INTRODUCTION TO GIBBS PHASE RULE (lecture 1)

The Gibbs phase rule describes the degrees of freedom available to describe a particular system with various phases and substances. To derive the phase rule, let us begin with a system that has "c" independent chemical species. We will assume no chemical reactions (which reduces the number of independent chemical species). Now let each chemical component exist in the maximum number of phases, "p". When we include the two quantities, temperature and pressure, as degrees of freedom that can describe our system, the total degrees of freedom of the system (before considering thermodynamic constraints) is

$$f = p \times c + 2$$

However, the number of chemical components in each phase is not truly independent since the total of the each mole fraction must equal one. $X_1 + X_2 + X_3 + \ldots X_c = 1$

This yields one constraint for each phase; therefore, the total degrees of freedom becomes

$$f = p \times c + 2 - p = p(c-1) + 2$$

References

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2. **Essentials of Physical Chemistry by B. S. Bahl, G. D. Tuli, Golden Jubilee Ed., S. Chand and Company**

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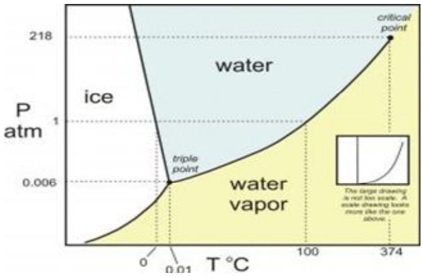
ONE COMPONENT SYSTEM (lecture 2)

Some Simple One-Component Examples: simple 1 component system for H_2O:



Phase diagram for the one component system H_2O. The system is entirely composed of H_2O, so there is only one component present. The phases present represent three states of matter: liquid (water), solid (ice), and vapor steam). All have distinct physical properties (e.g. density, structure--or lack of, etc.) and chemical properties (*e.g.* $\Delta G_{\text{formation}}$, molar volume etc.) so they must be considered distinct phases.

- Note that there is only one point on this diagram where all three phases coexist in equilibrium--this "triple point" is also referred to as an *invariant point*; because P and T are uniquely specified, there are zero degrees of freedom.
- Each of the curves represents a chemical reaction that describes a phase transformation: solid to liquid (melt/crystallization), liquid to vapor (boiling/condensation), solid to vapor (sublimation/deposition).There is only one degree of freedom along each of the univariant curves: you can independently change either T or P, but to maintain two coexisting phases along the curve the second variable must change by a corresponding fixed amount.
- There are three distinct areas where only ice, liquid, or vapor exit. These are *divariant* fields. T and P are both free to change within these fields and you will still have only one phase (a bit hotter or colder, or compressed or expanded, but nonetheless the same phase).
- The end of the "boiling curve", separating the liquid to vapor transition, is called the "critical point". This is a particularly interesting part of the phase diagram because beyond this region the physico-chemical properties of water and steam converge to the point where they are identical. Thus, beyond the critical point, we refer to this single phase as a "supercritical fluid".

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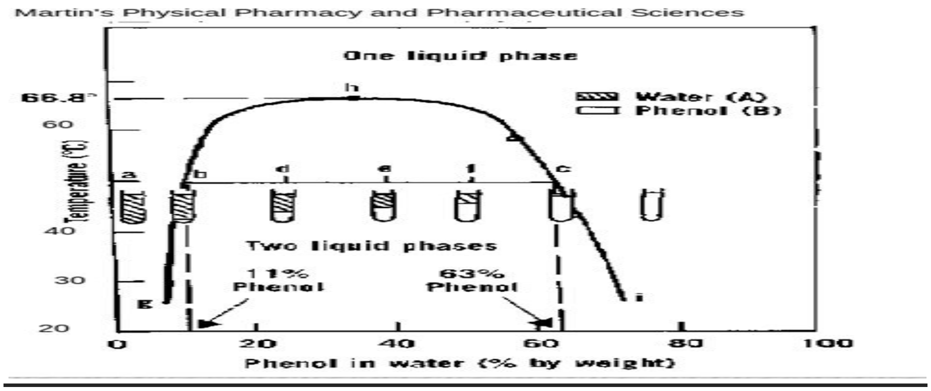
TWO COMPONENT SYSTEM (lecture 3)

Liquid water and phenol show limited miscibility below 70ºC. In an experiment, miscibility temperatures of several water-phenol mixtures of known composition were measured. The phase coexistence line (miscibility temperature versus composition) will be determined, and the critical composition and temperature calculated. An extension of the experiment will be to measure the effect of a third component on the water-phenol critical point.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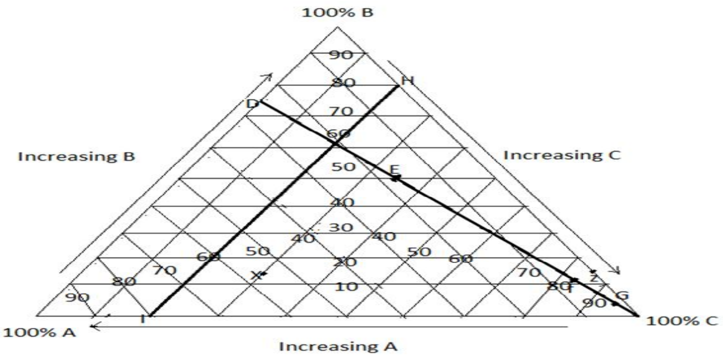
THREE COMPONENT SYSTEM (lecture 4)

In system containing 3 components but only one phase, F=3-1+2=4 for a non condensed system. The 4 degrees of freedom are temperature, pressure and concentration of tow of the three components. Only two concentration terms required because, the sum of this substracted from the total will give the concentration of the three component.



RULES FOR TRIANGULAR DIAGRAM

Considering the three component system, there are various rules for the ternary diagram which will be studied.

1) Each of three corners or apexes of the triangle represent 100% by weight of one component (A, B or C). 2) The three lines joining the corner points represents 2 component mixture of the 3 possible combinations of A,B,C. thus the lines AB, BC &CA are used for 2 components mixtures of A & B,B & C, C & A resp. 3) The area within the triangle represents. All the possible combinations of A , B and C to give 3 component systems.
4) If a line is draw through any apex to a point on the opposite side, then all system represented by points on such a line have a constant ratio of components, in this case A and B. furthermore, the continual addition of C to a mixture of A and B will produce system that lie progressively closer to apex C.
5) Any line draw parallel to 1 side of the triangle , for e.g. line HI in fig. represents ternary systems in which the proportion of 1 component is constant .In this instance, all system prepared along HI will contain 20% of C and varying concentrations of A and B.

References

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2. Essentials of Physical Chemistry by B. S. Bahl, G. D. Tuli, Golden Jubilee Ed., S. Chand and Company

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DEVIATION FROM GAS THEORY (lecture 5)

The deviation from ideal behavior of a gas is expressed in terms of the compressibility factor Z, which is defined as the ratio of the actual volume to the volume predicted by the ideal gas law.

$$Z = \text{Actual volume/volume predicted by ideal gas law} = v/RT/P = Pv/RT$$

For an ideal gas $Pv = RT$ and hence $Z = 1$ at all temperatures and pressures.

The experimental P-v-T data is used to prepare the compressibility chart.

Reduced pressure, $P_R = P/P_c$,

Reduced temperature, $T_R = T/T_c$

Reduced volume, $v_R = v/v_c$

Where P_c, T_c and v_c denote the critical pressure, temperature and volume respectively.

References

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LAW OF CORRESPONDING STATES (lecture 6)

According to van der Waals, the **theorem of corresponding states** (or **principle of corresponding states**) indicates that all fluids, when compared at the same reduced temperature and reduced pressure, have approximately the same compressibility factor and all deviate from ideal gas behavior to about the same degree.

Material constants that vary for each type of material are eliminated, in a recast reduced form of a constitutive equation. The reduced variables are defined in terms of critical variables.

