	3	(324)

TABLE 1—*Continued*

GAS	SOLVENT	PRESSURE	TEMPERATURE	VALUE	REFERENCES
			°C.		
B. Elementary gases—Continued:					
11. Air	Ethanol, petroleum fractions, H_2SO_4	Atm.	20	3	(49)
	Cottonseed, herring, cod-liver, maize, linseed, olive, and mineral oils	Atm.	20-50		(268)
C. Compound gases:					
12. Methane	Water	Atm.	0-100	4	(340)
		Atm.	20	3	(49)
		Atm.	0-25	2	(37, 38, 39, 40)
		Atm.	20		(90)
	Methanol	Atm.	20-50	4	(204)
	Ethanol	Atm.	20-40	4	(204)
		Atm.	20	3	(49)
		Atm.	0-25	2	(40)
		Atm.	0-25	2	(44)
	2-Propanol ("isopropanol")	Atm.	20-60	4	(204)
	Amyl alcohol	Atm.	20-30	4	(204)
				1	(94)
	Cyclohexanol	Atm.	25	4	(46)
	Isopentane	2200 to 7600 mm.	30	2	(245)
	Benzene	Atm.	10-60	4	(133)
		Atm.	20-50	4	(207)
	Toluene	Atm.	25-60	4	(204)
	m-Xylene	Atm.	20-60	4	(204)
	Hexane	Atm.	20-40	4	(204)
	Pinene	Atm.	20-55	4	(204)
	Diethyl ether	Atm.	0-15	4	(50)
		Atm.	−80-+20	4	(133)
		Atm.	20		(90)
	Acetone, methyl acetate	Atm.	−80-+40	4	(133)
	Carbon tetrachloride	Atm.	25	4	(134)
		Atm.	−20-+60	4	(133)
	Sulfuric acid	Atm.	20	3	(49)
	Chlorobenzene	Atm.	−40-+100	4	(133)
	Petroleum fractions	Atm.	20	3	(49)
		Atm.	10-20	2	(101)
		5100 to 6300 mm.	30	2	(245)
		Atm.	20		(90)
	Acidified aqueous sodium sulfate solution	Atm.	25	4	(158)
13. Ethane	Water	Atm.	0-100	4	(340)
		Atm.	0-20	2	(267)
		Atm.	0-25	2	(37, 38, 39, 40)
				1	(270)
	Cyclohexanol	Atm.	25	4	(46)
	Benzene	Atm.	0-50	4	(131, 133)
	Chlorobenzene	Atm.	0-80	4	(131, 133)
	Carbon tetrachloride	Atm.	0-40	4	(131, 133)
	Acidified aqueous sodium sulfate solution	Atm.	25	4	(158)

TABLE 1—*Continued*

GAS	SOLVENT	PRESSURE	TEMPERATURE	VALUE	REFERENCES
			°C.		
C. Compound gases—Continued:					
14. Propane......	Cyclohexanol	Atm.	25	4	(46)
	Ethanol, ether, benzene, chloroform, essence of terebenthine	Atm.	Room	1	(175)
15. Butane.......	Water	Atm.	0–25	2	(37, 38, 39, 40)
16. Ethylene......	Water	Atm.	25	4	(230)
		550 to 1000 mm.	25–37.5	4	(105)
		Atm.	15–80	2	(275)
		Atm.	15	2	(23)
		Atm.	0–25	2	(37, 38, 39, 40)
				1	(17)
					(227)
	Ethanol	Atm.	0–25	2	(40)
		Atm.	0–25	2	(44)
				1	(317)
				1	(17)
	Cyclohexanol	Atm.	25	4	(46)
	Acetone			1	(317)
	Benzene	Atm.	10–50	4	(131, 133)
		Atm.	20–50	4	(204)
	Xylene	50 to 750 mm.	−21–+40		(150)
	Carbon tetrachloride	Atm.	0–40	4	(131, 133)
	Chlorobenzene	Atm.	0–90	4	(131, 133)
	Petroleum fractions	Atm.	10–20	2	(101)
		50 to 750 mm.	−21–+40		(150)
	Blood fluids	550 to 1000 mm.	25–37.5	4	(105)
					(227)
	Aqueous solutions:				
	Solute:				
	KOH, NaOH, NH_4OH	Atm.	15	2	(23)
	Acidified sodium sulfate	Atm.	25	4	(158)
17. Propylene.....	Water	Atm.	0–20	2	(306)
				1	(17)
	Ethanol			1	(17)
	Xylene, petroleum fractions	50 to 750 mm.	−21–+40		(150)
18. Cyclopropane	Vegetable, animal, and mineral oils				(231)
19. Isobutylene...	Xylene, petroleum fractions	50 to 750 mm.	−21–+40		(150)
20. Acetylene.....	Water	Atm.	25	4	(198)
		Atm.	37.5	4	(105)
		Atm.	15	3	(23)
			25	3	(169)
			12–20	2	(220)
		Atm.	15–80	2	(275)
			0		(97)
	Ethanol			1	(317)
			18	1	(18)
		50 to 760 mm.	0		(151)

