

Factors affecting solubility

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There are three main factors that control solubility of a solute. (1) Temperature (2) Nature of solute or solvent (3) Pressure

EFFECT OF TEMPERATURE Generally in many cases solubility increases with the rise in temperature and decreases with the fall of temperature but it is not necessary in all cases. However we must follow two behaviours: In endothermic process, solubility increases with the increase in temperature and vice versa. For example: solubility of potassium nitrate increases with the increase in temperature. In exothermic process, solubility decrease with the increase in temperature.

For example: solubility of calcium oxide decreases with the increase in temperature. Gases are more soluble in cold solvent than in hot solvent.

NATURE OF SOLUTE AND SOLVENT Solubility of a solute in a solvent purely depends on the nature of both solute and solvent. A polar solute dissolved in polar solvent. Solubility of a non-polar solute in a solvent is large. A polar solute has low solubility or insoluble in a non-polar solvent.

EFFECT OF PRESSURE The effect of pressure is observed only in the case of gases. An increase in pressure increases of solubility of a gas in a liquid.

For example carbon dioxide is filled in cold drink bottles (such as coca cola, Pepsi 7up etc.) under pressure.

Properties of Solution Concentration The concentration of a solution is the measure of how much solute and solvent there is. A solution is concentrated if it contains a large amount of solute, or dilute if contains a small amount.

Molarity Molarity is the number of moles of solute per litre of solution. It is abbreviated with the symbol M, and is sometimes used as a unit of measurement, e. g. a 0.3 molar solution of HCl. In that example, there would be 3 moles of HCl for every 10 litres of water

(or whatever the solvent is). Molality Molality is the number of moles of solute per kilogram of solvent. It is abbreviated with the symbol m (lowercase), and is sometimes used as a unit of measurement, e. g. a 0.3 molal solution of HBr. In that example, there would be 3 moles of HBr for every 10 kilograms of water (or whatever the solvent was). Mole Fraction The mole fraction is simply the moles of solute per moles of solution. As an example, you dissolve one mole of NaCl into three moles of water. Remember that the NaCl will dissociate into its ions, so there are now five moles of particles: one mole Na^+ , one mole Cl^- , and three moles water.

The mole fraction of sodium is 0.2, the mole fraction of chloride is 0.2, and the mole fraction of water is 0.6. The mole fraction is symbolized with the Greek letter (χ), which is often written simply as an X. Dilution Dilution is adding solvent to a solution to obtain a less concentrated solution. Perhaps you have used dilution when running a lemonade stand. To cut costs, you could take a half-full jug of rich, concentrated lemonade and fill it up with water. The resulting solution would have the same total amount of sugar and lemon juice, but double the total volume. Its flavour would be weaker due to the added water.

The key concept is that the amount of solute is constant before and after the dilution process. The concentration is decreased (and volume increased) only by adding solvent. Thus, the number of moles of solute before and after dilution are equal. $\text{moles}_1 = \text{moles}_2$ By definition of molarity, you can find the moles of solvent. $M \times V = \text{moles}$ Substituting the second equation into the first gives the dilution equation. $M_1 \times V_1 = M_2 \times V_2$ To determine the

amount of solvent (usually water) that must be added, you must know the initial volume and concentration, and the desired concentration.

Solving for V_2 in the above equation will give you the total volume of the diluted solution. Subtracting the initial volume from the total volume will determine the amount of pure solvent that must be added.

Ionic Solutes

When ionic compounds dissolve in water, they separate into ions. This process is called dissociation. Note that because of dissociation, there are more moles of particles in the solution containing ions than there would be with the solute and solvent separated. If you have two glasses of water, and you dissolve salt into one and sugar into the other, there will be a big difference in concentration.

The salt will dissociate into its ions, but sugar (a molecule) will not dissociate. If the salt were NaCl, the concentration would be double that of the sugar. If the salt were $MgCl_2$, the concentration would be triple (there are three ions).

Solubility Rules

Not all ionic compounds are soluble. Some ionic compounds have so much attractive force between their anions and cations that they will not dissociate. These substances are insoluble and will not dissolve. Instead, they clump together as a solid in the bottom of solution. Many ionic compounds, however, will dissociate in water and dissolve.