It originated with the work of Johannes Diderik van der Waals in about 1873 when he used the critical temperature and critical pressure to characterize a fluid.

The most prominent example is the van der Waals equation of state, the reduced form of which applies to all fluids.

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LIQUEFACTION OF GASES (lecture 7)

The processes are used for scientific, industrial and commercial purposes. Many gases can be put into a liquid state at normal atmospheric pressure by simple cooling; a few, such as carbon dioxide, require pressurization as well. Liquefaction is used for analyzing the fundamental properties of gas molecules (intermolecular forces), for storage of gases, for example: LPG, and in refrigeration and air conditioning. There the gas is liquefied in the *condenser*, where the heat of vaporization is released, and evaporated in the *evaporator,* where the heat of vaporization is absorbed. Ammonia was the first such refrigerant, but it has been replaced by compounds derived from petroleum and halogens.

Liquid oxygen is provided to hospitals for conversion to gas for patients with breathing problems, and liquid nitrogen is used in the medical field for cryosurgery, and by inseminators to freeze semen. Liquefied chlorine is transported for eventual solution in water, after which it is used for water purification, sanitation of industrial waste, sewage and swimming pools, bleaching of pulp and textiles and manufacture of carbon tetrachloride, glycol and numerous other organic compounds as well as phosgene gas.

Liquefaction of helium (^{4}He) with the Hampson-Linde cycle led to a Nobel Prize for Heike Kamerlingh Onnes in 1913. At ambient pressure the boiling point of liquefied helium is 4.22 K (-268.93°C). Below 2.17 K liquid ^{4}He becomes a superfluid (Nobel Prize 1978, Pyotr Kapitsa) and shows characteristic properties such as heat conduction through second sound, zero viscosity and the fountain effect among others.

The liquefaction of gases is a complicated process that uses various compressions and expansions to achieve high pressures and very low temperatures, using, for example, turbo expanders.

The liquefaction of air is used to obtain nitrogen, oxygen, and argon and other atmospheric noble gases by separating the air components by fractional distillation in an cryogenic air separation unit.

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2. **Essentials of Physical Chemistry by B. S. Bahl, G. D. Tuli, Golden Jubilee Ed., S. Chand and Company**
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COMPRESSIBILITY FACTOR AND CRITICAL CONSTANTS (lecture 8)

The **compressibility factor** (**Z**), also known as the **compression factor**, is the ratio of the molar volume of a gas to the molar volume of an ideal gas at the same temperature and pressure. It is a useful thermodynamic property for modifying the ideal gas law to account for the real gas behavior.[1] In general, deviation from ideal behavior becomes more significant the closer a gas is to a phase change, the lower the temperature or the larger the pressure. Compressibility factor values are usually obtained by calculation from equations of state (EOS), such as the virial equation which take compound specific empirical constants as input. For a gas that is a mixture of two or more pure gases (air or natural gas, for example), the gas composition must be known before compressibility can be calculated. Alternatively, the compressibility factor for specific gases can be read from generalized compressibility charts[1] that plot Z as a function of pressure at constant temperature.

As for the compressibility of gases, the principle of corresponding states indicates that any pure gas at the same reduced temperature, T_r, and reduced pressure, P_r, should have the same compressibility factor.

The reduced temperature and pressure are defined by

$$T_r = \frac{T}{T_c} \text{ and } P_r = \frac{P}{P_c}.$$

Here T_c and P_c are known as the critical temperature and critical pressure of a gas

References

1. **Martins Physical Pharmacy and Pharmaceutical Sciences, 5/Ed., Patric J. Sinka, Lippincott Williams and Wilkins**
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PRINCIPLE OF AEROSOL-TWO AND THREE PHASE SYSTEM (lecture 9)

Pharmaceutical aerosols are products that are packaged under pressure and contain therapeutically active ingredients that are released upon activation of an appropriate valve system. They are intended for topical application to the skin as well as local application into the nose (nasal aerosols), mouth (lingual aerosols), or lungs (inhalation aerosols). These products may be fitted with valves enabling either continuous or metered-dose delivery; hence, the terms “[DRUG] Metered Topical Aerosols,” “[DRUG] Metered Nasal Aerosols,” etc. The term “aerosol” refers to the fine mist of spray that results from most pressurized systems. The basic components of an aerosol system are the container, the propellant, the concentrate containing the active ingredient(s), the valve, and the actuator. The nature of these components determines such characteristics as particle size distribution, uniformity of dose for metered valves, delivery rate, wetness and temperature of the spray, spray pattern and velocity or plume geometry, foam density, and fluid viscosity.

Types of Aerosols: Aerosols consist of two-phase (gas and liquid) or three-phase (gas, liquid, and solid or liquid) systems. I) The two-phase aerosol consists of a solution of active ingredients in liquefied propellant and the vaporized propellant. The solvent is composed of the propellant or a mixture of the propellant and cosolvents such as alcohol, propylene glycol, and polyethylene glycols, which are often used to enhance the solubility of the active ingredients. ii) Three-phase systems consist of a suspension or emulsion of the active ingredient(s) in addition to the vaporized propellants. A suspension consists of the active ingredient(s) that may be dispersed in the propellant system with the aid of suitable excipients such as wetting agents and/or solid carriers such as talc or colloidal silicas.

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2. **Essentials of Physical Chemistry by B. S. Bahl, G. D. Tuli, Golden Jubilee Ed., S. Chand and Company**
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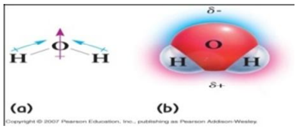
VARIOUS BINDING FORCES (lecture 10)

Intermolecular forces are weaker than ionic or covalent bonds and are responsible for the physical state of a compound (solid, liquid or gas).

Types of intermolecular forces

- Van der Waals Forces
    - Dipole Interactions
    - Dispersion Forces
- Hydrogen Bonds

Van der Waals Forces



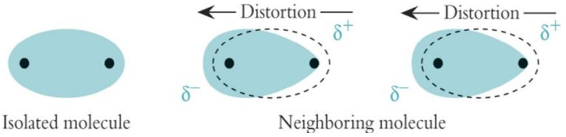
- They are the weakest attractions between molecules.
- Dipole: A polar molecule that has two poles.

Van der Waals-Dipole Interactions

- **Electrostatic** interaction between the oppositely charged regions of polar molecules (dipoles).

Which Molecules have Dipole Interactions?

Van der Waals Forces-**Dispersion** Forces

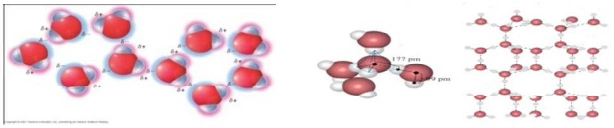


- **Dispersion** Forces
    - Caused by the motion of electrons.
    - Increase as the number of electrons increases.
    - Weakest of all intermolecular forces.

Hydrogen Bonding

- Hydrogen bonding is the attraction between a hydrogen atom of a molecule to an unshared pair of electrons in another molecule.
- Hydrogen bonding occurs in molecules where hydrogen is covalently bonded to a very electronegative element.
- Hydrogen bonding occurs in molecules containing N, O, F.
- Hydrogen bonds are the strongest of all intermolecular forces.
- Hydrogen bonds are possible because in hydrogen atoms there is no shielding of the nucleus.
- Hydrogen bonds are responsible for the physical properties of many biological substances and, more importantly, water.

