

SOLUTIONS

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Mixtures

-a random assortment of matter that you can physically separate

- magnetism
- filtration
- floatation
- gravitation

-**homogeneous** and **heterogenous** are mixture types

-**homogeneous**: any one part represents the whole

-**heterogenous**: no one part represents the whole

-must use context to distinguish

-**mixture** vs. **compound**: compounds are bonded, mixtures are physical

-all phases of matter can form mixtures with any other phase of matter

-**Solution** is when you can't physically separate. Only chemically can they be separated

-Solution is a type of mixture

-homogenous = irreversible

-homogenous is preferred over heterogenous (max entropy)

It's all about size!

-If compounds can readily mix in atomic sizes, mixtures become inseparable! This is a true solution.

Size	Type	Light
<1 nm	Solution (homo)	no effect
1nm-1000nm	Colloid (hetero)	tyndal effect
>1000 nm	suspension (hetero)	blocks passage of light

-**colloids** diffract light

-**Brownian motion**: Einstein explained particles in random motion all the time

Electrolytes

-**Solvent + Solute = Solution**

-**Solvent**: that which dissolves the solute

-**Solute**: the thing being dissolved

-Solute is typically named

-**Electrolytic** solution improves the electrical conductivity of the solution

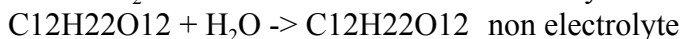
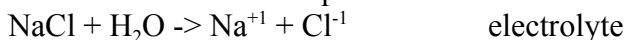
-**Non - electrolytic** solutions do not change or inhibit the conductivity of the solution

-universal solvent = H₂O

Dissociation of Ions

-chemicals with ionic bonds (or very polar covalent bonds) tend to dissociate in polar solvents

-**dissociation** means to "break apart"



-more dissociation (ionic bonds) mean more electrolytic

-**electrolitics** is a measure of dissociation

- strong electrolyte = about 100% dissociation
- weak electrolyte = <100% -> 5-10%
- strong acids are strong because they are strong electrolytes
- It's all about bond types
- solvents and solutes fall into two categories:

- | | |
|--|---|
| <p>Polar</p> <ul style="list-style-type: none"> -opposite charges across molecules -lone pairs very negative -electronegativity counts | <p>Non-Polar</p> <ul style="list-style-type: none"> -no charge difference across the molecule |
|--|---|

It's all about crossing the molecule + to -.

- like dissolves like (polar dissolves polar, water dissolves water like substances)
- CCl₄ is non-polar

Non-Additivity of Volumes

- volume changes cannot be predicted for solid and liquid phases
- solutions must be made up carefully (solvent to solute until final volume is reached)

Factors Affecting "Solution Making" Rate

- Temperature: $T \propto 1/t$ -not gases, chalk (Calcium carbonate), or Ce₂SO₄ (cesium sulphate)
- Surface Area: $S.A._{solute} \propto 1/t$
- Agitation: $Agitation_{solution} \propto 1/t$
- take solute and add solvent**
- immiscible solution:** heterogenous mixture of two liquids, "don't go together"
- miscible solution:** homogenous mixture of two liquids
- most solutes are more soluble when you increase temperature.
 - gases are less soluble when you increase temperature
 - chalk
 - cesium sulphate

Solubility

- a measure of solvent capacity
- expressed in grams solute per 100g solvent
- the capacity of a solvent is unique to a solute, and cannot be predicted
 - Unsaturated:** -when more solute can fit into a solvent
 - Saturated:** -when no more solute can fit into a solvent
 - equilibrium condition, rate of dissociation = rate of recrystallization
 - Supersaturated:** -when more solute is in a solvent at a given temp.
 - when change temp from hot to cold quickly, crystals can't form fast enough
- solute/solvent capacity can not be predicted

Henry's Law

- the amount of gas dissolved in a solvent is directly proportional to the partial pressure of the gas

at the solvent surface.

$$\text{Mol}_{\text{gas}} \propto P_{\text{p-gas}}$$

-**Recrystallization** is typically exothermic

-**Dissociation** is usually endothermic

-solid \rightarrow solution is always endothermic

Solubility

-soluble: $\geq 1\text{g}/100\text{g}$ solvent

-sparingly soluble: $1\text{g} > 0.1\text{g}/100\text{g}$ solvent

-insoluble $\leq 0.1\text{g}/100\text{g}$ solvent

Concentration

-it's a measure of amount of solute in a given solution

-it has two measures in general chemistry

Molarity (M)

$\frac{\text{mol (solute)}}{\text{vol (solution) L}}$

Molality (m)

