

Dr. Adcock Has a Web Page

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Downloaded from Any Internet Enabled
Computer!**

e-mail: Jadcock@utk.edu

U05 CPS Class Key: W10124Q488

**PHYSICAL PROPERTIES
OF SOLUTIONS**

Chapter 12

Hill, Petrucci, McCreary & Perry 4th Ed.

Solutions = *Homogeneous Mixtures*

May be gases, liquids or solids

Solvent = *Dissolving Medium*

Usually largest quantity component!

Solute = *Substance Dissolved*

Usually smallest quantity

Definitions

Solubility = *Amount that Dissolves!*
Mass or Moles per Unit of Volume (or
mass of solvent).

**Saturated Solution = A Solution which is
in Equilibrium with undissolved solute.**

**Unsaturated Solution = A Solution in
which more solute could dissolve before
equilibrium is reached.**

Supersaturated Solution

**A Solution in which the solute is
dissolved in greater amount than the
*Equilibrium Amount!***

***A metastable state* usually formed by
cooling a saturated solution in the
absence of crystallization. An *entropy
effect* as the disordered solution forms
a highly ordered crystal.**

Solubility Rules

Like Dissolves Like

***Polar Solvents* dissolve *Polar* and *Ionic
Solutes!***

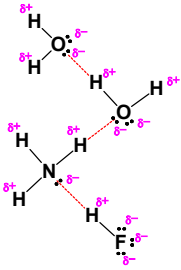
Water* ⇔ *alcohol* or *salt

***Nonpolar Solvents* dissolve *Nonpolar
Solutes!***

Kerosene* ⇔ *Oil* or *tar

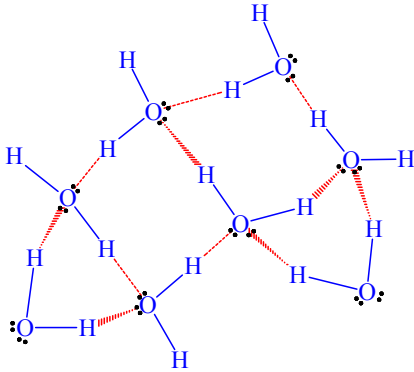
Hydrogen Bonding

A special kind of dipolar attraction which causes molecules so capable of strongly attracting themselves and each other. This will cause strong solubility attractions and exclusions to occur.

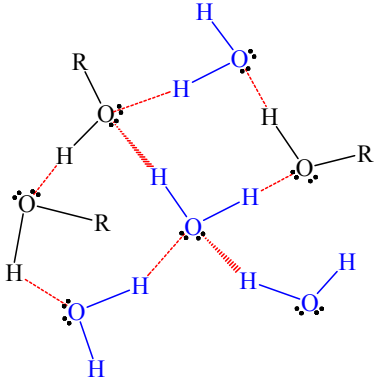


Three small electronegative atoms N O F draw electrons away from the hydrogens bonded to them making the hydrogen atoms positively charged, and also concentrate negative charge on their small lone pairs of electrons. A "Hydrogen bond" will form between the hydrogen atom and a lone-pair on a neighbor.

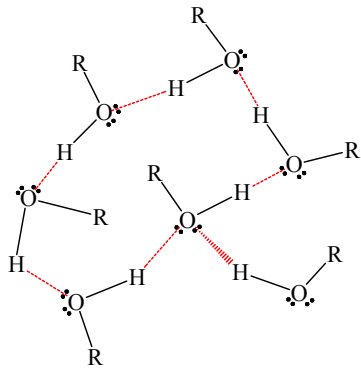
Water forms Hydrogen Bonds



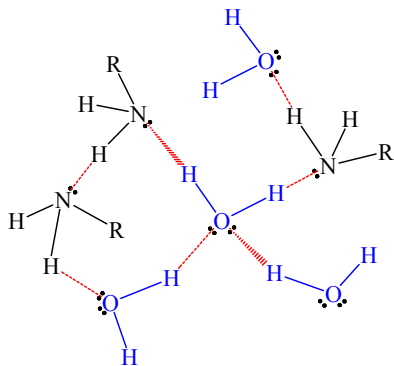
Hydrogen Bonds – Alcohols in Water



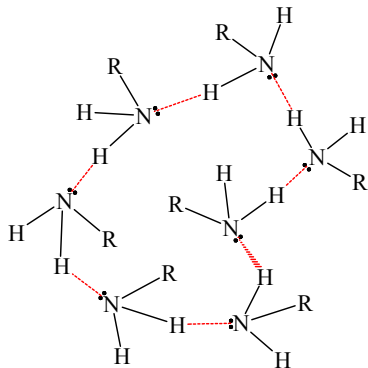
Hydrogen Bonds – Pure Alcohols



Hydrogen Bonds – Amines in Water



Hydrogen Bonds – Pure Amines



Temperature Effects

Increasing Temperature:

- **Solid or Liquid solutes:** *usually increase* in solubility, converse is also true.
- **Gaseous solutes:** *decrease* in solubility, converse is also true.