Dipole Interactions



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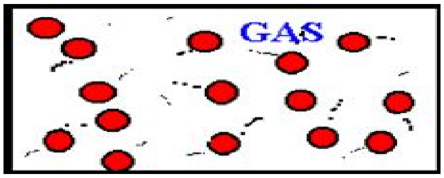
GASEOUS STATE (lecture 11)

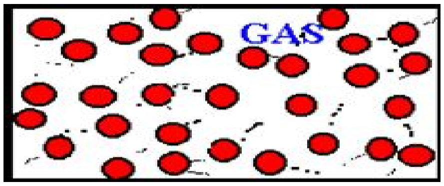
- **A gas has no fixed shape or volume, but always spreads out to fill any container.**
- There are almost **no forces of attraction between the particles** so they are completely free of each other. The particles are **widely spaced and scattered** at random throughout the container so there is no order in the system. The particles **move rapidly in all directions**, **frequently colliding** with each other and the side of the container. With **increase in temperature**, the **particles move faster as they gain kinetic energy**.

Using the particle model to explain the properties of a Gas

- **Gases have a very low density** ('light') because the particles are so spaced out in the container (density = mass / volume). **Density order: solid > liquid >>> gases**
- **Gases flow freely** because there are no effective forces of attraction between the gaseous particles - molecules. **Ease of flow order**: **gases > liquids >>> solids** (no real flow in solid unless you powder it!). Because of this **gases and liquids are described as fluids**.
- **Gases have no surface**, and **no fixed shape or volume**, and because of lack of particle attraction, they always spread out and fill any container (so gas volume = container volume).
- **Gases are readily compressed** because of the 'empty' space between the particles.
- **Ease of compression order**: **gases >>> liquids > solids** (almost impossible to compress a solid)
    - **Gas pressure**
- When a gas is confined in a container the particles will cause and exert a **gas pressure** which is measured in atmospheres (atm) or Pascals (Pa = N/m^2) - pressure is force/area on which force is exerted.
- The **gas pressure is caused by the force created by millions of impacts of the tiny individual gas particles on the sides of a container**.
- For example - **if the number of gaseous particles in a container is doubled, the gas pressure is doubled** because doubling the number of molecules doubles the number of impacts on the side of the container so the total impact force per unit area is also doubled. This doubling of the particle impacts doubling the pressure is pictured in the two diagrams below.



2 x particles ===> P x 2



For gas volume-pressure calculations

- **If the volume of a sealed container is kept constant and the gas inside is heated to a higher temperature, the gas pressure increases**.
    - The reason for this is that **as the particles are heated they gain kinetic energy and on average move faster**.
    - Therefore **they will collide with the sides of the container with a greater force of impact**, so increasing the pressure.

APPLICATION OF LIQUIFICATION (lecture 45)

H bonding

ice (ice is less dense than water because of this)

London Disp

when you need to compress noble gases, or non polar molecules (He, Ne, Ar, C2H4, propane,etc.)

Dipole

water is an example of this again. also HCl, etc. (pretty much anything ionically bonded)

coordinate

when mixing ammonia with a H+ ion, you get ammonium

The most important advantage of liquefying gases is that they can then be stored and transported in much more compact form than in the gaseous state. Two kinds of liquefied gases are widely used commercially for this reason, liquefied **natural gas** (LNG) and liquefied **petroleum** gas (LPG). LPG is a mixture of gases obtained from natural gas or petroleum that has been converted to the liquid state. The mixture is stored in strong containers that can withstand very high pressures. LPG is used as a fuel in motor homes, boats, and homes that do not have access to other forms of fuel.

Liquefied natural gas is similar to LPG, except that it has had almost everything except methane removed. LNG and LPG have many similar uses.

In principle, any gas can be liquefied, so their compactness and ease of transportation has made them popular for a number of other applications. For example, liquid **oxygen** and liquid **hydrogen** are used in rocket engines. Liquid oxygen and liquid acetylene can be used in **welding** operations. And a combination of liquid oxygen and liquid nitrogen can be used in aqualung devices.

Liquefaction of gases is also important in the field of research known as **cryogenics**. Liquid helium is widely used for the study of behavior of **matter** at temperatures close to absolute zero—0K (-459°F [-273°C]).

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INTRODUCTION TO CRYSTALLISATION (lecture 13)

Definition : **"Crystallization"** is a sequence of physical operations which allow to obtain in the form of a **crystalline solid** one or several substances initially contained in a liquid or gaseous phase

- ❖ Crystallization is one of the oldest **unit operations of thermal separation** used to prepare or concentrate a substance in the solid state

- ❖ Preliminary **step** of crystallization process :

= preparation of a **supersaturated solution** (= which contains "too much" dissolved solid)

Two ways for this….

- ➢ Solvent elimination
- ➢ **Shift of the equilibrium sokid-liquid equilibrium via temperature variation**
- ➢ **Crystallization** ≠ **precipitation** (involves chemical steps)

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CRYSTAL GROWTH VARIOUS CRYSTAL AND ITS SIZES (lecture 14)

An ideal solution one in which the molecules attract one another with equal force irrespective of their nature. Thus, a solution composed of two components A and B will be an ideal one if the forces between A and A, B and B should be the same. An ideal solution possesses the following characteristics:

(i) Volume change of mixing should be zero. $V_{solvent} + V_{solute} = V_{solution}$

(ii)Heat change on mixing should be zero. $?H_{mix} = 0$ (Heat is neither absorbed nor evolved).

(iii)There should be no chemical reaction between solvent and solute. solvent + solute → solution

$H_2O + NH_3 \rightarrow NH_4OH$
$H_2O + CO_2 \rightarrow H_2CO_3$ Non-ideal
$H_2O + CaO \rightarrow Ca(OH)_2$

(iv)Solute molecules should be not dissociate in the ideal solution.

Aqueous medium

$NaCl \longrightarrow Na^+ + Cl^-$

Non-ideal

Aqueous medium

$H_2SO_4 \longrightarrow 2H^+ + SO_4^{2-}$

(v)Solute molecules should not associate in ideal solution.

(vi)Ideal solutions must obey Raoult's law at all concentrations.

The following are some of the binary mixtures which show the properties of ideal solutions.

(a) Benzene and toluene. (b) Carbon tetrachloride, and silicon tetrachloride, (c) N-Hexane and n-heptane, Ethylene dibromide and tethylene dichloride.

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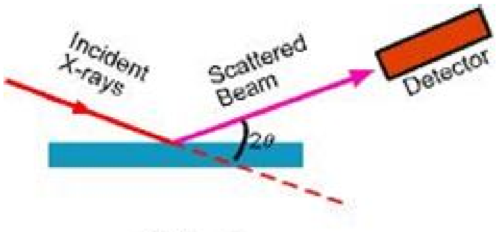
METHODS OF CRYSTAL ANALYSIS (lecture 15)

Powder Diffraction

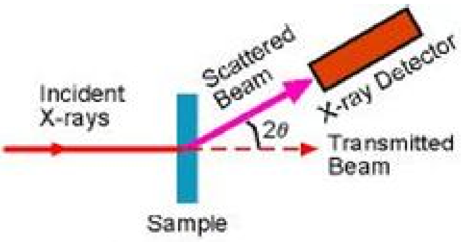
Powder XRD (X-ray Diffraction) is perhaps the most widely used x-ray diffraction technique for characterizing materials. As the name suggests, the sample is usually in a powdery form, consisting of fine grains of single crystalline material to be studied. The technique is used also widely for studying particles in liquid suspensions or polycrystalline solids (bulk or thin film materials).

The term 'powder' really means that the crytalline domains are randomly oriented in the sample. Therefore when the 2-D diffraction pattern is recorded, it shows concentric rings of scattering peaks corresponding to the various d spacings in the crystal lattice. The positions and the intensities of the peaks are used for identifying the underlying structure (or phase) of the material. For example, the diffraction lines of graphite would be different from diamond even though they both are made of carbon atoms. This phase identification is important because the material properties are highly dependent on structure (just think of graphite and diamond).

Powder diffraction data can be collected using either transmission or reflection geometry, as shown below. Because the particles in the powder sample are randomly oriented, these two methods will yield the same data. In the MRL x-ray facility, powder diffraction data are measured using the Philips XPERT MPD diffractometer, which measures data in reflection mode and is used mostly with solid samples, or the custom built 4-circle diffractometer, which operates in transmission mode and is more suitable for liquid phase samples.