TABLE 1—*Continued*

GAS	SOLVENT	PRESSURE	TEMPERATURE	VALUE	REFERENCES
			°C.		
C. Compound gases—Continued:					
20. Acetylene ...	Amyl alcohol		18	1	(18)
	Cyclohexanol	Atm.	25	4	(46)
	Aniline		−6		(96)
	Dimethylaniline		2		(96)
	Cyclohexane		6		(96)
	Nitrobenzene		5		(96)
	Benzene	Atm.	10–45	4	(133)
			18	1	(18)
			5.5		(97)
	Acetaldehyde, propionaldehyde, butyraldehyde, methylal, acetal, methyl formate, ethyl formate, isoamyl formate, methyl acetate, ethyl acetate, isoamyl acetate, ethyl mustard oil, acetoacetone, ethylidene cyanohydrin, methyl propyl ketone	Atm.	−10	2	(142)
	Carbon tetrachloride	Atm.	0–40	4	(133)
			18	1	(18)
	Chlorobenzene	Atm.	10–45	4	(133)
	Formic acid		7		(97)
	Acetic acid		15		(97)
	Bromoform		7		(97)
	Acetophenone		16		(97)
	Acetone		25	3	(169)
				1	(317)
		Atm.	15–50	1	(52)
		1.3 atm.	−80	1	(51)
			−20–+40		(210)
	Ethylene dichloride	50 to 760 mm.	0		(151)
	Pentane, carbon disulfide, chloroform, styrolene		18	1	(18)
	Petroleum fractions		0	2	(220)
		50 to 760 mm.	0		(151)
	Stannic chloride		30		(98)
	Blood fluids	Atm.	37.5	4	(105)
	Aqueous solutions:				
	Solute:				
	Acetone		25	3	(169)
	NaCl		0	2	(220)
	NH_4Cl, KCl, NaCl, $MgCl_2$, $CaCl_2$, $BaCl_2$, $AlCl_3$, $FeCl_3$, NH_4Br, NaBr, KBr, $NaNO_3$, KNO_3, $Mg(NO_3)_2$, $Ca(NO_3)_2$, $Zn(NO_3)_2$, $Al(NO_3)_3$, $(NH_4)_2SO_4$, Na_2SO_4, K_2SO_4, $MgSO_4$, $ZnSO_4$, $MnSO_4$, $NiSO_4$, $CoSO_4$, $FeSO_4$, $Al_2(SO_4)_3$, $Cr(SO_4)_3$, $Fe_2(SO_4)_3$	Atm.	25	4	(198)
	$Ba(OH)_2$, $Ca(OH)_2$, NH_4OH, NaOH, KOH, Na_2SO_4, K_2SO_4	Atm.	15	3	(23)
	Acidified sodium sulfate solution	Atm.	25	4	(158)

TABLE 1—*Continued*

GAS	SOLVENT	PRESSURE	TEMPERATURE	VALUE	REFERENCES
			°C.		
C. Compound gases—Continued:					
21. Dimethyl ether........	Acetone, methyl acetate	230 to 1000 mm.	25	4	(133)
	Benzene	90 to 1000 mm.	25	4	(133)
	Carbon tetrachloride	100 to 1000 mm.	25	4	(133)
	Chlorobenzene	10 to 1000 mm.	25	4	(133)
	Olive and sesame oils		17-37	3	(209)
22. Methyl chloride.........	Acetone, methyl acetate	230 to 1000 mm.	25	4	(133)
	Benzene	90 to 1000 mm.	25	4	(133)
	Carbon tetrachloride	100 to 1000 mm.	25	4	(133)
	Chlorobenzene	10 to 1000 mm.	25	4	(133)
	Olive and sesame oils		17-37	3	(209)
	Chloroform	200 to 900 mm.	25	4	(133)
23. Chloroethylene....	Ethanol, ethylene dichloride, petroleum fractions	50 to 760 mm.	0		(151)
24. Fluoroethane.	Water		14	2	(214)
	Ethanol, ethyl bromide, diethyl ether			1	(214)
25. Fluoroethylene....	Ethanol, acetone		20	1	(301)
26. Carbon monoxide........	Water	Atm.	0-100	4	(340)
		Atm.	0-60	4	(339)
		Atm.	20-25	3	(147)
		Atm.	20	3	(49)
		Atm.	0-25	2	(37, 38, 39, 40)
		900 to 8200 mm.	20-25	2	(45)
		Atm.	20	2	(185)
	Methanol	Atm.	25	3	(289)
		Atm.	20-25	3	(147)
	Ethanol	Atm.	20-25	3	(147)
		Atm.	20	3	(49)
		Atm.	25	3	(289)
		Atm.	0-25	2	(44)
		Atm.	0-25	2	(40)
	Glycerol	Atm.	25	3	(289)
	Amyl alcohol	Atm.	20-25	3	(147)
	Cyclohexanol	Atm.	25	4	(46)
	Acetic acid	Atm.	25	3	(289)
		Atm.	20-25	3	(147)
	Methyl acetate	Atm.	−80–+40	4	(133)
	Ethyl acetate, isobutyl acetate, amyl acetate	Atm.	20-25	3	(147)
	Benzene	Atm.	10-60	4	(133)
		Atm.	25	3	(289)
		Atm.	20-25	3	(147)
	Toluene, xylene	Atm.	20-25	3	(147)
	Chlorobenzene	Atm.	−40–+80	4	(133)

TABLE 1—*Continued*

GAS	SOLVENT	PRESSURE	TEMPERATURE	VALUE	REFERENCES
			°C.		
C. Compound gases—**Continued:**					
26. Carbon monoxide.......	Carbon tetrachloride	Atm.	−20-+60	4	(133)
	Chloroform	Atm.	20-25	3	(147)
		Atm.	25	3	(289)
	Ethylene dichloride	Atm.	25	3	(289)
	Diethyl ether	Atm.	−80-+20	4	(133)
		Atm.	0-15	4	(50)
	Acetone	Atm.	−80-+40	4	(133)
		Atm.	25	3	(289)
		Atm.	20-25	3	(147)
	Carbon disulfide	Atm.	25	3	(289)
		Atm.	20-25	3	(147)
	Nitrobenzene	Atm.	25	3	(289)
		Atm.	20-25	3	(147)
	Aniline	Atm.	20-25	3	(147)
	Petroleum fractions	Atm.	20	3	(49)
		Atm.	10-20	3	(101) (88)
	Sulfuric acid	Atm.	20	3	(49)
	Blood	1 to 70 mm.	39	1	(138)
	Aqueous solutions:				
	Solute:				
	Alcohol	Atm.	20	2	(185)
	Acidified sodium sulfate	Atm.	25	4	(158)
	Cuprous ammonium carbonate	150 to 2500 mm.	0-75	4	(107)
	Cuprous ammonium carbonate and formate	12 to 370 mm.	0-60	4	(173)
	Non-aqueous solutions:	Atm.	25	3	(289)
	In benzene: phenanthrene, nitrobenzene, α-naphthol, β-naphthol, ethanol				
	In toluene: naphthalene, phenanthrene, aniline, α-naphthol, acetic acid				
	In acetone: phenanthrene, aniline, nitrobenzene, β-naphthol				
	In acetic acid: nitrobenzene, aniline, chloroform, benzene				
	In acetone: chloroform, carbon disulfide				
	In methanol: glycerol, chloroform				
	In carbon disulfide: ethylene dichloride				
27. Carbon dioxide.......	Water	Atm.	25	4	(282)
		750 to 1400 mm.	25	4	(81)
		250 to 1000 mm.	25	4	(84)
		750 to 1400 mm.	25	4	(85)
		Atm.	25	4	(82)
		750 to 1400 mm.	25	4	(86)