Pressure Effects

- **Pressure** has little effect on the solubility of liquid or solid solutes.
- **Gas Solubility** is proportional to the *Partial Pressure* of the gas in equilibrium with the solution. This is called Henry's Law.

Gravimetric Concentration Units

Molality, $m = (\text{moles solute})/(\text{kilograms solvent})$

Mole Fraction, $X_A = (\text{moles A})/(\text{total moles})$

Mole Percent A = $X_A \times 100$

Mass Percent = $(\text{mass solute})(100)/(\text{total mass})$

Molality, Mole Fraction, Mole Percent and Mass Percent are interconvertible units if the components of the solution and their molecular (formula) weights are given.

Conversions between Various Gravimetric Concentration Units

To convert between concentration units a “convenient quantity” must be assumed for the *known concentration value*.

Examples, Converting from:

Percent by weight: Assume exactly 100 grams soln.

Mole Fraction: Assume exactly one mole total.

Molality: Assume exactly 1 kilogram of solvent.

Gravimetric Concentration Interconversions

Calculate the mole fraction and molality of a 40.0% NaOH (40.0 g/mol) solution in water (18.0 g/mol).

Colligative Properties

Colligative properties result from interactions between groups of molecules and are mostly properties of the *Solvents!*

Vapor Pressure Lowering

Boiling Point Elevation

Freezing Point Depression

Changes in Osmotic Pressure

Vapor Pressure Lowering

$$\text{Raoult's Law: } P_A = P_A^\circ X_A$$

The partial pressure of a solvent, P_A , over a solution equals the vapor pressure of the pure solvent, P_A° , times the mole fraction of the solvent in the solution.

A nonvolatile solute lowers the vapor pressure to the degree it replaces solvent molecules in the solution.

Boiling Point Elevation

Boiling Occurs when: $P_{\text{vapor}} = P_{\text{atmospheric}}$

Normal Boiling Point: $P_{\text{vapor}} = 1 \text{ atm.}$

$$\Delta T_b = K_b \times m$$

$$T_{b, \text{soln}} = T_{b, \text{solvent}} + \Delta T_b$$

Molal bp elevation constant is a characteristic of the *solvent!* A constant of proportionality between molality and the *change* in the boiling point.

Explanation

A nonvolatile solute lowers the vapor pressure of a solvent to the degree it replaces solvent molecules throughout but most importantly at the surface of the liquid. In order for boiling to resume, a higher temperature is needed to compensate for the loss of vapor pressure, boiling resumes when the vapor pressure again reaches atmospheric pressure.

Freezing Point Lowering

$$\Delta T_f = - K_f \times m$$

The molal freezing point lowering constant is a characteristic of the *solvent!*

$$T_{\text{fp, soln}} = T_{\text{fp, solvent}} + \Delta T_f$$

Note: Some texts do not put the minus sign in the first equation, instead they subtract the ΔT_f in the second equation

Explanation

A solution has greater disorder (more entropy) than the pure solvent. To cause crystallization of pure solvent from the solution, more energy must be removed (cooling) relative to freezing of the pure solvent.

Example: Fresh water freezes at 0° C, but seawater produces pure ice at -5° C

Calculation of Molecular Weights

$$\Delta T_f = - K_f \times m$$

$$\Delta T_f = - K_f \times (\text{moles solute}) / (\text{kilograms solvent})$$

$$\Delta T_f = - K_f \times (\text{grams solute}) / (\text{MW solute})(\text{kg solvent})$$

Rearranging:

$$\text{MW}_{\text{solute}} = - K_f \times (\text{grams solute}) / (\Delta T_f) (\text{kg solvent})$$

Example Problem – A solution contains 0.0653 g of a compound in 9.75 g ethanol. The molality of the solution is 0.0368 m. Calculate the molecular weight of the compound.

Note: you are given the molality, which you would normally have to derive from ΔT_f or ΔT_b i.e. $m = \Delta T_f / -K_f$

So: $m = (\text{grams}_{\text{solute}}) / (\text{MW}_{\text{solute}})(\text{kg}_{\text{solvent}})$

Example Problem: Urea, $(\text{NH}_2)_2\text{CO}$, is dissolved in 100.0 g of water. The solution freezes at -0.085°C . How many grams of urea were dissolved to make this solution?

First: What are you trying to calculate?

Second: What are the general equations governing this problem?

Third: Do you need data not supplied in the problem, but available in the text.

To work this problem you will need to look up freezing point data for water from Table 12.2, p 507.

You will need to calculate ΔT_f for the solution, you will need to calculate the mole weight for urea from the formula.

From the ΔT_f and K_f you can calculate the molality, or simply plug into a rearranged combined formula to get grams urea.

Example problem – A solution was prepared by dissolving 0.915 g sulfur, S₈, in 100.0 g acetic acid. Calculate the freezing point and boiling point of this solution.

You will need to get the mp, bp, K_f, K_b for acetic acid from Table 12.2, p 507.

You will need to calculate the number of moles of sulfur as S₈ molecules, and the molality of the solution.