Reflection



Transmission

A powder XRD scan from a $K_2Ta_2O_6$ sample is shown below as a plot of scattering intensity vs. the scattering angle 2θ or the corresponding d-spacing. The peak positions, intensities, widths and shapes all provide important information about the structure of the material.

References

1. **Martins Physical Pharmacy and Pharmaceutical Sciences, 5/Ed., Patric J. Sinka, Lippincott Williams and Wilkins**
2. **Essentials of Physical Chemistry by B. S. Bahl, G. D. Tuli, Golden Jubilee Ed., S. Chand and Company**

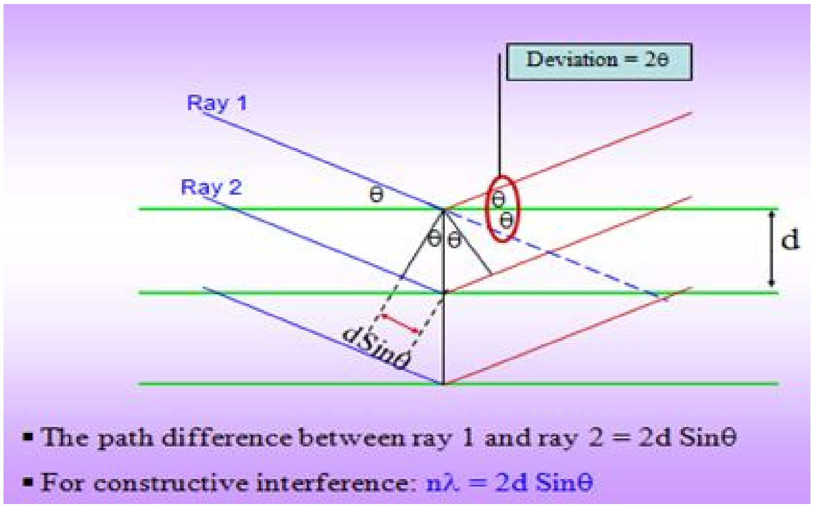
BRAGG’S EQUATION (lecture 16)

This law relates the wavelength of electromagnetic radiation to the diffraction angle and the lattice spacing in a crystalline sample. These diffracted X-rays are then detected, processed and counted. By scanning the sample through a range of 2θangles, all possible diffraction directions of the lattice should be attained due to the random orientation of the powdered material. Conversion of the diffraction peaks to d-spacings allows identification of the mineral because each mineral has a set of unique d-spacings. Typically, this is achieved by comparison of d-spacings with standard reference patterns. All diffraction methods are based on generation of X-rays in an X-ray tube. These X-rays are directed at the sample, and the diffracted rays are collected. A key component of all diffraction is the angle between the incident and diffracted rays.

Note that in the Bragg’s equation:

- The interatomic spacing (a) along the plane does not appear
- Only the interplanar spacing (d) appears

⇒ *Change in position or spacing of atoms along the plane should not affect Bragg’s condition !!*



- The path difference between ray 1 and ray 2 = $2d\ \mathrm{Sin}\theta$
- For constructive interference: $n\lambda = 2d\ \mathrm{Sin}\theta$

References

1. Martins Physical Pharmacy and Pharmaceutical Sciences, 5/Ed., Patric J. Sinka, Lippincott Williams and Wilkins
2. Essentials of Physical Chemistry by B. S. Bahl, G. D. Tuli, Golden Jubilee Ed., S. Chand and Company
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X-RAY DIFFRACTION (lecture 16)

- X-ray powder diffraction (XRD) is a rapid analytical technique primarily used for phase identification of a crystalline material and can provide information on unit cell dimensions. The analyzed material is finely ground, homogenized, and average bulk composition is determined.
- For electromagnetic radiation to be diffracted the spacing in the grating should be of the same order as the wavelength
- In crystals the typical interatomic spacing ~ 2-3 Å so the suitable radiation is X-rays
- Hence, X-rays can be used for the study of crystal structures
- A beam of X-rays directed at a crystal interacts with the electrons of the atoms in the crystal
- The electrons oscillate under the influence of the incoming X-Rays and become secondary sources of EM radiation
- The secondary radiation is in all directions
- The waves emitted by the electrons have the same frequency as the incoming X-rays ⇒ *coherent*
- The emission will undergo constructive or destructive interference with waves scattered from other *atoms*

References

1. **Martins Physical Pharmacy and Pharmaceutical Sciences, 5/Ed., Patric J. Sinka, Lippincott Williams and Wilkins**
2. **Essentials of Physical Chemistry by B. S. Bahl, G. D. Tuli, Golden Jubilee Ed., S. Chand and Company**
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METHODS FOR POWDER DIFFRACTION AND PROBLEMS (lecture 17)

Monochromatic X-rays → Many θs (orientations) Powder specimen → POWDER METHOD

Monochromatic X-rays → θ Varied by rotation → ROTATING CRYSTAL METHOD

- Powders are billions of tiny crystallites in all possible orientations.When placed in an x-ray beam, all possible interatomic planes will be seen. By systematically changing the experimental angle, we will produce all possible diffraction peaks from the powder.

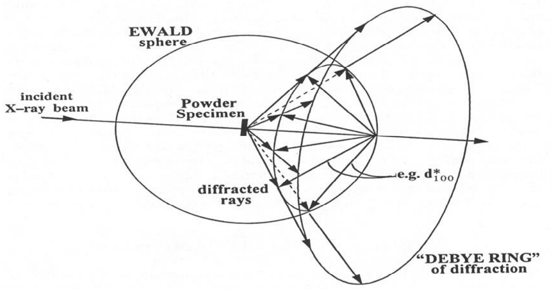

Figure 3.5. The intersection of d_{100} vectors from a powder with the Ewald sphere.

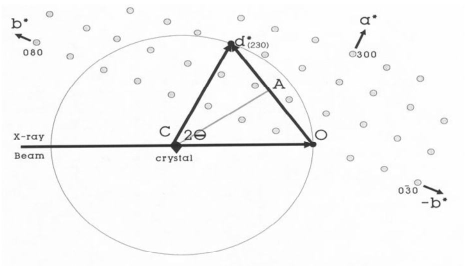

Figure 3.7. The Ewald sphere of reflection with the crystal rotated so that the (230) reciprocal lattice point touches it, permitting it to diffract.

- Each crystallite located in the center of the sphere has its own reciprocal lattice with its orientation determined by the orientation of the crystallite with respect to the X-ray beam
- Think of the diffractometer as a device for measuring diffractions occurring along the sphere – it’s function is to move all of the crystallites in the powder and their associated reciprocal lattices, measuring diffractions as they intersect the sphere
- Because of the operational geometry of diffractometers, there must be a very large number of small crystallites for the diffractometer to “see” all of the possible diffractions.
- By convention diffraction angles are recorded as 2θ. Data are commonly recorded as 2θ and intensity

References

1. **Martins Physical Pharmacy and Pharmaceutical Sciences, 5/Ed., Patric J. Sinka, Lippincott Williams and Wilkins**
2. **Essentials of Physical Chemistry by B. S. Bahl, G. D. Tuli, Golden Jubilee Ed., S. Chand and Company**
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POLYMORPHISM INTRODUCTION (lecture 18)

Polymorphism in crystalline solid is defined as materials with the same chemical composition different lattice structures and/or different molecular composition. Pseudopolymorphism is a term that refers to crystalline forms with solvent molecules as an integral part of the structure. Knowledge of crystal structure has also been applied to further understand chemical stability and dehydration or solvent loss. There is a renewed interest in polymorph; this is partly due to increased economic pressure faced by pharmaceutical companies and the greater awareness of the effect that polymorphs may have on the bioavailability, manufacturability and stability of the product.