TABLE 1—*Continued*

GAS	SOLVENT	PRESSURE	TEMPERATURE °C.	VALUE	REFERENCES
C. Compound gases—Continued:					
27. Carbon dioxide.......	Water	260 to 760 mm.	25	4	(87)
		Atm.	0-25	4	(152)
		Atm.	25	4	(159)
		Atm.	18-36	4	(171)
		Atm.	25	4	(216)
		60 to 800 mm.	0-25	4	(215)
		Atm.	25	4	(230)
		Atm.	0	4	(247)
		Atm.	15	4	(284)
		Atm.	20	4	(320)
		Atm.	0-40	4	(201)
		Atm.	25	4	(61)
		Atm.	38	4	(326)
		Atm.	15	3	(112)
		Atm.	25	3	(100)
		Atm.	15	3	(48)
		Atm.	15-25	3	(147)
		Atm.	20	3	(49)
		520 to 720 mm.	0-20	2	(40)
		Atm.	0-20	2	(37, 38, 39, 40)
				2	(6)
		Atm.	0-60	2	(25)
		560 to 875 mm.	15	2	(279)
		Atm.	15-80	2	(275)
		4 atm.	0	2	(233)
		Atm.	20-25	2	(207)
		Atm.	15-21	2	(219)
		Atm.	8-30	2	(191)
		Atm.	0-40	1	(113)
		700 to 1300 mm.	15	1	(149)
		Relative pressures	11	1	(238)
		500 to 800 mm.	12	1	(36)
			0		(97)
			20		(343)
			17-20		(34)
		Atmospheric to very low	20		(35)
	Heavy water	Atm.	25	4	(61)
	Sea water	500 to 800 mm.	12	1	(36)
	Methanol	Atm.	18-36	4	(171)
		Atm.	15-20	3	(168)
		50 to 760 mm.	−78–−59	3	(296)
		Atm.	15-25	3	(147)
	Ethanol	Atm.	18-36	4	(171)
		Atm.	15-25	3	(147)
		50 to 760 mm.	−78–−59	3	(296)
			15-20	3	(168)
		Atm.	20	3	(49)
		Atm.	0-20	2	(40)
		Atm.	0-25	2	(44)
		Atm.	−67-45	2	(26)
		Atm.	20-25	2	(207)
	Propanol	Atm.	15-25	3	(147)

TABLE 1—*Continued*

GAS	SOLVENT	PRESSURE	TEMPERATURE	VALUE	REFERENCES
			°C.		
C. Compound gases—**Continued:**					
27. Carbon dioxide.......	Isobutyl alcohol	Atm.	15–25	3	(147)
	Amyl alcohol	Atm.	15–25	3	(147)
	Isoamyl alcohol	Atm.	18–36	4	(171)
	Cyclohexanol	Atm.	25	4	(46)
	Acetic acid	Atm.	18–36	4	(171)
		Atm.	15–25	3	(147)
	Propionic acid	Atm.	15–25	3	(147)
	Butyric acid	Atm.	15–25	3	(147)
	Formic acid	Atm.	7		(97)
	Acetic acid	Atm.	15		(97)
	Methyl acetate	Atm.	25	4	(130)
		Atm.	15–25	3	(147)
		50 to 760 mm.	−78–−59	3	(296)
	Ethyl acetate	50 to 760 mm.	−78–−59	3	(296)
	Amyl acetate	Atm.	18–36	4	(171)
		Atm.	15–25	3	(147)
	Amyl formate	Atm.	15–25	3	(147)
	Isobutyl acetate	Atm.	15–25	3	(147)
	Benzene	Atm.	5.5		(97)
			5.5		(98)
		Atm.	15–25	3	(147)
	Toluene	Atm.	15–25	3	(147)
	Chloroform	Atm.	18–36	4	(171)
		Atm.	0	4	(130)
		Atm.	15–25	3	(147)
		36 to 760 mm.	13	3	(329)
	Bromoform	Atm.	9		(97)
			9		(98)
	Carbon tetrachloride	Atm.	0–25	4	(130)
		Atm.	25	4	(131, 133)
		Atm.	15–25	3	(147)
	Ethyl chloride	Atm.	17.5	2	(307)
	Ethylene dichloride	Atm.	15–25	3	(147)
	Ethylene dibromide	Atm.	18–36	4	(171)
		Atm.	15–25	3	(147)
	Acetone	Atm.	10–25	4	(130)
		Atm.	18–36	4	(171)
		Atm.	15–25	3	(147)
		50 to 760 mm.	−78–−59	3	(296)
	Acetophenone	Atm.	16		(97)
	Diethyl ether	Atm.	0	4	(130)
			−64–+15	2	(307)
	Pyridine	Atm.	18–36	4	(171)
		Atm.	15–25	3	(147)
	Carbon disulfide	100 to 900 mm.	7–20	3	(328)
		Atm.	15–25	3	(147)
	Nitrobenzene	Atm.	15–25	3	(147)
			6		(98)
	p-Azoxyphenetole	Atm.	145–170	3	(128)
	Chlorobenzene	Atm.	25	4	(130)
		Atm.	15–25	3	(147)
	Aniline	Atm.	18–36	4	(171)
		Atm.	15–25	3	(147)

