Solutions

Unit 6

Solution: a homogeneous mixture composed of two or more substances

Compounds

- chemical compounds are *substances* composed of two or more *elements* that are *chemically bonded* in a *definite proportion*
- a chemical compound is formed by a chemical change and new substances with new properties are formed
- a single chemical compound is *one* single substance

Mixtures

- two or more *substances* that can be separated by physical means
- the *proportions* in a mixture can be *varied*
- the substances are not *chemically bonded*
- two or more components in a mixture retain their original properties

Types of Mixtures			
	Heterogeneous		Homogeneous
	Simple Suspension	Colloid	Solution
Particle Size	Visible Greater than 1 µm	Microscopic 1 µm down to 1 nm	Molecular / Ionic 1 nm down to 62 pm
Settles Out	Yes	No	No
Key Feature	Easily Separated by Filtering	Tyndall Effect and Brownian Motion	Clear
Examples	river delta oil and vinegar	smoke, gelatin, mayonnaise, ink	salt water, sugar water, soda water, vinegar, tincture of iodine, alloys, marshmallows, rubbing alcohol, amalgums

Solution: a homogeneous mixture

- dissolved ≡ in solution
- (aq) ≡ in *water* solution

Solution concentrations can be:

- concentrated ≡ high solute proportion
- dilute ≡ low solute proportion

Parts of a Solution:

- solute ≡ minor part of a solution, the part that gets dissolved
- solvent ≡ major part of a solution, it does the dissolving
  water is the most common solvent – often called the universal solvent

Types of Solutions (based on solvent phase):

- gaseous solution**:** the solute must be a gas (or multiple gases)
- liquid solution: the solute may be solid, liquid, or gas
- solid solution: the solute may be a solid, liquid, or gas

Solubility of a solute:

- soluble**:** able to be dissolved
- insoluble: *no* appreciable dissolution observed

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Saturation

The term saturated defines the *solubility* of a solute at a given temperature

Degrees of saturation:

- saturated**:** no more solute can be dissolved in the solute under the given conditions (added solute will not dissolve)
- unsaturated**:** a solution with less solute than a saturated solution (added solute will dissolve)
- supersaturated**:** a metastable condition in which more solute is dissolved than is normally possible under the given conditions (added solute will cause crystallization)

Concentration

Solubility in 100 g of water (Solubility Curves): Table G

- show the most solute that can be dissolved in 100 g of water at a specific temperature
  - as temperature increases, solubility of *solids* increases
    - most compounds on Table G are salts and therefore solids at room temperature
  - as temperature increases, solubility of *gases* decreases
    - only three compounds on Table G are gases
- saturated solution is shown *on the line*
- unsaturated solution is shown *below the line*
- supersaturated solution is shown *above the line*

Example: How much ammonia, NH_4Cl, can be dissolved in water at 70° C?

At 70°C, the solubility curve crosses at 62 g NH_4Cl per 100 g H_2O

- saturated = 62 g NH_4Cl per 100 g H_2O
  any added NH_4Cl would just sink to the bottom – a dynamic equilibrium would occur: the rate at which solid solute dissolves would be equal to the rate at which dissolved solute recrystallizes
- unsaturated < 62 g NH_4Cl per 100 g H_2O
  any added NH_4Cl would dissolve
- supersaturated > 62 g NH_4Cl per 100 g H_2O
  any added NH_4Cl would cause any excess amount of NH_4Cl to precipitate

Example: In terms of saturation, describe a solution made by adding 27 grams of KNO_3 to 50. grams of water at 40° C

From Table G:

$$\frac{64\ g_{KNO_3}}{100\ g_{H_2O}} <=> \frac{27\ g_{KNO_3}}{50.\ g_{H_2O}} \times \frac{2}{2} = \frac{54\ g_{KNO_3}}{100\ g_{H_2O}} \qquad \therefore \text{ the solution is } \textit{unsaturated}$$

Example: In terms of saturation, describe a solution made by completely dissolving 2.5 grams of $KClO_3$ in 25 grams of water at 25° C