References

1. **Martins Physical Pharmacy and Pharmaceutical Sciences, 5/Ed., Patric J. Sinka, Lippincott Williams and Wilkins**
2. **Essentials of Physical Chemistry by B. S. Bahl, G. D. Tuli, Golden Jubilee Ed., S. Chand and Company**
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APPLICATIONS (lecture 33)

Electrolytic solutions conducts electricity due to presence of cations and anions. They are used in electrical cells,electroplating,extraction of metals,amalgamation etc. Some non electrolytes may be used as solvents for organic solutes.

Purification and extraction of metals. Such as copper and aluminium

Non electrolyte solutions is a strange description. Liquids that don't conduct electricity are used for insulation in electrical systems such as transformers switchgear etc.

References

1. **Martins Physical Pharmacy and Pharmaceutical Sciences, 5/Ed., Patric J. Sinka, Lippincott Williams and Wilkins**
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VARIOUS TYPES OF SOLUTIONS (lecture 22)

- **Solution** - a homogeneous mixture
  - **Solute:** the component that is dissolved
  - **Solvent:** the component that does the dissolving

Generally, the component present in the greatest quantity is considered to be the solvent. *Aqueous* solutions are those in which *water* is the solvent.

- Electrolytes and Nonelectrolytes
  - ***Electrolyte***: substance that dissolved in water produces a solution that conducts electricity.
  - Contains ions $\mathrm{NaCl}(s) \xrightarrow{\mathrm{H_2O}} \mathrm{Na^+}(aq) + \mathrm{Cl^-}(aq)$
  - ***Nonelectrolyte***: substance that dissolved in water produces a solution that does not conduct electricity
  - Does not contain ions

$$\mathrm{C_{12}H_{22}O_{11}}(s) \xrightarrow{\mathrm{H_2O}} \mathrm{C_{12}H_{22}O_{11}}(aq)$$

- ***Dissociation*** - ionic compounds separate into constituent ions when dissolved in solution

$$\mathrm{NaCl}(s) \xrightarrow{\mathrm{H_2O}} \mathrm{Na^+}(aq) + \mathrm{Cl^-}(aq)$$

- ***Ionization*** - formation of ions by molecular compounds when dissolved

$$\mathrm{HCl}(g) \xrightarrow{\mathrm{H_2O}} \mathrm{H^+}(aq) + \mathrm{Cl^-}(aq)$$

- Strong and weak electrolytes

  Strong Electrolyte: 100% dissociation

  All water soluble ionic compounds, strong acids and strong bases
- ***Weak electrolytes***

  Partially ionized in solution

  Exist mostly as the molecular form in solution

  Weak acids and weak bases

References

1. **Martins Physical Pharmacy and Pharmaceutical Sciences, 5/Ed., Patric J. Sinka, Lippincott Williams and Wilkins**
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PROPERTIES OF SOLUTIONS (lecture 23)

An ideal solution one in which the molecules attract one another with equal force irrespective of their nature. Thus, a solution composed of two components A and B will be an ideal one if the forces between A and A, B and B should be the same. An ideal solution possesses the following characteristics:

(i) Volume change of mixing should be zero. $V_{solvent} + V_{solute} = V_{solution}$

(ii)Heat change on mixing should be zero. $H_{mix} = 0$ (Heat is neither absorbed nor evolved).

(iii)There should be no chemical reaction between solvent and solute. solvent + solute → solution

$H_2O + NH_3 \rightarrow NH_4OH$

$H_2O + CO_2 \rightarrow H_2CO_3$ Non-ideal

$H_2O + CaO \rightarrow Ca(OH)_2$

(iv)Solute molecules should be not dissociate in the ideal solution.

Aqueous medium

$NaCl \longrightarrow Na^+ + Cl^-$

Non-ideal Aqueous medium

$H_2SO_4 \longrightarrow 2H^+ + SO_4^{2-}$

(v)Solute molecules should not associate in ideal solution.

(vi) Ideal solutions must obey Raoult's law at all concentrations.

The following are some of the binary mixtures which show the properties of ideal solutions.

(a) Benzene and toluene. (b) Carbon tetrachloride, and silicon tetrachloride, (c) N-Hexane and n-heptane, Ethylene dibromide and tethylene dichloride.

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1. **Martins Physical Pharmacy and Pharmaceutical Sciences, 5/Ed., Patric J. Sinka, Lippincott Williams and Wilkins**
2. **Essentials of Physical Chemistry by B. S. Bahl, G. D. Tuli, Golden Jubilee Ed., S. Chand and Company**
3. **Essentials of Physical Chemistry and Pharmacy, H. J. Arnikar, S. S. Kadam, K. N. Gujar, Orient Longman Pvt. Ltd, India**

IDEAL AND REAL SOLUTIONS (lecture 24)

An ideal solution can be said to follow Raoult's Law but it must be kept in mind that in the strict sense ideal solutions do not exist. The fact that the vapor is taken to be ideal is the least of our worries. Interactions between gas molecules are typically quite small especially if the vapor pressures are low. The interactions in a liquid however are very strong. For a solution to be ideal we must assume that it does not matter whether a molecule A has another A as neighbor or a B molecule. This is only approximately true if the two species are almost identical chemically.

Raoult's Law may be adapted to non-ideal solutions by incorporating two factors that will account for the interactions between molecules of different substances. The first factor is a correction for gas non-ideality, or deviations from the ideal-gas law.

References

1. **Martins Physical Pharmacy and Pharmaceutical Sciences, 5/Ed., Patric J. Sinka, Lippincott Williams and Wilkins**
2. **Essentials of Physical Chemistry by B. S. Bahl, G. D. Tuli, Golden Jubilee Ed., S. Chand and Company**
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VARIOUS CONCENTRATION TERMS (lecture 25)

"Parts-per" concentration:

One common method of expressing the concentration is based on the quantity of solute in a fixed quantity of solution. The "quantities" referred to here can be expressed in weight, in volume, or both (the weight of solute in a given volume of solution.) In order to distinguish among these posibilities, the abbreviations (w/w), (v/v) and (w/v) are used.

Percent:

Means "parts per 100"; we can also use parts per thousand (ppt) for expressing concentrations in grams of solute per kilogram of solution. For more dilute solutions, parts per million (ppm) and parts per billion (10^{9}; ppb) are used. These terms are widely employed to express the amounts of trace pollutants in the environment.

Weight/volume and volume/volume basis:

It is sometimes convenient to base concentration on a fixed volume, either of the solution itself, or of the solvent alone. In most instances, a "5% by volume" solution of a solid will mean 5 g of the solute dissolved in 100 ml of the solvent.

Molarity: mole/volume basis:

This is the method most used by chemists to express concentration. Molar concentration (mo larity) is the number of moles of solute per liter of solution . The important point to remember is that the volume of the solution is different from the volume of the solvent; the latter quantity can be found from the molarity only if the densities of both the solution and of the pure solvent are known. Similarly, calculation of the weight-percentage concentration from the molarity requires density information; you are expected to be able to carry out these kinds of calculations, which are covered in most texts.

Mole fraction: mole/mole basis:

This is the most fundamental of all methods of concentration measure, since it makes no assumptions at all about volumes.

Molality:

mole/weight basis. A 1-molal solution contains one mole of solute per 1 kg of solvent. . Molality is a hybrid concentration unit, retaining the convenience of mole measure for the solute, but expressing it in relation to a temperature-independent mass rather than a volume.

Molality:

like mole fraction, is used in applications dealing with certain physical properties of solutions;

References

1. **Martins Physical Pharmacy and Pharmaceutical Sciences, 5/Ed., Patric J. Sinka, Lippincott Williams and Wilkins**
2. **Essentials of Physical Chemistry by B. S. Bahl, G. D. Tuli, Golden Jubilee Ed., S. Chand and Company**
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ROULTS LAW AND ITS DEVIATION (lecture 26)

The partial vapour pressure of each volatile constituent is equal to vapour pressure of pure constituent multiplied by mole fraction.

$$P_A=P_A^oX_A, \quad P_B = P_B^O X_B$$

Raoult's law states that the relatives lowering of the vapour pressure of a dilute solution is equal to the mole fraction of the solute present in a dilute solution and Raoult's law can be expressed mathematically in the form.

$$\frac{P-Ps}{P} = \frac{n}{n+N}$$

Where, n = number of moles or molecules of solute.