TABLE 1—*Continued*

GAS	SOLVENT	PRESSURE	TEMPERATURE	VALUE	REFERENCES
			°*C.*		
C. Compound gases—Continued: 27. Carbon dioxide.......	Benzaldehyde	Atm.	18-36	4	(171)
		Atm.	15-25	3	(147)
	Glycerol, bromobenzene, iodobenzene, benzyl chloride, propylene bromide, amyl bromide, amyl chloride, isobutyl chloride, benzotrichloride, *o*-toluidine, *m*-toluidine, acetic anhydride, dichlorohydrin, cumene, eugenol	Atm.	15-25	3	(147)
	Sulfuric acid	Atm.	25	4	(201a)
		Atm.	20	3	(49)
	Petroleum fractions	Atm.	20	3	(49)
		Atm.	2-25	2	(170)
		Atm.	10-20	2	(101)
		Atm.	20-50		(268)
					(88)
	Cottonseed, herring, olive, linseed, maize, and cod-liver oils	Atm.	20-50		(268)
	Cottonseed oil, corn oil, lard		23-45		(327)
	Blood	Atm.	38	4	(326)
	Beer	750 to 1400 mm.	25	4	(85)
	Aqueous solutions:				
	Solute:				
	Ethanol	Atm.	15	4	(284)
		750 to 1400 mm.	25	4	(85)
		Atm.	0-25	4	(152)
		Atm.	15-21	2	(219)
			0		(310)
	Dextrose, levulose, sucrose	Atm.	15	4	(284)
	Glycerol, acetone, urea	Atm.	0-25	4	(152)
	Glycerol, chloral hydrate	Atm.	15	3	(112)
	Sucrose, chloral hydrate, KCl, $Fe(NH_4)_2(SO_4)_2$, NH_4Cl, $BaCl_2$	750 to 1400 mm.	25	4	(86)
	Propanol, acetic acid, acetamide, antipyrene, urea, thiourea, urethan, catechol, resorcinol, quinol, pyrogallol, glycine, mannitol, dextrose, sucrose	Atm.	20	4	(320)
	Ethanol, acetone, urea, glycerol		0		(310)
	KCl	750 to 1400 mm.	25	4	(81)
	NaCl, KCl, Na_2HPO_4, K_2HPO_4, sodium lactate, potassium lactate	Atm.	38	4	(326)
	H_2SO_4, NaCl, Na_2SO_4, Na_3PO_4, $CaCl_2$, $MgCl_2$, $ZnCl_2$, $AlCl_3$, $Al_2(SO_4)_3$	Atm.	25	4	(159)
	HNO_3, HCl, H_2SO_4, CsCl, KNO_3, KI, RbCl, KBr, KCl	Atm.	15-25	3	(100)

TABLE 1—*Continued*

GAS	SOLVENT	PRESSURE	TEMPERATURE	VALUE	REFERENCES
			°C.		
C. Compound gases—Continued:					
27. Carbon dioxide.......	Sucrose, LiCl, NaCl, KCl, KBr, KI, KNO_3, H_2SO_4, $MgSO_4$, $CuSO_4$, $ZnSO_4$, $(NH_4)_2SO_4$, $KHSO_3$, $KHSO_4$, KH_2AsO_4, KH_2PO_4, K_2HAsO_4, K_2HPO_4, Na_3PO_4, $Na_4P_2O_7$, $Na_2B_4O_7$	Atm.	15	3	(48)
	NaCl, $NaNO_3$, Na_2SO_4, KCl, KNO_3, $Mg(NO_3)_2$, $MgSO_4$	Atm.	0-40	4	(201)
	H_2SO_4, $HClO_4$	Atm.	25	4	(201a)
	Citric, tartaric, metaphosphoric acids; $NaNO_3$, NaBr, Na_2SO_4, LiCl, $MgCl_2$, $MgSO_4$, $Ca(NO_3)_2$, $CaCl_2$, $Co(NO_3)_2$, $K_4Fe(CN)_6$, $ZnSO_4$, $Zn(NO_3)_2$, $Pb(NO_3)_2$, NH_4Cl, $(NH_4)_2SO_4$, KCl, KBr, KI, KCNS, $NaClO_3$	Atm.	15	2	(279)
	NaCl, KCl, NH_4Cl	Atm.	8-22	2	(191)
	$CaCl_2$, $SrCl_2$, $BaCl_2$	Atm.	8-30	2	(191)
	NaCl	Atm.	0-60	2	(25)
	$(NH_4)_2CO_3$	3 to 12 mm.	25	3	(33)
	$MgSO_4$, $CaSO_4$	500 to 800 mm.	12	1	(36)
	NaCl, $CaCl_2$		20		(343)
	NaCl				(6)
	Acidified sodium sulfate solution	Atm.	25	4	(158)
	Sugar liquors				(195)
	Ternary solutions:				
	Dextrose-ethanol-water	Atm.	15	4	(284)
	Sucrose–ethanol-water				
	Non-aqueous solutions:				
	Methanol and ethanol solutions of LiCl, LiBr, LiI, NaCl, NaBr, NaI		15-20	3	(168)
	Acetic acid-carbon tetrachloride and carbon disulfide–ethylene dichloride solutions	Atm.	15	3	(48)
28. Carbonyl sulfide......	Water	Atm.	13.5	2	(116)
			20	2	(299)
	Ethanol		20	1	(299)
	Toluene		20	1	(299)
			−14	1	(330)
	Carbon disulfide		20	1	(299)
	Aqueous sodium chloride solution		20	1	(299)

