65 Example: If a solution is 0.688 m citric acid, what is the molar concentration (M) of the solution? The density of the solution is 1.049 g/mL

$H_3C_6H_5O_7$: 192.125 g/mol "CA"

$\frac{0.688 \text{ mol CA}}{\text{kg solvent}} \longrightarrow \frac{? \text{ mol CA}}{\text{L solution}}$

molality (definition) molarity (definition)

1 - ASSUME A BASIS of 1 kg of solvent. Each kg of solvent contains 0.688 moles of CA.
2 - Find volume of SOLUTION. How? We know the density of the solution, so if we can figure out the mass of the solution, we'll know the volume too. To get the mass of solution, find the mass of CA (convert from moles to mass) and add it to the 1000g (kilogram) of solvent.

$$0.688 \text{ mol CA} \times \frac{192.125 \text{ g CA}}{\text{mol CA}} = 132.182 \text{ g CA}$$

mass solution = 1000 g solvent + 132.182 g CA = 1132.182 g

Find volume using density ...

$$1132.182 \text{ g} \times \frac{\text{mL}}{1.049 \text{ g}} = 1079.296473 \text{ mL} = 1.079296473 \text{ L}$$

$$M = \frac{\text{mol CA}}{\text{L solution}} = \frac{0.688 \text{ mol CA}}{1.079296473 \text{ L}} = \boxed{0.637 \ \underline{M} \text{ CA}}$$

66 An aqueous solution is 8.50% ammonium chloride by mass. The density of the solution is 1.024 g/mL
Find: molality, mole fraction, molarity.

NH_4Cl : 53.491 g/mol H_2O: 18.016 g/mol

To find molality:

$$\frac{8.50\text{ g } NH_4Cl}{100\text{ g solution}} \longrightarrow \frac{\text{mol } NH_4Cl}{\text{kg } H_2O}$$

mass percent   molality

Assuming 100g solution, find moles ammonium chloride:

$$8.50\text{ g } NH_4Cl \times \frac{\text{mol } NH_4Cl}{53.491\text{ g } NH_4Cl} = 0.158905236\text{4 mol } NH_4Cl$$

Find ,mass water:

$$100\text{ g solution} - 8.50\text{ g } NH_4Cl = 91.50\text{ g } H_2O = 0.09150\text{ kg } H_2O$$

Molality is:

$$\frac{0.1589052364\text{ mol } NH_4Cl}{0.09150\text{ kg } H_2O} = \boxed{1.74\text{ m } NH_4Cl}$$

To find mole fraction:

$$\frac{8.50\text{ g } NH_4Cl}{100\text{ g solution}} \longrightarrow \frac{\text{mol } NH_4Cl}{\text{mol } NH_4Cl + \text{mol } H_2O}$$

mass percent   mole fraction

An aqueous solution is 8.50% ammonium chloride by mass. The density of the solution is 1.024 g/mL
Find: molality, mole fraction, molarity.

NH_4Cl : 53.491 g/mol H_2O: 18.016 g/mol

To find mole fraction, keep the 100g solution basis, so we already know both the mass and moles (from the last calculation) of ammonium chloride. All that's left is to convert the mass water to moles!

Find moles water: $91.50\ g\ H_2O \times \frac{mol\ H_2O}{18.016\ g\ H_2O} = 5.078818828\ mol\ H_2O$

$$X_{NH_4Cl} = \frac{0.158905236 4\ mol\ NH_4Cl}{0.1589052364\ mol\ NH_4Cl + 5.078818828\ mol\ H_2O} = \boxed{0.0303}$$

Find molarity: $\frac{8.50\ g\ NH_4Cl}{100\ g\ solution} \rightarrow \frac{mol\ NH_4Cl}{L\ solution}$

mass percent   molarity

Keeping the 100g solution basis means we know the moles ammonium chloride. We just need to convert 100g soluiton to volume wiht density.

$100\ g\ solution \times \frac{mL}{1.024\ g} = 97.65625\ mL$ (0.09765625 L)

So molarity is...

$$\frac{0.1589052364\ mol\ NH_4Cl}{0.09765625\ L} = \boxed{1.63\ M\ NH_4Cl}$$

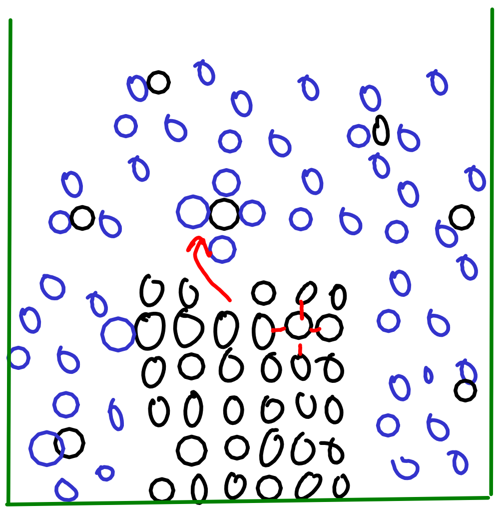
HOW THINGS DISSOLVE

- Let's look at how things dissolve into water, since aqueous solutions are quite common.

sucrose (table sugar)

$$C_{12}H_{22}O_{11}(s) \overset{H_2O}{\rightleftharpoons} C_{12}H_{22}O_{11}(aq)$$

... what happens?