N = number of moles & molecules of solvent.

A solution which obeys Raoult's law strictly called ideal solution and a solution which shows deviation from Raoult's law is called non-ideal or real solution.

Suppose, the molecules of solvent and solute are represented by 'A' and 'B' respectively. Now let 'ʃAB' be the attractive force between 'A' and 'B' and ʃAA between 'A' and 'A'

ʃAB = ʃAA

It is an ideal solution. How ever , if

ʃAB = ʃAA

Molecule a will escape less readily and vapour pressure will be less than that predicted by Raoult's law (Negative deviation).A molecule will escape from solution surface more readily and the vapour pressure of solution will be higher than predicted by Raoult's law (positive deviation)

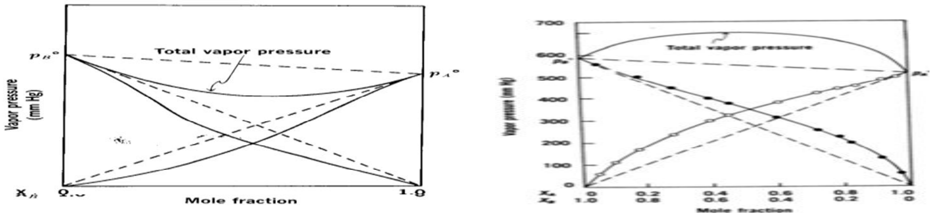

Fig – Negative and positive deviation from Raoult's law

References

1. **Martins Physical Pharmacy and Pharmaceutical Sciences, 5/Ed., Patric J. Sinka, Lippincott Williams and Wilkins**
2. **Essentials of Physical Chemistry by B. S. Bahl, G. D. Tuli, Golden Jubilee Ed., S. Chand and Company**
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B.P.DETERMINATION (lecture 27)

The boiling point of a compound is the temperature at which it changes from a liquid to a gas. This is a physical property often used to identify substances or to check the purity of the compound.

It is difficult, though, to find a boiling *point*. Usually, chemists can only obtain a boiling range of a 2 - 3^{O}C accuracy. This is usually sufficient for most uses of the boiling point.

Purpose:

The purpose of this experiment is to determine the boiling points of various organic compounds and to use these to identify unknowns.

References

1. **Martins Physical Pharmacy and Pharmaceutical Sciences, 5/Ed., Patric J. Sinka, Lippincott Williams and Wilkins**
2. **Essentials of Physical Chemistry by B. S. Bahl, G. D. Tuli, Golden Jubilee Ed., S. Chand and Company**
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COLLIGATIVE PROPERTIES (lecture 28)

Properties of solutions that depend on the number of molecules present and not on the kind of molecules are called colligative properties. These properties include boiling Point elevation, freezing point depression, and osmotic pressure. Historically, colligative properties have been one means for determining the molecular weight of unknown compounds. In this chapter we discuss using colligative properties to measure the molecular weight of polymers. Because colligative properties depend on the number of molecules, we expect, and will show, that colligative property experiments give a number average molecular weight.

References

1. **Martins Physical Pharmacy and Pharmaceutical Sciences, 5/Ed., Patric J. Sinka, Lippincott Williams and Wilkins**
2. **Essentials of Physical Chemistry by B. S. Bahl, G. D. Tuli, Golden Jubilee Ed., S. Chand and Company**
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INTRODUCTION TO CONDUCTANCE (lecture 29)

Power of electrolytes to conduct electric current is called conductivity or conductance of electrolytes.

CONDUCTOMETRIC TITRATION

Titration in which conductance measurements are made use of in determining the end point of acid-alkali reaction and some displacement reaction or precipitation reaction are called **conduct metric titration.**

1. Titration of strong acid against strong base-

The reason for this is that before the addition of alkali the conductance of the solution is due to presence of H^+ and Cl^- ions. End point is conductance against volume of base.

2. Titration of weak acid against a strong alkali-

The initial conductance of the solution is low because of the poor dissociation of the weak acid on adding alkali; highly ionized sodium acetate is formed. The conductance value shows sharp increases. End point is conductance against volume of base.

3 .Titration of strong acid against a weak base

The curve obtained for the titration of the titration of the strong acid against a weak base. The conductance of the solution will first decrease due to the fixing up of a fast moving H^+ ions and their replacement by slow moving NH_4^+ ions. End point is conductance against volume of ammonium hydroxide.

4. Titration of weak acid against weak base-

Conduct metric method is particularly suitable as such titration do not give a sharp end point with indicator. Initial conductance of the solution in this case is also low due to the poor dissociation of the weak acid. End point is conductance against volume of ammonium hydroxide.

5. Precipitation reaction-

End point in precipitation reaction can be accurately determined by conduct metric titration. In the titration of potassium chloride against silver nitrate. End point is conductance against volume of potassium chloride.

References

1. **Martins Physical Pharmacy and Pharmaceutical Sciences, 5/Ed., Patric J. Sinka, Lippincott Williams and Wilkins**
2. **Essentials of Physical Chemistry by B. S. Bahl, G. D. Tuli, Golden Jubilee Ed., S. Chand and Company**
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SPECIFIC AND EQIVALENT CONDUCTANCE (lecture 30)

1. Specific Conductance-

The reciprocal of specific resistance is termed specific conductance.

It is denoted by k(kappa). Unit=mhos or ohm^{-1}

2 .Equivalent Conductance

The conductance of electrolyte obtained by dissolving one gram of in water.

It is denoted by ^. Unit = ohm^{-1}cm^{2}equt^{-1}.

Advantages of conduct metric titration

1) This method is useful for the titration of weak acid against weak base which do not give a sharp change of color with indicator in ordinary volumetric analysis.

2) More accurate results are obtained because the end point is determined graphically.

References

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2. **Essentials of Physical Chemistry by B. S. Bahl, G. D. Tuli, Golden Jubilee Ed., S. Chand and Company**
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ARRHENIUS THEORY (lecture 31)

We know that the conductivity of a solution is as a result of the mobility of cations and anions in aqueous solution. The greater the number of ions, the greater is the conductivity. The degree of dissociation is defined as the fraction of the total number of molecules dissociated into ions. Strong electrolytes are used to express substances that completely ionize when dissolved with no neutral molecules formed in solution. A good example is the ionic solid NaCl. The solute of NaCl is completely ionized because no molecules of NaCl are present in NaCl solid or NaCl solution.

$NaCl\ (s) \rightarrow Na^+\ (aq) + Cl^-\ (aq)$

Some other examples of ionic solids are $CuSO_4$, NH_4Cl, KBr, $NaCH_3COO$ and $NaHCO_3$.

Weak electrolytes only partially ionize hence some neutral molecules are still present in their solutions (in this equation below, there are still H_2CO_3 molecules present).

$H_2CO_3\ (aq) \rightarrow H^+(aq) + HCO_3^-\ (aq)$

Examples are H_2CO_3, NH_3, NH_4OH, CH_3COOH and most organic acids and bases.