TABLE 1—*Continued*

GAS	SOLVENT	PRESSURE	TEMPERATURE	VALUE	REFERENCES
			°C.		
C. Compound gases—Continued:					
29. Carbonyl chloride.....	Ethanol, acetic acid, benzene	Atm.	20	2	(11)
	Toluene	Atm.	12-31	3	(7)
		Atm.	20	2	(11)
	Xylene	Atm.	12-31	3	(7)
	Chloroform, carbon tetrachloride	Atm.	20	2	(11)
	Chlorobenzene, acetylene tetrachloride, creosote	Atm.	12-31	3	(7)
	Petroleum fractions	Atm.	12-31	3	(7)
		Atm.	20	2	(11)
30. Cyanogen.....	Water and aqueous hydrochloric acid solution		18	1	(223)
31. Hydrogen cyanide.......	Water	1 to 7 mm. (?)	25	2	(179)
32. Silane.........	Cyclohexanol	Atm.	25	4	(46)
33. Ammonia.....	Water	6 to 500 mm.	0-60	4	(241)
		1 to 100 mm.	0-25	4	(215)
		750 to 4000 mm.	0-40	4	(226)
		0.024 to 1.6 mm.	0	4	(154)
		8 to 60 mm.	60	4	(66)
		6 to 14 mm.	25	4	(99)
		Distillation under atmospheric pressure		4	(42)
		20 to 1800 mm.	0-61	3	(239)
		200 to 1000 mm.	20-60	3	(240)
		1 to 15 mm.	25	3	(183)
		200 to 2000 mm.	0-100	3	(287)
		About 2 to 70 mm.	20	3	(184)
		10 to 2000 mm.	0	3	(260)
		Atm.	0-56	3	(260)
		13 mm.	25	3	(1)
		Atm.	0-29	2	(254)
		Atm.	0-20	1	(40)
		Relative pressures	16	1	(238)
	Methanol	Atm.	0-28	3	(182)
	Ethanol	Atm.	0-28	3	(182)
			0	1	(9)
	Cyclohexanol	Atm.	25	4	(46)
	Quinoline	Relative pressures	16	1	(238)
	Benzene, toluene, hexane, octane, dodecane, cetane, carbon tetrachloride, chloroform, ethylene dichloride, chlorobenzene, bromobenzene, benzyl chloride	Atm.	20	4	(15)

TABLE 1—*Continued*

GAS	SOLVENT	PRESSURE	TEMPERATURE	VALUE	REFERENCES
			°C.		
C. Compound gases—Continued:					
33. Ammonia	Aqueous solutions				
	Solute:				
	$NaOH$, NH_4Cl, NH_4NO_3, NH_4I, NH_4CNS, $(NH_4)_2SO_4$, $(NH_4)_2C_2O_4$, $(NH_4)_2C_4H_4O_6$, KCl, $NaCl$, $BaCl_2$, $CaCl_2$, $SrCl_2$, $MgCl_2$, $AgCl$, $CuCl$, $CuSO_4$, $ZnSO_4$, $CdSO_4$	6 to 14 mm.	25	4	(99)
	$CuSO_4$	1 to 15 mm.	25	3	(183)
	Urea, mannitol, K_2SO_4, NH_4Cl, $CuSO_4$	200 to 1000 mm.	20–60	3	(240)
	NH_4CNS	Up to 2 atm.	10–30	3	(91)
	KCl, KBr, KI, KOH, KF, $NaCl$, $NaBr$, NaI, $NaOH$, $LiCl$, $LiBr$, LiI, $LiOH$, KNO_3, KNO_2, KCN, $KCNS$, KBO_2, K_2SO_4, K_2CO_3, K_2CrO_4, $K_2C_2O_4$, CH_3COOK, K_2HPO_4, Na_2S, $KClO_3$, $KBrO_3$, KIO	13 mm.	25	3	(1)
	$(NH_4)_2CO_3$	8 to 23 mm.	25	3	(33)
	NH_4Cl, $NaNO_3$, NH_4NO_3, KOH, $NaOH$, $Ca(NO_3)_2$	Atm.	0–29	2	(254)
	KOH, $NaOH$, K_2CO_3, CH_3COOK, $(COOK)_2$, KCl, Na_2CO_3, CH_3COONa, $HCOONa$, $NaCl$, $BaCl_2$, $SrCl_2$, $CaCl_2$, $LiCl$, NH_4Cl, $(NH_4)_2SO_4$, K_2SO_4, KCl, KNO_3, KBr, KI, $Cd(NO_3)_2$, $Zn(NO_3)_2$, $(HCOO)_2Ba$, $(CH_3COO)_2Ba$, $NiCl_2$, $Cu(NO_3)_2$, $AgNO_3$, $NiSO_4$, $CuCl_2$, $CuSO_4$, $(CH_3COO)_2Cu$	About 60 mm.	60		(165, 166)
	Distribution data:				
	Water-ether		20	4	(184)
	Water-chloroform		25	4	(16)
			20	3	(64)
	Chloroform-aqueous solutions:				
	Solute: $CuSO_4$, $CuCl_2$, CdI_2, $NiSO_4$, Na_2SO_4, CuO, $ZnSO_4$		20	3	(64)
34. Methylamine.	Water	8 to 60 mm.	60	4	(66)
		2 to 17 mm.	25	4	(80)
			12.5	1	(345)
	Distribution data:				
	Water-ether		25	4	(290)
	Water-xylene		25	4	(290)