In these cases, the attractive force between ion and water is greater than that between cation and anion. There are several rules to help you determine which compounds will dissolve and which will not.

Solubility Rules

1. All compounds with Group 1 ions or ammonium ions are soluble.
2. Nitrates, acetates, and chlorates are soluble.
3. Compounds containing a halogen are soluble, except those with fluorine, silver, or mercury.

If they

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have lead, they are soluble only in hot water. 4. Sulfates are soluble, except when combined with silver, lead, calcium, barium, or strontium. . Carbonates, sulfides, oxides, silicates, and phosphates are insoluble, except for rule #1. 6. Hydroxides are insoluble except when combined with calcium, barium, strontium, or rule #1. Sometimes, when two different ionic compounds are dissolved, they react, forming a precipitate that is insoluble. Predicting these reactions requires knowledge of the activity series and solubility rules. These reactions can be written with all ions, or without the spectator ions (the ion that don't react, present on both sides of the reaction), a format known as the net ionic equation.

For example, silver nitrate is soluble, but silver chloride is not soluble (see the above rules). Mixing silver nitrate into sodium chloride would cause a cloudy white precipitate to form. This happens because of a double replacement reaction. Electrolytes When solutes dissociate (or if a molecule ionizes), the solution can conduct electricity. Compounds that readily form ions, thus being good conductors, are known as strong electrolytes. If only a small amount of ions are formed, electricity is poorly conducted, meaning the compound is a weak electrolyte.

A strong electrolyte will dissolve completely. All ions dissociate. A weak electrolyte, on the other hand, will partially dissociate, but some ions will remain bonded together. Colligative Properties Some properties are the same for all solute particles regardless of what kind. These are known as the colligative properties. These properties apply to ideal solutions, so in reality, the properties may not be exactly as calculated. In an ideal solution, there

are no forces acting between the solute particles, which is generally not the case. Vapor Pressure

All liquids have a tendency for their surface molecules to escape and evaporate, even if the liquid is not at its boiling point. This is because the average energy of the molecules is too small for evaporation, but some molecules could gain above average energy and escape. Vapor pressure is the measure of the pressure of the evaporated vapor, and it depends on the temperature of the solution and the quantities of solute. More solute will decrease vapor pressure. The vapor pressure is given by Raoult's Law, where X is the mole fraction of the solvent.

Notice that the vapor pressure equals that of the pure solvent when there is no solute ($X = 1$). If $X = 0$, there would be no vapor pressure at all. This could only happen if there were no solvent, only solute. A solid solute has no vapor pressure. $P_{\text{solution}} = P_{\text{pure solvent}} (X_{\text{solvent}})$ If two volatile substances (both have vapor pressures) are in solution, Raoult's Law is still used. In this case, Raoult's Law is essentially a linear combination of the vapor pressures of the substances. Two liquids in solution both have vapor pressures, so this equation must be used. $P_{\text{solution}} = P_1 (X_1) + P_2 (X_2)$

The second equation shows the relationship between the solvents. If two liquids were mixed exactly half-and-half, the vapor pressure of the resulting solution would be exactly halfway between the vapor pressures of the two solvents. Another relation in Henry's Law, which shows the relationship between gas and pressure. It is given by $C_g = k P_g$, where C is concentration and P is pressure. As the pressure goes up, the concentration of gas in solution must also increase. This is why soda cans release gas when

they are opened - The decrease in pressure results in a decrease in concentration of CO₂ in the soda.

Boiling Point Elevation A liquid reaches its boiling point when its vapor pressure is equal to the atmosphere around it. Because the presence of solute lowers the vapor pressure, the boiling point is raised. The boiling point increase is given by: $\Delta T_{\text{solution}} = K_b \times m_{\text{solution}}$ The reduced vapor pressure increases the boiling point of the liquid only if the solute itself is non-volatile, meaning it doesn't have a tendency to evaporate. For every mole of non-volatile solute per kilogram of solvent, the boiling point increases by a constant amount, known as the molal boiling-point constant (K_b).