- Water molecules pull the sugar molecules out of the sugar crystal and into solution.

- Attractions between sugar molecules and water allow this to happen.

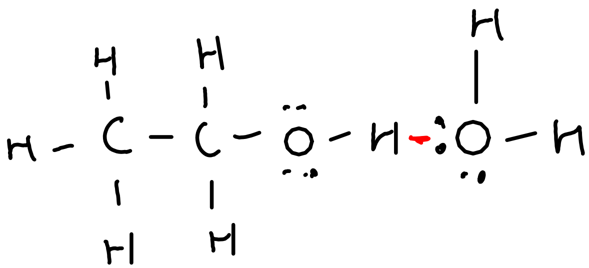
- The solubility of the sugar depends on how well water and sugar interact (HYDRATION) versus how well the sugar molecules are held in the crystal (LATTICE ENERGY)

- "like dissolves like": Substances held together by similar (or at least compatible) kinds of attractive forces can dissolve in each other. Substances that are held together by very different kinds of attractive forces will not dissolve in one another!

Consider WATER:

HYDROGEN BONDS ↓

Water mixes well with other substances that can hydrogen bond, like ETHANOL!



POLAR ↓

Water can dissolve polar substances! (SUCROSE is polar!)

↓

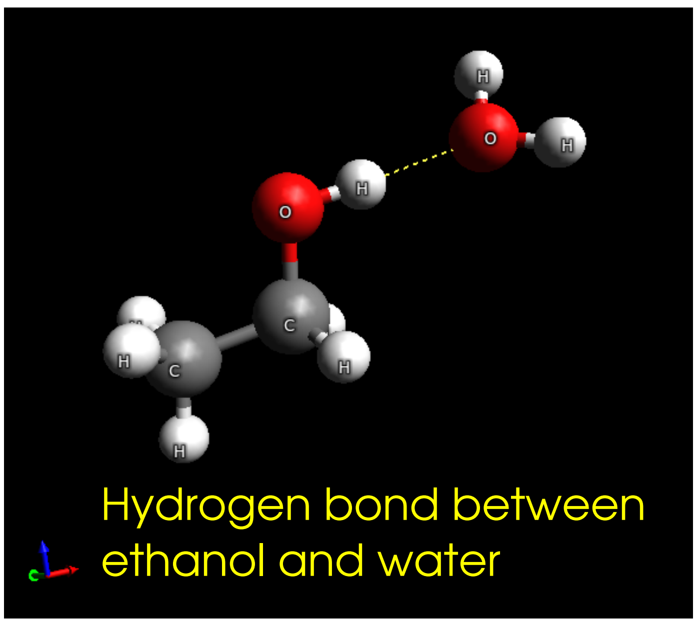
Since IONIC BONDS are also interactions between opposite charges (You can think of an ionic bond here as an extreme case of dipole-dipole interaction), many IONIC SUBSTANCES will also dissolve in water!

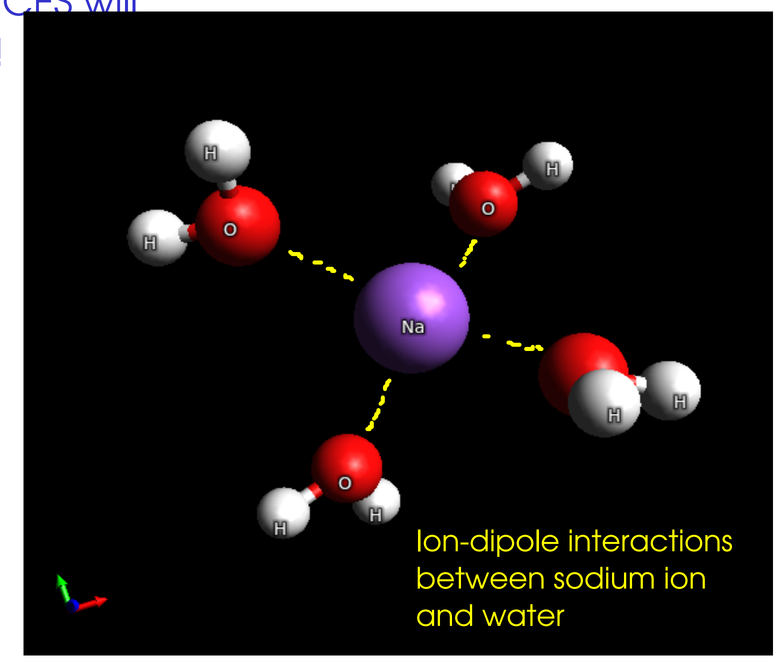
$Na^+ - Cl^-$

SMALL (little London force) ↓

large and/or nonpolar solutes do not dissolve well in water!