The Arrhenius theory had limitations interpreting for strong electrolytes. Here are some of the anomalies:

1. Arrhenius theory assumes ionization of electrolytes only in aqueous solution but x-ray diffraction studies prove that crystalline alkali halides (e.g. NaCl) exists in ionized forms even in the solid state.
2. Strong electrolytes (e.g. NaOH, KCl) conduct electricity in molten states also which means that their ionization must have taken place in the absence of water
3. Arrhenius theory assumes the existence of ions in solution but it does not explain how or why ions are formed in solution

References

1. **Martins Physical Pharmacy and Pharmaceutical Sciences, 5/Ed., Patric J. Sinka, Lippincott Williams and Wilkins**
2. **Essentials of Physical Chemistry by B. S. Bahl, G. D. Tuli, Golden Jubilee Ed., S. Chand and Company**
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DEBYE HUCKLE THEORY (lecture 32)

Debye and Hukel derived an equation based upon the principle that strong electrolytes are completely ionized in dilute solution and that the deviations of electrolytic solutions from ideal behaviour are due to the electrostatic effects of the oppositely charged ions. The equation relates the activity coefficient of a particular ion or the mean ionic activity coefficient of an electrolyte to the valence of the ions, the ionic strength of the solution, and the characteristics of the solvent.

According to the theory of Debye and Hukel, the activity coefficient γ_i of an ion of valence z_i is given by the expression

$$\log \gamma_i = -Az_i^2\sqrt{\mu}$$

Yields a satisfactory measure of the activity coefficient of an species upto an ionic strength μ of about 0.02.

The form of Debye-Huckel equation for a binary electrolyte, consisting of an ion with valances of z_+ and z_- and present in a dilute solution ($\mu<0.02$), is

$$\log \gamma_i = -Az_+ z_- \sqrt{\mu}$$

The symbol of z_+ and z_- stand for the valences or charges

This formula applies up to an ionic strength of perhaps 0.1.

$$\mathrm{Log}\ \gamma = -\frac{Az+z_-\sqrt{\mu}}{1+a_iB\sqrt{\mu}}$$

Further higher concentration is , at ionic strength above 0.1.

$$\mathrm{Log}\ \gamma = -\frac{Az+z_-\sqrt{\mu}}{1+a_iB\sqrt{\mu}} + C\mu$$

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VARIOUS EXAMPLES AND APPLICATIONS (lecture 19)

Most important application of X-ray crystallography is its use in synthesizing (blending) substances. Many of the medicinal chemicals that have been discovered by scientists are very difficult to produce naturally in large amounts. When this happens, it becomes necessary to create the chemicals in the laboratory through synthesis. Before a chemist can synthesize a substance, a map of its atomic structure must be obtained. This map can only be drawn by using X-ray crystallography.

Few scientists have been more successful at this, Hodgkin and her colleagues determined the structure of **penicillin.** The synthesis of this drug was necessary for mass wartime production. Since then Hodgkin's team has worked on the mapping of **vitamin B12,** the vitamin prescribed to prevent pernicious anemia (a chronic blood disorder characterized by weakness and pallor). The team also worked on mapping **insulin.** Insulin is used in the treatment of diabetes (another blood disorder).

Other researchers have used X-ray technologies to record the structures of proteins, hemoglobin, and the double-helix of DNA structure (deoxyribonucleic acid).

Creating a New Science

The development of X-ray crystallography also created the science of mineralogy. Once the inner structures of many minerals were determined, mineralogists were able to define the major mineral groups. The understanding that stems from crystallography has also allowed scientists to construct the man-made minerals used in industry.

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EFFECT OF pH AND SOLVENT (lecture 44&45)

The solubility as a function of pH, temperature, and solvent composition was examined to determine the pK_a of the salt from the solubility profile at various temperatures and in several solvent systems, for example the effect of alcohol on the solubility of phenobarbital. The results showed that the pKa of Phenobarbital, 7.41, is raised to 7.92 in a hydroalcoholic solution containing 30% by volume of alcohol. Also the solubility, $S_\circ$ of un-ionized phenobarbital is increased from 0.12 g/100 mL (0.005 M) in water to 0.64% (0.28 M) in a 30% alcoholic solution.

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DISTRIBUTION PHENOMENON (lecture 34)

Partition coefficient: rate of movement of X in aqueous phase to rate of movement organic phase until the equilibrium achieved

$$K\,\frac{conc\ \ of\ X\ in\ solvent}{conc\ \ 2of\ X\ in\ solvent}$$

K is also known as distribution ratio, distribution coefficient or partition coefficient

Conc. of acid in ether(C_1)	Conc. of acid in water(C_2)	C_1/C_2
0.0046	0.024	0.191
0.013	0.069	0.188
0.022	0.119	0.185

Also called as Nenst equation /Partition coefficient/distribution phenomenon

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DISSOCIATION AND ASSOCIATION (lecture 35)

When interaction occurs between like molecules of one of the component in a solution is referred as 'association'.

The solute may exit partly or wholly as associated molecules in one of the phase or it may dissociate into ions in either of the liquid phases. The distribution law applies only to concentration of the species common to both the phases, namely 'monomer'.

Consider distribution of benzoic acid between an oil phase and water phase. When it is neither associated in oil nor dissociated in ion in the water then,

$$\frac{C_1}{C_2} = K$$

Can be used to compute the distribution constant

According to Garrett and Woods benzoic acid is considered to be distributed between two phases, peanut oil and water. Although benzoic acid undergoes dimerization in many dipolar solvent it does not associate in peanut oil. It ionizes into water to a degree, however depending on pH of the solution. The species common to both the oil and water are unassociated and undissociated benzoic acid molecule. Distribution is expressed as:

$$K = \frac{[HA]_O}{[HA]_W} = \frac{C_O}{[HA]_W}$$

In which K is *true distribution coefficient*

$[HA]_o = C_O$ is molar concentration of simple benzoic acid molecule in the oil phase

$[HA]_w$=molar concentration of undissociated acid in the water phase the total acid concentration obtained by analysis of aqueous phase is

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APPLICATIONS OF DISRTIBUTION PHENOMENON (lecture 36)

1. Extraction –

To determine the efficiency with which one solvent can extract a compound from a second solvent.

$$w_n = w\left(\frac{KV_1}{KV_1+V_2}\right)^n$$

2. Solubility and partition coefficient –

To determine the aqueous solubility of liquid or crystalline organic compound.

$$\log S = -\log K - 1.11\frac{\Delta S_f(mp-25)}{1364} + 0.54$$

3. Preservative action of weak acids in oil-water system –

Solution of foods, drugs and cosmetics are subject to deterioration by enzyme of microorganism that act as a catalyst in decomposition reactions. Sterilization and addition of chemical preservatives are common method used in pharmacy to preserve drug solution against attack by various microorganisms.

4. Drug action and partition coefficient –

Mayer and Overton propose the hypothesis that narcotic action of a non specific drug is a function of distribution coefficient of compound between a lipoidal medium and water.

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SOLUTE SOLVENT INTERACTIONS (lecture 37)

- **- Solubility depends on chemical, electrical & structural effects that lead to mutual** interactions between the solute and the solvent.
- - In pre-or early formulation, selection of the most suitable solvent is based on the principle of **"like dissolves like". That is, a solute dissolves best in a solvent** with similar chemical properties. i.e. Polar solutes dissolve in polar solvents. E.g **salts & sugar dissolve in water .** Non polar solutes dissolve in non polar solvents. Eg**. naphtalene dissolves in benzene.**
- **- To explain the above rule, consider the forces of attraction between solute and solvent molecules.**
- If the solvent is **A & the solute is B, and t**he forces of attraction are represented by **A-A, B-B and A-B,**one of the following conditions will occur:
- 1. If **A-A >> A-B The solvent molecules will be attracted to each other &** the solute will be excluded. Example: Benzene & water, where benzene molecules are unable to penetrate the closely bound water aggregates.
- 2. If **B-B >> A-A The solvent will not be able to break the binding forces** between solute molecules. Example: NaCl in benzene, where the NaCl crystal is held by strong electrovalent forces which cannot be broken by benzene.
- 3. If **A-B >> A-A or B-B, or the three forces are equal The solute will** disperse & form a solution. Example: NaCl in water.