TABLE 1—*Continued*

GAS	SOLVENT	PRESSURE	TEMPERATURE	VALUE	REFERENCES
			°C.		
C. Compound gases—Continued:					
35. Dimethylamine.......	Water	3 to 26 mm.	25	4	(80)
	Distribution data: water-ether, water-xylene		25	4	(290)
36. Trimethylamine.......	Water	17 to 133 mm.	25	4	(80)
		35 to 60 mm.	16-22	2	(174)
	Methanol	9 to 13 mm.	25	3	(108)
	Ethanol	7 to 12 mm.			
	Propanol	8 to 13 mm.			
	Isoamyl alcohol	16 mm.			
	Benzyl alcohol	1 to 25 mm.			
	Hexane	22 to 42 mm.			
	Benzene	28 to 33 mm.			
	Chloroform	5 to 9 mm.			
	α-Bromonaphthalene	50 to 85 mm.			
	Acetone	52 to 77 mm.			
	Acetophenone	70 to 120 mm.			
	Diethyl ether	30 to 50 mm.			
	Acetonitrile	35 to 50 mm.			
	Nitromethane	35 to 56 mm.			
	o-Nitrotoluene	85 mm.			
	Nitrobenzene	65 to 95 mm.			
	Ethyl acetate	50 to 60 mm.			
	Ethyl benzoate	40 mm.			
37. Ethylamine...	Water	8 to 60 mm.	60	4	(66)
		2 to 18 mm.	25	4	(62)
	Distribution data: water-ether, water-xylene		25	4	(290)
38. Diethylamine.......	Water	4 to 30 mm.	25	4	(62)
39. Triethylamine.......	Water	5 to 30 mm.	25	4	(62)
		15 to 170 mm.	6-50	3	(174)
	Hexane	2 mm.	25	3	(108)
	Nitromethane	8 to 14 mm.	25	3	(108)
40. Propylamine..	Water	8 to 60 mm.	60	4	(66)
41. Nitrous oxide.	Water	750 to 1400 mm.	25	4	(81)
		250 to 1000 mm.	25	4	(83)
		Atm.	0-40	4	(201)
		Atm.	25	4	(230)
		Atm.	25	4	(198)
		Atm.	18-36	4	(171)
		Atm.	25	3	(100)
		Atm.	20	3	(157)
			5-20	3	(262)
		Atm.	15-80	2	(275)
		Atm.	8-22	2	(104)

TABLE 1—*Continued*

GAS	SOLVENT	PRESSURE	TEMPERATURE	VALUE	REFERENCES
			°C.		
C. Compound gases—Continued:					
41. Nitrous oxide.	Water	Atm.	0-25	2	(40)
		Atm.	0		(97) (285)
	Ethanol	Atm.	18-36	4	(171)
		Atm.	0-25	2	(40)
		Atm.	0-25	2	(44)
		Atm.	16-18	2	(186)
	Benzene	Atm.	10-40	4	(130, 133)
		Atm.	5.5		(97)
	Acetone	Atm.	0-40	4	(130, 133)
		Atm.	18-36	4	(171)
	Acetic acid	Atm.	18-36	4	(171)
		Atm.	15		(97)
	Methyl alcohol	Atm.	18-36	4	(171)
	Isoamyl alcohol	Atm.	18-36	4	(171)
	Cyclohexanol	Atm.	25	3	(46)
	Formic acid	Atm.	7		(97)
	Methyl acetate	Atm.	10-40	4	(130, 133)
	Amyl acetate	Atm.	18-36	4	(171)
	Chloroform	Atm.	18-36	4	(171)
	Carbon tetrachloride	Atm.	10-40	4	(130, 133)
	Bromoform	Atm.	7		(97)
	Ethylene dibromide	Atm.	18-36	4	(171)
	Chlorobenzene	Atm.	10-55	4	(130, 133)
	Acetophenone	Atm.	16		(97)
	Pyridine	Atm.	18-36	4	(171)
	Benzaldehyde	Atm.	18-36	4	(171)
	Aniline	Atm.	18-36	4	(171)
	Petroleum fractions	Atm.	10-20	2	(101)
	Olive and sesame oils	Atm.	17-37	3	(209)
	Aqueous solutions:				
	Solute:				
	Propionic acid, chloral hydrate	Atm.	20	3	(157)
	Urea, oxalic acid, glycerol		5-25	3	(262)
	NH_4Cl, KCl, $CaCl_2$, $NaCl$, $BaCl_2$, NH_4Br, KBr, $NaBr$, NH_4NO_3, KNO_3, $NaNO_3$, $Mg(NO_3)_2$, $Ca(NO_3)_2$, $Zn(NO_3)_2$, $Cd(NO_3)_2$, $Cu(NO_3)_2$, $Al(NO_3)_3$, $(NH_4)_2SO_4$, K_2SO_4, Na_2SO_4, $MgSO_4$, $ZnSO_4$, $MnSO_4$, $FeSO_4$, $CoSO_4$, $NiSO_4$, $Al_2(SO_4)_3$, $Fe_2(SO_4)_3$, $Cr_2(SO_4)_3$, KIO_4, Na_2HPO_4, Na_3PO_4	Atm.	25	4	(198)
	$NaCl$, Na_2SO_4, KCl, KNO_3, $Mg(NO_3)_2$, $MgSO_4$	Atm.	0-40	4	(201)
	KNO_3, $NaNO_3$	Atm.	20	3	(157)
	H_3PO_4, $NaCl$		5-25	3	(262)
	$LiCl$, $NaCl$, KCl, Na_2SO_4, K_2SO_4, Li_2SO_4, $CaCl_2$, $SrCl_2$, $MgSO_4$	Atm.	8-22	3	(104)