(example: OILS and WAXES)





MOLECULAR AND IONIC SOLUTIONS

- MOLECULAR solutions:

Contain MOLECULES dissolved in one another.

(1) - Any mixture of GASES

- all gases mix with one another, since gas molecules (effectively) do not interact with one another.

(2) - Liquids

- Liquids dissolve well in one another only if they are held together by similar kinds of forces

(3) - Solids and liquids

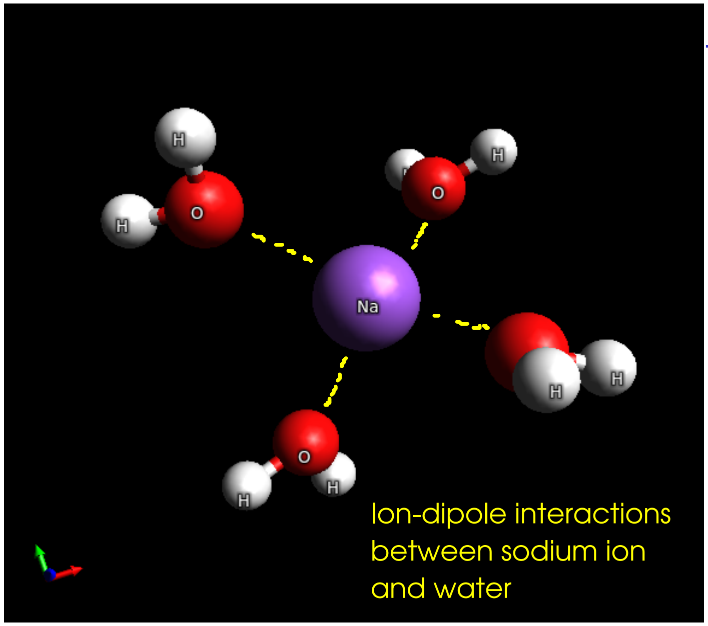
- MOLECULAR SOLIDS will dissolve well in liquids if they are held together by similar forces.

- IONIC SOLIDS will sometimes dissolve in POLAR liquids, but not in nonpolar liquids

- COVALENT NETWORK solids don't generally dissolve well in other substances

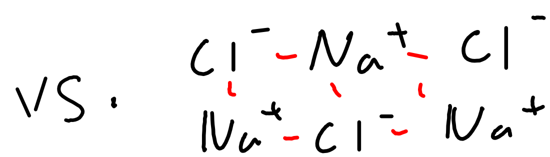
IONIC solutions

- form when ions from IONIC SUBSTANCES interact with POLAR solvents - often WATER.



Ion-dipole interactions between sodium ion and water

The charged ends of the water molecule HYDRATE the ions.



- The solubility of an ionic compound depends on whether HYDRATION (attraction of water molecules for an ion) is greater than LATTICE ENERGY - the attrraction of ions in a crystal lattice for one another..

- SMALLER IONS are usually easier to enclose in water than larger ones, and ions with larger charges are attracted to water molecules.

- But solubility is also determined by LATTICE ENERGY - which holds the solid ionic compound together. Ions with high charges tend to be strongly attracted to other ions in a crystal, meaning lattice energy is high. Smaller ions also tend to have higher lattice energies. Lattice energy and hydroation are competing trends!

COLLIGATIVE PROPERTIES

- properties unique to solutions.

- depend only on the CONCENTRATION of a solution and not the IDENTITY of the solute**

**ionic solutes: Remember that they dissociate into MULTIPLE IONS!

1. Freezing point depression
   - The freezing temperature of a SOLUTION gets lower as the CONCENTRATION of a solution increases.

2. Vapor pressure lowering
   - The vapor pressure of a solution (pressure of sovent vapor over a liquid surface) goes DOWN as solution concentration goes UP

3. Boiling point elevation
   - The boiling temperature of a solution increases as the concentration of the solution increases.