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SOLUBILITY OF GAS, SOLID, LIQUIDS (lecture 38 & 39)

Solubility of gases in liquids

Examples of pharmaceutical solutions of gases include: HCl, ammonia water & effervescent preparations containing CO2 maintained in solution under pressure.

- The solubility of a gas in a liquid ***is the concentration of dissolved gas when it is in equilibrium with some of the pure gas above the solution.***
- The solubility depends on the pressure, temperature, presence of salts & chemical reactions that sometimes the gas undergoes with the solvent

Solubility of liquids in liquids

Preparation of pharmaceutical solutions involves mixing of 2 or more liquids (alcohol & water to form hydroalcoholic solutions, volatile oils & water to form aromatic waters, volatile oils & alcohols to form spirits …)

- Liquid-liquid systems may be divided into 2 categories:
- 1) Systems showing *complete miscibility such as alcohol & water, glycerin* & alcohol, benzene & carbon tetrachloride..
- 2) Systems showing *Partial miscibility as phenol and water;* two liquid layers are formed each containing some of the other liquid in the dissolved state.
- ***Complete miscibility occurs when: The adhesive forces between different molecules***
- (A-B) **>>cohesive forces between like molecules (A-A or B-B)**

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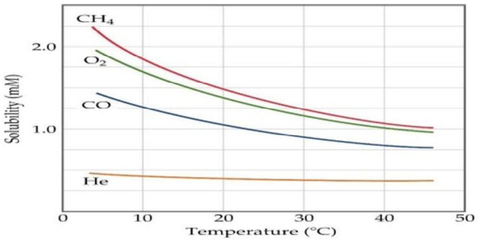
FACTORS AFFECTING SOLUBILITY (lecture 40)

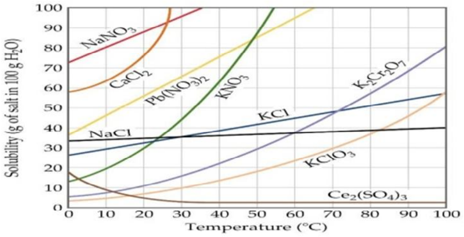
1. Nature of solute and solvent:
   Like dissolves like, i.e. same polarity
   - Polar solutes are soluble in polar solvents.
   - Nonpolar solutes are soluble in nonpolar solvents.

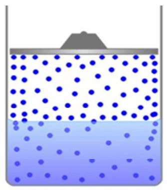
The amount of solute that dissolves depends on what type of solute it is. While only 1 gram of lead (II) chloride can be dissolved in 100 grams of water at room temperature, 200 grams of zinc chloride can be dissolved. This means that a greater amount of zinc chloride can be dissolved in the same amount of water than lead II chloride.

2. Temperature Effects: Generally, an increase in the temperature of the

solution increases the solubility of a solid solute. For example, a greater amount of sugar will dissolve in warm water than in cold water. A few solid solutes, however, are less soluble in warmer solutions. For all gases, solubility decreases as the temperature of the solution rises. An example of this is Soda. The solubility of the carbon dioxide gas decreases when a soda is warm, making the soda flat.







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3. Pressure: For solid and liquid solutes, changes in pressure have practically no

effect on solubility. For gaseous solutes, an increase in pressure increases solubility and a decrease in pressure decreases solubility. Example: When the cap on a bottle of soda pop is removed, pressure is released, and the gaseous solute bubbles out of solution. This escape of a gas from solution is called **effervescence**.

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SOLUBILITY OF SLIGHTLY SOLUBLE ELECTROLYTES (lecture 41)

In the case of the sparingly soluble compound part of the compound that has been dissolved, is in solution virtually completely dissociated into ions. Between the precipitate (solid) and its aqueous solution exists dynamic equilibrium, this means that the rates of dissolution and precipitation are equal to one another, e.g.

$AgCl\downarrow \leftrightarrow Ag^+ + Cl^-$

This is a reversible reaction, and thus the thermodynamic equilibrium expression is:

$$K_{rozp} = \frac{[Ag^+][Cl^-]}{[AgCl]}$$

Since this is a solid phase, the [AgCl] is constant and then K_{eg} $[AgCl] = K_{sp}$ that is, $K_{sp} = [Ag^+][Cl^-]$.

K_{sp} is called the solubility product of sparingly soluble compound.

In the saturated solution of sparingly soluble electrolyte (compound) the product of the numerical values of the concentrations of the ions, to which the electrolyte (compound) dissociates, is constant at given temperature and is called the solubility product.

If $[Ag^+][Cl^-] < K_{sp,}$ no precipitate is form. If $[Ag^+][Cl^-] > K_{sp,}$ a precipitate is form.

A measure of the solubility of a substance is its concentration in the saturated solution (at a given temperature).

Factors affecting the solubility of the precipitate:

1) temperature – temperature increase generally increases the solubility.
2) common ion effect e.g.: $K_{sp} = [Ba^{2+}][SO_4^{2-}]$ = const. Increasing the concentration of one of the ions Ba^{2+} or SO_4^{2-}, decreases the solubility.
3) salt effect.
4) pH effect
5) hydrolysis of precipitate .
6) complex formation reactions.

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SOLUBILITY OF WEAK ELECTROLYTES (lecture 42)

A weak electrolyte, on the other hand, will ionize only slightly in water and produce few ions, so their solutions are poor conductors of electricity. In a weak electrolyte, this dissociation into ions occurs only to a small extent, typically much less than 10%. The vast bulk of the weak electrolyte remains in its original unionized form in solution. In such cases, there are insufficient ions present to carry an electric current.

The electrolytic strength of a substance can readily be determined by measuring the electrical conductance of a solution of the substance of known concentration. Tables are available listing the Q_{sp} or ion constant of many substances, which is a measure of their degree of ionization.

Telling a weak electrolyte apart from a strong one just by inspecting its chemical formula is not so straightforward. A weak electrolyte will usually be composed of covalent bonds, or chemical bonds in which electrons are shared by two atoms.

In general, though, most organic acids and their salts, and the salts of organic bases, are weak electrolytes. A substance with low water solubility may also be classified as a weak electrolyte. Technically, however, solubility is not the same as electrolytic strength.

In addition to weak and strong electrolytes, non-electrolytes are substances that don't ionize to any appreciable extent in aqueous solution, and their solutions don't conduct electricity at all. Most organic substances, unless they contain an acid or base functionality, are non-conductors and therefore non-electrolytes. Sugar and alcohol, for example, are organic compounds with no acid or base functionality and will therefore not produce ions in solution.

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EFFECT OF pH,SOLVENT (lecture 43)

pH affect the solubility of slightly soluble compounds containing anions that are conjugate bases of weak acids, such as F^-, NO_2^-, OH^-, SO_3^{2-}, and PO_4^{3-}, but not those containing anions which are conjugate bases of strong acids, such as SO_4^{2-}, Cl^-, Br^-. For example, in a saturated solution of calcium fluoride, CaF_2, the following equilibrium exists:

$$CaF_{2(s)} \rightleftarrows Ca^{2+}_{(aq)} + 2F^-_{(aq)}$$

When a strong acid is added to the saturated solution, the following reaction occurs:

$$H^+_{(aq)} + F^-_{(aq)} \rightarrow HF_{(aq)}$$

This reaction has the net effect of reducing the concentration of F^- ions, which causes the equilibrium to shift to the right and resulting in more of CaF_2 to dissolve.

Weak electrolytes can behave like strong electrolytes or like nonelectrolytes in solution. When the solution is of such a pH that the drug is entirely in the ionic form, it behaves as a solution of a strong electrolyte, and solubility does not constitute a problem. However, when the pH is adjusted to a value at which un-ionized molecules are produced in sufficient concentration to exceed the solubility of this form, precipitation occurs.

Frequently, a solute is more soluble in a mixture of solvents than in one solvent alone. This phenomenon is known as *cosolvency*, and the solvents that, in combination, increase the solubility of the solute are called *cosolvents*.

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