TABLE 1—*Continued*

GAS	SOLVENT	PRESSURE	TEMPERATURE	VALUE	REFERENCES
			°C.		
C. Compound gases—Continued:					
41. Nitrous oxide.	HNO_3, HCl, H_2SO_4, CsCl, KNO_3, KI, RbCl, KBr, KCl, LiCl, NH_4Cl, KOH	Atm.	15–25	3	(100)
	H_2SO_4, $FeSO_4$, NaOH, pyrogallol (alkaline)		16–18	2	(186)
	Acidified sodium sulfate solution	Atm.	25	4	(158)
	Blood fluids	Atm.	37.5	4	(230) (285)
42. Nitric oxide..	Water	Atm.	0–100	3	(340)
		Atm.	0–60	3	(339)
			20	2	(319)
		250 to 2000 mm.	0–16	2	(200)
			0		(98)
	Ethanol	Atm.	0–25	2	(40)
		Atm.	0–25	2	(44)
	Benzene	520 to 1000 mm.	9–35	3	(156)
			5		(98)
	Carbon tetrachloride	450 to 1000 mm.	9–35	3	(156)
	Nitrobenzene	450 to 1000 mm.	20–90	3	(156)
			5		(98)
	Bromoform		8		(98)
	Cyclohexane		6		(98)
	Aqueous solutions:				
	Solute:				
	H_2SO_4	Atm.	18	3	(314)
		Atm.	18	2	(187)
			0	2	(199)
				1	(197)
	$FeSO_4$, $FeCl_2$	250 to 2000 mm.	0–16	2	(200)
	$NiSO_4$, $CoSO_4$, $MnCl_2$, ferrous salt	550 to 700 mm.	20	2	(137)
	Ferrous salts			1	(308)
	Ethanol solution of ferrous chloride	700 to 2000 mm.	2–28	2	(200)
43. Phosphine....	Water		15	1	(73)
	Cyclohexanol	Atm.	25	4	(46)
44. Methylphosphine..	Ethanol, ether		0	1	(126)
45. Arsine........	Water	150 to 760 mm.	20	4	(99a)
		200 mm.	0–25	4	(146)
		100 to 360 mm.	20	4	(343a)
46. Stibine........	Water		Room	2	(298)
				1	(143)
	Ethanol, benzene		Room	2	(298)
	Carbon disulfide		0	2	(298)
47. Hydrogen sulfide......	Water	270 to 3500 mm.	5–60	4	(344)
		Atm.	0	4	(247)
		Atm.	25	4	(148)

TABLE 1—*Continued*

GAS	SOLVENT	PRESSURE	TEMPERATURE	VALUE	REFERENCES
			°C.		
C. Compound gases—Continued:					
47. Hydrogen sulfide......	Water	Atm.	25	4	(244)
		Atm.	0–25	4	(152)
		Atm.	0–40	3	(37, 38, 39, 40)
		Relative pressures	12	1	(238)
			0		(97)
	Ethanol	Atm.	0–25	3	(40)
		Atm.	0–25	3	(44)
	Glycerol		25	3	(205)
	Ether	740 mm.	26	3	(232)
	Aniline	100 to 1200 mm.	22	4	(10)
	Pyridine	750 mm.	22	1	(269)
	Benzene, hexane, cyclohexane, octane, dodecane, cetane, carbon tetrachloride, chloroform, chlorobenzene, bromobenzene, toluene, ethylene dichloride, trichloroethylene, tetrachloroethylene, pentachloroethane, ethyl bromide, bromoform, *s*-tetrachloroethane, *s*-tetrabromoethane	Atm.	20	4	(15)
	Benzene		5.5		(97)
	Bromoform		7		(97)
	Formic acid		7		(97)
	Acetic acid		15		(97)
	Acetophenone		16		(97)
	Sulfur		440	1	(236)
	Aqueous solutions:				
	Solute:				
	HCl	Atm.	25	4	(148)
	HI	Atm.	25	4	(244)
	NaHS		15–45	2	(102)
	Ethanol	Atm.	0–25	4	(152)
			0		(310)
	Glycerol	Atm.	0–25	4	(152)
			0		(310)
	Acetone	Atm.	0–25	4	(152)
			0		(310)
	Urea	Atm.	0–25	4	(152)
			0		(310)
	HCl, NaCl, NH_4Cl, $NaNO_3$, KNO_3, NH_4NO_3, NaBr, KBr, NH_4Br, KI, CH_3COONH_4, H_2SO_4, Na_2SO_4, K_2SO_4, $(NH_4)_2SO_4$		25	3	(205)
48. Sulfur dioxide.......	Water	250 to 2500 mm.	10–27	4	(189)
		100 to 1500 mm.	20–110	4	(22)
		30 to 800 mm.	0–25	4	(215)
		Atm.	10–90	4	(135)
		50 to 1100 mm.	5–60	4	(291)

TABLE 1—*Continued*

GAS	SOLVENT	PRESSURE	TEMPERATURE °C.	VALUE	REFERENCES
C. Compound gases—Continued:					
48. Sulfur dioxide	Water	Atm.	25-35	4	(92)
		78 mm.	25	4	(92)
		0 to 180 mm.	15-25	4	(21)
		20 to 2000 mm.	7-50	3	(287)
		0 to 2800 mm.	0-12	3	(13)
		Atm.	0-40	2	(38, 39, 40)
		Relative pressures	12	1	(238)
		Atm.	0-100		(229)
	Methanol	Atm.	0-26	3	(182)
	Ethanol	Atm.	0-26	3	(182)
		Atm.	0-40	2	(40)
		Atm.	0-25	4	(133)
	Methyl acetate, acetone	200 to 1000 mm.	25	4	(133)
	Benzene	100 to 1000 mm.	25	4	(133)
		Atm.	25	4	(134)
		Atm.	30-60	4	(181)
	Toluene	Atm.	20-60	4	(181)
	Nitrobenzene, *o*-nitrotoluene	Atm.	15-60	4	(181)
	Carbon tetrachloride	Atm.	25-40	4	(130)
		100 to 950 mm.	25	4	(133)
	Chlorobenzene	10 to 1050 mm.	25	4	(133)
	Acetic anhydride	Atm.	−5–+30	4	(181)
	Camphor	700 mm.	4-24	2	(271)
	Sulfuric acid	Atm.	20	3	(214)
	Aqueous solutions:				
	Solute:				
	H_2SO_4	Atm.	20	3	(214)
		Atm.	10-15	2	(72)
	KCl, Na_2SO_4	Atm.	10-90	4	(135)
	$Ca(HSO_3)_2$	0 to 180 mm.	15-25	4	(21)
	$Ca(HSO_3)_2$, $Mg(HSO_3)_2$	50 to 1100 mm.	5-60	4	(291)
	KI, KBr, KCl, KCNS, NH_4NO_3, KNO_3, $(NH_4)_2SO_4$, CdI_2, Na_2SO_4, $CdBr_2$, $CdCl_2$, $CdSO_4$	Atm.	25-35	4	(92)
	KI, KCNS, KBr, KCl, KNO_3, $(NH_4)_2SO_4$	78 mm.	25	4	(92)
	Acidified sodium sulfate solution	Atm.	25	4	(158)
49. Hydrogen selenide	Water, hydriodic acid	Atm.	15-35	4	(203)
	Selenium		580	1	(235)
50. Hydrogen chloride	Water	0.04 to 560 mm.	50	4	(346)
		0.01 to 4 mm.	25-30	4	(12)
		0.001 to 0.1 mm.	30	3	(95)
		60 to 1300 mm.	0-100	3	(260)
		0 to atmospheric	30	3	(67)
			−12-0	1	(19)
		Relative pressures	11	1	(238)