$\frac{\text{mol (solute)}}{\text{mass (solvent) kg}}$

-molarity changes with temperature changes, molality does not

-concentration is inversely proportional to volume

-concentration is directly proportional to amount solute

- $M/p = m$, or $M/m = p$ (p, density)

[] = molarity

{ } = molality

Solubility Rules (reasons for double replacement reactions) (p 427)

1. Most sodium (Na), potassium (K), and ammonium (NH₄) compounds are **soluble** in water.

2. Most nitrates (NO₃), acetates (CH₃COO), and chlorates (ClO₃) are **soluble**.

3. Most chlorides (___Cl) are **soluble**, except those of silver (Ag), mercury I (Hg), and lead (Pb).
PbCl₂ is soluble in warm water.

4. Most sulfates (SO₄) are **soluble**, except those of barium (Ba), strontium (Sr), and lead (Pb).

5. Most carbonates (___CO₃), phosphates (PO₄), and silicates (SiO₄) are **insoluble**, except those of sodium (Na), potassium (K), and ammonium (NH₄).

6. Most sulfides (___S) are **insoluble**, except those of calcium (Ca), strontium (Sr), sodium (Na), potassium (K), and ammonium (NH₄).

Double Replacement reactions are driven by solubility

-double replacement only takes place if one (or more) possible configurations of ions is insoluble in the solvent (or forms water) (precipitate or gas).

Reasons for Solubility

-bond types: ionic bonds-very soluble, covalent bonds-very insoluble

-enthalpy changes: $\Delta G = \Delta H - T \Delta S$, want ΔH (-)

$$\Delta H = \Delta H_{\text{dissociation}} + \Delta H_{\text{solvation}}$$

endo (+) exo (-)
 to break to make

Net Ionic Equations

1. Chemical:
2. Total Ionic:
3. Net Ionic:

Vapor Pressure Crap

- vapor pressure: pressure caused by the vapor of the liquid (more pressure = more pure)
- Vanerwall's forces: reasons that things liquify. When things cool down, they form bonds with each other (strong with H₂O) **Intramolecular** bond
- at liquid surface, less attraction than in the liquid, so at the surface, it's easier to get "kicked"
- when they get kicked out, vapor phase (absorbed energy, but NOT gas)
- evaporation is a cooling process (endothermic)
- condensation is a heating process (exothermic)
- If you add a solute, less vapor pressure (remember dissociation!!!)
- Vapor takes **mass** and **energy** when a molecule gets kicked.
- When you compress a vapor, you get a liquid (with a gas, you get hot gas)
- Example**: adding sulfuric acid to ice (class demonstration)

The sulfuric acid acts as a solute, decreasing the vapor pressure. Since vapor pressure is decreased, the freezing point is decreased. Because of this, the ice is not at the temperature to remain an ice (it's not cold enough), so some of it melts (and further reacts with sulfuric acid, making it hot), and the rest of it stays as ice by absorbing energy from the surroundings to keep in a solid phase (because freezing is endothermic). This is why the ice got colder, even though it was hot beneath.

Colligative Properties

- a colligative property is one that depends only on the concentration of the solute particles, not on their chemical nature (note: this is a lie, because if the solute dissociates, ie: ionic bonds, the property will be increased)
- molality is used as the concentration term (because it is very "state" stable)
- most of these properties rely on the vapor pressure effects
- Freezing Point Depression: solute doesn't matter (not true...)

$$\Delta T = k_f m \quad k_f \text{ is solvent specific (H}_2\text{O is } -1.86 \text{ k/m)}$$

- Boiling Point Elevation:

$$\Delta T = k_b m \quad k_b \text{ is solvent specific (H}_2\text{O is } 0.51 \text{ k/m)}$$

- with these, you can find the molar mass of the unknowns, just find the kg of solvent and solve.

OK, so it does matter what solute you use!

- it's an ion effect: every ion can cause the same colligative property (if it dissociates...)
- strong electrolytes only: multiply the molality by the number of product ions

$\Delta H_{\text{solvation}}$ is typically exothermic.

Equations

M=mol/vol (vol of entire solution)

m=mol/kg (kg of solvent)

$$M/m = p \text{ (p = density)}$$

$$\Delta H = \Delta H(\text{dissociation}) + \Delta H(\text{solvation})$$

$$\text{molarity}_1 \times \text{volume}_1 = \text{molarity}_2 \times \text{volume}_2 \text{ (Used for dilutions)}$$

$$\text{Beer's Law: } A = \epsilon l [A]$$

A = absorbance

ϵ = molar absorbtivity constant

l = path length of light

[A] = molarity of the substance