4. Osmotic pressure
   - The pressure required to PREVENT the process of osmosis

FREEZING POINT DEPRESSION

$$\Delta T_f = K_f \times c_m$$

- c_m: concentration of solute (molality)
- K_f: Freezing point depression constant (for SOLVENT)
- ΔT_f: Freezing point depression: The amount the freezing temperature is LOWERED by the solute.

- Applications: In chemistry, this effect is often used to determine the molecular weight of an unknown molecule.

74 A solution of 2.500g of unknown dissolved in 100.0 g of benzene has a freezing point of 4.880 C. What is the molecular weight of the unknown?

$K_{f,\,benzene} = 5.065\ ^{\circ}C/m$, $T_{f,\,benzene} = 5.455^{\circ}C$ (see p500 9th p509, 10th)

$\Delta T_f = K_f \times C_m$

$K_f = 5.065\ ^{\circ}C/m$

$\Delta T_f = 5.455^{\circ}C - 4.880^{\circ}C = 0.575^{\circ}C$

$C_m = \dfrac{\text{mol unknown}}{\text{kg benzene}}$

To find molecular weight, we need to know moles of unknown. We can get that from Cm

First, calculate Cm ... the molality:

$0.575^{\circ}C = (5.065^{\circ}C/m) \times C_m$

$C_m = 0.113524186\ m \text{ unknown}$

To find moles unknown, moltiply the mass of solvent by the molality:

$0.1000 \text{ kg benzene} \times \dfrac{0.113524186 \text{ mol unk}}{\text{kg benzene}} = 0.0113524186 \text{ mol unk}$

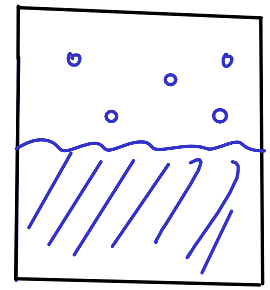
Since MOLECULAR WEIGHT is mass per mole ...

$MW = \dfrac{\text{mass unk}}{\text{mol unk}} = \dfrac{2.500 \text{ g unk}}{0.0113524186 \text{ mol unk}} = \boxed{220.\ g/mol}$

VAPOR PRESSURE LOWERING

- Described by RAOULT'S LAW

$$P_A = P_A^* \times X_A$$



P_A = partial pressure of the VAPOR of solvent molecules.

- X_A: mole fraction of component A
- P_A^*: vapor pressure of pure component A (depends on temperature)
- P_A: partial pressure of component A in a solution

... but component "A" above is actually the SOLVENT. If we want to describe this as a colligative property, we want to express Raolt's law in terms of the SOLUTE! Assuming a two-component mixture, we get...

$$\Delta P = P_A^* \times X_B$$

- X_B: mole fraction of component B (the SOLUTE in a two-component mixture)
- ΔP: Vapor pressure lowering. This is the DECREASE in the vapor pressure of the solvent due to the presence of solute.

BOILING POINT ELEVATION

- Since the vapor pressure is lowered by the presence of a solute, AND since boiling occurs when the vapor pressure of a liquid equals the external pressure - solutes also cause BOILING POINT ELEVATION.

- The equation for boiling point elevation looks almost exactly like the equation for the freezing point depression, and is used in almost the same way.

$$\Delta T_b = K_b \times c_m$$

c_m: concentration of solute (molality)

K_b: Boiling point elevation constant (for SOLVENT) (p500, 9th; p509, 10th)

ΔT_b: Boiling point elevation: The amount the boiling temperature is RAISED by the solute.

What is the boiling point of a solution that contains 2.817 g of molecular sulfur S_8) dissolved in 100.0 grams of acetic acid?

$T_b = 118.5°C$ $K_b = 3.08 °C/m$ (see p500 for data) p509, 10th

$\Delta T_b = K_b \times C_m$ (3.08 °C/m)

$C_m = \frac{mol\ S_8}{kg\ HC_2H_3O_2}$ — 0.1000 kg $HC_2H_3O_2$

First, find the moles of sulfur. Then find Cm and calculate boiling point elevation.

S_8: 8 x 32.07 = 256.56 g S_8 = mol S_8

$2.817\ g\ S_8 \times \frac{mol\ S_8}{256.56\ g\ S_8} = 0.01097988777\ mol\ S_8$

Find Cm:

$C_m = \frac{0.01097988777\ mol\ S_8}{0.1000\ kg\ HC_2H_3O_2} = 0.1097988775\ m$

Find boiling point ELEVATION:

$\Delta T_b = (3.08\ °C/m)(0.1097988775\ m) = 0.338\ °C$

To get the new boiling temperature, add the elevation to the boiling temperature of the pure acetic acid:

$T_{b,solution} = 118.5\ °C + 0.338\ °C = \boxed{118.8\ °C}$