From Table G:

$$\frac{10.\ g_{KClO_3}}{100\ g_{H_2O}} <=> \frac{2.5\ g_{KClO_3}}{25\ g_{H_2O}} \times \frac{4}{4} = \frac{10.\ g_{KClO_3}}{100\ g_{H_2O}} \qquad \therefore \text{ the solution is } \textit{saturated}$$

Example: In terms of saturation, describe solution containing 23 grams of KNO_3 completely dissolved in 50. grams of $KClO_3$ in 25 grams of water at 25° C

From Table G:

$$\frac{41\ g_{KNO_3}}{100\ g_{H_2O}} <=> \frac{23\ g_{KNO_3}}{50.\ g_{H_2O}} \times \frac{4}{4} = \frac{46\ g_{KNO_3}}{100\ g_{H_2O}} \qquad \therefore \text{ the solution is } \textit{supersaturated}$$

Factors that Affect Rate of Solubility

- temperature: an increase in temperature increases the rate of dissolution
- surface area: an increase in the surface area increases the rate of dissolution
- agitation: an increase in agitation (stirring) increases the rate of dissolution
- pressure: increasing gas pressure increases the rate of dissolution (gases only)

Factors that Affect Solubility

- temperature
  as temperature increases, solubility of *solids* increases
  as temperature increases, solubility of *gases* decreases
  } see Table G section on previous page
- nature of the solvent
- nature of the solute
  } has to do with the polarity of the particles
  nonpolar molecules have even charge distribution (symmetrical)
  polar molecules have uneven charge distribution (asymmetrical)
  solubility rule: like dissolves like
  - ethyl alcohol (polar) dissolves well in water (polar)
  - salt (ionic) dissolves well in water (polar)
  - oil (nonpolar) dissolves well in hexane (nonpolar)
  - oil (nonpolar) does *not* dissolve in water (polar)
- as pressure increases, solubility of *gases* increases (Henry's Law)
  this effect is negligible for solid and liquid solutes

Identifying solutes that are soluble or insoluble *in water*: Table F

Na_2S	sodium sulfide	soluble
PbS	lead(II) sulfide	insoluble
$Al(OH)_3$	aluminum hydroxide	insoluble
Hg_2Cl_2	mercury(I) chloride	insoluble
$MgCrO_4$	magnesium chromate	soluble

Solubility in Percent Mass (or pph): Table G (curves) and Table T (equation)

Concentration in Table G is written as:

$$\frac{x\ \mathrm{g_{solute}}}{100\ \mathrm{g_{solvent}}}\ or\ \frac{\text{mass of solute}}{\text{mass of solvent}}$$

this is a mass/mass concentration

Concentration in Percent by Mass in Table T is written as:

$$\%\ \text{composition by mass} = \frac{\text{mass of part}}{\text{mass of whole}} \times 100 = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

a slightly different mass/mass concentration (measured in parts per hundred, pph)

Compare and contrast Table G and Table T $\%_{mass}$

Compare: the numerators of both equations are the same

Contrast: Table G denominator is solute only and Table T $\%_{mass}$ denominator includes both solute and solvent

What is the solubility of $NaNO_3$ at 20°C?

$$\frac{x\ \mathrm{g\ _{solute}}}{100\ \mathrm{g\ _{solvent}}}\ or\ \frac{\text{mass of solute}}{\text{mass of solvent}} = \frac{88\ \mathrm{g\ _{NaNO_3}}}{100\ \mathrm{g\ _{H_2O}}}$$

Determine the percent mass of $NaNO_3$ in a saturated solution at 20°C

$$\%_{\mathrm{mass}} = \frac{\text{mass of solute}}{\text{mass of solvent}} \times 100 = \frac{88\ \mathrm{g\ _{NaNO_3}}}{100\ \mathrm{g\ _{H_2O}} + 88\ \mathrm{g\ _{NaNO_3}}} \times 100 = 47\%\ _{\mathrm{NaNO_3}}$$