TABLE 1—*Continued*

GAS	SOLVENT	PRESSURE	TEMPERATURE °C.	VALUE	REFERENCES
C. Compound gases—Continued:					
50. Hydrogen chloride....	Methanol	Atm.	0–32	3	(182)
	Ethanol	Atm.	0–32	3	(182)
	Chloroform		10	2	(334)
	Benzene, hexane, cyclohexane, octane, dodecane, cetane, carbon tetrachloride, chloroform, bromoform, chlorobenzene, bromobenzene, toluene, ethylene dichloride, trichloroethylene, tetrachloroethylene, pentachloroethane, ethyl bromide, *s*-tetrachloroethane, *s*-tetrabromoethane, benzotrichloride, benzyl chloride	Atm.	20	4	(15)
	1,1,2-Trichloroethane, pentachloroethane	500 to 700 mm.	12–20		(110)
	Ethylene dichloride, ethylene dibromide, acetylene tetrachloride, carbon tetrachloride	300 to 800 mm.	15–25		(109)
	Diethyl ether		−9–+30	3	(273)
	Isoamyl ether	Atm.	0–25	3	(237)
	Aqueous solutions:				
	Solute:				
	Ethanol	0.04 to 5 mm.	25	3	(145)
	Sulfuric acid		17–70		(56)
					(60)
51. Hydrogen bromide....	Water	0.001 to 0.1 mm.	25	4	(12)
			10	1	(19)
	Benzene	8 to 630 mm.	30–50		(147a)
52. Hydrogen iodide.......	Water	0.0005 to 0.1 mm.	25	4	(12)
			10	1	(19)
D. Radioactive gases:†					
1. Radium emanation......	Water		0–100		(302)
					(315)
			0–80		(272)
			0–40		(29)
			15–30		(124)
			0–80		(127)
			0–91		(161)
			0–100		(208)
			0–18		(252)
	Sea water		14		(29)
	Methanol		15–30		(124)
	Ethanol		−18–+18		(252)
			−18–+18		(208)
			−18–+50		(272)
			14		(29)
					(315)

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† The pressures are not given directly. The methods used were all similar, and there is relatively little basis for evaluating the results, except the temperature control.

TABLE 1—*Continued*

GAS	SOLVENT	PRESSURE	TEMPERATURE	VALUE	REFERENCES
			°C.		
***D. Radioactive gases†* —Continued:** 1. Radium emanation......	Propanol, butanol, formic acid, acetic acid, propionic acid, butyric acid		15-30		(124)
	Amyl alcohol		14		(29)
	Glycerol		18		(252)
	Ethyl acetate		−18-+18		(208)
			−18-+18		(252)
			−18-+60		(272)
	Benzene		6-73		(302)
			18		(252)
			18		(272)
	Toluene		−18-+18		(208)
			14		(29)
			−18-+60		(272)
			−18-+18		(252)
	Xylene		−18-+18		(252)
			−20-+70		(208)
	Hexane		−18-+30		(272)
			−18-+18		(252)
	Cyclohexane		−18-+18		(252)
	Chloroform		−18-+18		(208)
			−18-+18		(252)
			−20-+50		(272)
	Aniline		−18-+18		(252)
					(315)
			0-18		(272)
	Acetone		−18-+18		(208)
			−20-+40		(272)
			−18-+18		(252)
	Carbon disulfide		−18-+40		(272)
			−18-+18		(208)
			−18-+18		(252)
	Diethyl ether		−18-+18		(208)
			−18-+18		(252)
			−18-+30		(272)
	Petroleum fractions		−18-+18		(252)
			−21-+60		(127)
					(315)
	Aqueous solutions: Solute:				
	$CuSO_4$				(315)
	NaCl, $Ba(NO_3)_2$, NH_4NO_3, urea		5-30		(162)
	Ethanol, sucrose, KCl, NaCl, NH_4Cl, $Pb(NO_3)_2$, $AgNO_3$, $HgCl_2$, $ZnSO_4$, $CuSO_4$, $FeSO_4$, $KMnO_4$, $K_4Fe(CN)_6$		18		(160)
2. Thorium emanation......	Water		Room		(155)
					(28)
	Ethanol				(28)

TABLE 1—*Concluded*

GAS	SOLVENT	PRESSURE	TEMPERATURE °C.	VALUE	REFERENCES
D. Radioactive gases† —Continued:					
2. Thorium emanation......	Petroleum fractions		Room		(155) (28)
	Aqueous solutions of H_2SO_4, $CuSO_4$, $CaCl_2$				(28)
3. Actinium emanation..	Water, acetone, benzene, ethanol, amyl alcohol, benzaldehyde, toluene, carbon disulfide, petroleum, sulfuric acid, aqueous KCl solution		Room		(119)

† The pressures are not given directly. The methods used were all similar, and there is relatively little basis for evaluating the results, except the temperature control.

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