Because this is a colligative property, K_b is not affected by the kind of solute.

Freezing Point Depression A liquid reaches its freezing temperature when its vapor pressure is equal to that of its solid form. Because the presence of the solute lowers the vapor pressure, the freezing point is lowered. The freezing point depression is given by: $\Delta T_{\text{solution}} = K_f \times m_{\text{solute}}$ Again, this equation works only for non-volatile solutes. The temperature of the freezing point decreases by a constant amount for every one mole of solute added per kilogram solvent. This constant (K_f) is known as the molar freezing-point constant.

Osmosis If you studied biology, you would know that osmosis is the movement of water through a membrane. If two solutions of different molarity are placed on opposite sides of a semipermeable membrane, then water will travel through the membrane to the side with higher molarity. This happens because the water molecules are "attached" to the solvent

molecules, so they cannot travel through the membrane. As a result, the water on the side with lower molarity can more easily travel through the membrane than the water on the other side. The pressure of this osmosis is given in the equation. $\pi = MRT$

Where π is the pressure, M is molarity, R is the gas constant, and T is temperature in Kelvin. Electrolytes and Colligative Properties When one mole of table salt is added to water, the colligative effects are double those that would have occurred if sugar were added instead. This is because the salt dissociates, forming twice as many particles as sugar would. This dissociation, called the Van't Hoff Factor describes how many particles that are dissociated into the solution and must be multiplied into the Boiling Point Elevation or Vapor Pressure Lowering equations. Different ways of expressing the concentration of solutions

Mass Percentage The mass percentage of a component in a given solution is the mass of the component per 100g of the solution. For e. g. , if W_A is the mass of the component A, W_B is the mass of the component B in a solution. Then, Example: A 10% solution of sodium chloride in water (by mass) means that 10g of sodium chloride are present in 100g of the solution. Volume percentage This unit is used in case of a liquid dissolved in another liquid. The volume percentage is defined as the volume of the solute per 100 parts by volume of solution. For e. g. , If V_A is the volume of component A present is V_{sol} volume of the solution.

Then, For e. g. , a 10% solution of ethanol C_2H_5OH , in water (by volume) means that 10cm³ of ethanol is present in 100cm³ of the solution. Strength of a solution is defined as the amount of the solute in gms, present in one

litre of the solution. It is expressed as g L^{-1} . Mathematically, Molarity Molarity of a solution is defined as the number of moles of solute dissolved per litre of solution. Mathematically, For e. g. , If 'a' is the weight of the solute (in gms) present in VCC volume of the solution. Then, Molarity is expressed by the symbol M. It can also be expressed as, Normality

Normality of a solution is defined as the number of gram equivalents (gm. e) of a solute dissolved per litre of the given solution. Mathematically it is, For e. g. , If a is the weight of the solute (in gms) present in VCC volume of the solution. Then, Normality is expressed by the symbol N. It can also be expressed as, Relationship between molarity and normality The molarity and normality of a solution is related to each other as follows: Molality Molality of a solution is defined as the number of moles of solute dissolved in 1000g of a solvent. Mathematically, it is expressed as Molality is expressed by the symbol m.

Molality does not change with temperature. Formality In case of ionic compounds like KCl, CaCO_3 etc. Formality is used in place of molarity. It is the number of gram formula masses of solute dissolved per liter of the solution. It is denoted by the symbol F. Mathematically it is given as, Mole Fraction It is the ratio of number of moles of one component (solute or solvent) to the total number of moles of all the components (solute and solvent) present in the solution. It is denoted by the symbol X. Let us suppose that a solution contains two components A and B and suppose that n_A moles of A and n_B moles of B are present in the solution then,

Adding eq (i) and (ii) we get $X_A + X_B = 1$ Parts per million (ppm) When a solute is present in very small amounts, its concentration is expressed in

parts per million. It is defined as the amount of the solute present in one million parts of the solution. It may be noted that the concentration units like molarity, mole fraction etc. are preferred as they involve the weight of the solute and solvent, which is independent of the temperature. But units like, molarity, normality etc. , involve volume of the solution, hence changes with temperature.