Describe how to make 750 g of a solution that is 15% NaOH by mass

$$\%_{\text{mass}} = \frac{\text{mass of solute}}{\text{mass of solvent}} \times 100 = 15\% = \frac{x\ \text{g}_{\text{ NaOH}}}{750\ \text{g}} \times 100\%$$

$$\frac{15\%}{100\%} = \frac{x\ \text{g}_{\text{ NaOH}}}{750\ \text{g}}$$

$$0.15 = \frac{x\ \text{g}_{\text{ NaOH}}}{750\ \text{g}}$$

x g $_{\text{NaOH}}$ = 0.15 × 750 g = 112.5 g or 110 g $_{\text{NaOH}}$ to 2 S.F.

Dissolve 110 grams of NaOH in 640 grams of water

In terms of saturation, describe a solution that is 27.5% NaCl by mass at 20°C

$$\%_{\text{mass}} = \frac{\text{mass of solute}}{\text{mass of solvent}} \times 100 = 27.5\% = \frac{x\ \text{g}_{\text{ NaCl}}}{100\ \text{g}_{\text{ H}_2\text{O}} + x\ \text{g}_{\text{ NaCl}}} \times 100\%$$

$$\frac{27.5\%}{100\%} = \frac{x\ \text{g}_{\text{ NaCl}}}{100\ \text{g}_{\text{ H}_2\text{O}} + x\ \text{g}_{\text{ NaCl}}}$$

$$0.275 = \frac{x\ \text{g}_{\text{ NaCl}}}{100\ \text{g}_{\text{ H}_2\text{O}} + x\ \text{g}_{\text{ NaCl}}}$$

x g $_{\text{NaCl}}$ = 0.275(100 g + x g $_{\text{NaCl}}$) = 27.5 g + 0.275 x g $_{\text{NaCl}}$

0.725 x g $_{\text{NaCl}}$ = 27.5 g

x g $_{\text{NaCl}}$ = 27.5 g ÷ 0.725 = 37.9 g

$\frac{37.9\ \text{g}_{\text{ NaCl}}}{100\ \text{g}_{\text{ H}_2\text{O}}}$ is a saturated solution at 20°C according to Table G

Solubility in Parts Per Million: Table T (equation)

$$\text{parts per million} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 1\,000\,000 \ \textit{or}\ \text{ppm} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 10^6$$

another mass/mass concentration

Scientists use parts per million (ppm) and parts per billion (ppb) to measure extremely small concentrations

Water is unfit to drink at 10 ppb Cd concentration

This would be 0.01 ppm or 0.000 001 % so it is much easier to use 10 ppb

Because percent literally means parts per hundred, comparing the equations for pph and ppm shows that ppm isn't really as strange as it first appears:

$$\text{pph} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100 \ \ \textit{and}\ \ \text{ppm} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 1\,000\,000$$

If the mercury concentration in a pond is 300 ppm, calculate the percent by mass concentration of the mercury

$$\text{ppm} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 10^6 = \frac{300\ \text{g}_{\text{Hg}}}{1\,000\,000\ \text{g}} \times 10^6$$

$$\text{pph} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100 = \frac{300\ \text{g}_{\text{Hg}}}{1\,000\,000\ \text{g}} \times 100 = 0.03\ \%$$

If 0.050 grams of NaCl is dissolved in 100. grams of water, express the concentration in parts per million

$$\text{ppm} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 10^6 = \frac{0.050\ \text{g}_{\text{Hg}}}{0.050\ \text{g} + 100.\ \text{g}} \times 10^6$$

Notice that by significant figure rules, 100. g + 0.050 g = 100. g

The added mass of the salt is *insignificant*

$$\text{ppm} = \frac{0.050\ \text{g}_{\text{Hg}}}{0.050\ \text{g} + 100.\ \text{g}} \times 10^6 = \frac{0.050\ \text{g}_{\text{Hg}}}{100.\ \text{g}} \times 10^6 = 500\ \text{ppm}$$

Assume we hadn't rounded off the denominator:

$$\text{ppm} = \frac{0.050\ \text{g}_{\text{Hg}}}{0.050\ \text{g} + 100.\ \text{g}} \times 10^6 = 499.75\ \text{ppm, which to 3 S.F.} = 500\ \text{ppm}$$

What mass of lead(II) nitrate, $Pb(NO_3)_2$, is required to produce 500.0 grams of a 155 ppm solution of lead(II) nitrate?

$$\text{ppm} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 10^6 = 155\ \text{ppm} = \frac{x\ \text{g}_{\text{Pb(NO}_3)_2}}{500.0\ \text{g}} \times 10^6$$

$$x\ \text{g}_{\text{Pb(NO}_3)_2} = 155 \times 500.0\ \text{g} \times 10^{-6} = 0.0775\ \text{g}$$

Not rounding the denominator gives 0.0775089 g or 0.0775 g to 3 S.F.

Molar Concentration: Table T

$$\text{molarity} = \frac{\text{moles of solute}}{\text{liter of solution}} \quad \textit{or} \quad \underline{\text{M}} = \frac{\text{mol}}{\text{L}}$$

which is a mass/volume concentration

Describe how to make 1 liter of a 1 molar NaOH solution

Add 1 mole of NaOH in a small amount of water to dissolve the NaOH, then dilute the solution with enough water to reach exactly 1 liter

Do NOT add 1 liter of water

For Molar solutions:

- volume *must* be in liters
- volumes in mL *must* be changed to liters

What is the molarity of a solution made by dissolving 0.575 moles of NaOH in enough water to form 350 mL of solution?

First, 350 mL = 0.350 L

$$\underline{\text{M}} = \frac{\text{mol}}{\text{L}} = \frac{0.575\ \text{mol}_{\text{NaOH}}}{0.350\ \text{L}_{\text{solution}}} = 1.64\ \underline{\text{M}}$$

How many grams of NaOH are required to make 150 mL of a 0.25 molar solution?

First, 150 mL = 0.150 L

$$0.15\ \text{L} \times \frac{0.25\ \text{mol}}{\text{L}} = 0.0375\ \text{mol}$$

$$0.0375\ \text{mol} \times \frac{40.0\ \text{g}_{\text{NaOH}}}{\text{mol}} = 1.5\ \text{g}_{\text{NaOH}}$$

Colligative Properties

Physical properties that depend only on the ratio of the number of solute particles to the number of solvent particles

- the *kind* of solute particles does not matter, only the *number* of solute particles
- the presence of dissolved particles changes the behavior of the solvent
- more solute particles result in a greater change

Types of solute particles include:

- molecules
- atoms
- ions (an *electrolyte* if water is the solvent)

Four colligative properties:

1. Boiling point elevation "The high gets higher"
2. Freezing point depression "The low gets lower"
3. Osmotic pressure
4. Vapor pressure depression

Boiling point elevation:

- adding a solute to water *raises* the boiling point "The high gets higher"
- the boiling point rises in direct proportion to the number of solute *particles*

Example: adding antifreeze to a car radiator also raises the boiling point of the water

- the optimal temperature for radiators is 93.3°C
- the boiling point of pure water is 100.0°C
- the boiling point of an antifreeze solution is 106.1°C, a safety margin of 6°C

Freezing point depression:

- adding a solute to water *lowers* the freezing point "The low gets lower"
- the freezing point drops in direct proportion to the number of solute *particles*

Example: salt is added to icy roads

- the average Syracuse temperature in January is 1°C to −9°C
- the freezing point of pure water is 0.0°C
- the freezing point of an ice/salt mixture on the roads is −15°C (~5°F), which works on all but the very coldest days

Comparing types of particles

- $Al_2(PO_4)_{3\,(s)} \rightarrow Al_2(PO_4)_{3\,(s)}$ (0 particles because $Al_2(PO_4)_{3\,(s)}$ is insoluble in water)
- $C_6H_{12}O_{6\,(s)} \rightarrow C_6H_{12}O_{6\,(aq)}$ (1 particle because glucose is molecular but soluble)
- $NaCl_{(s)} \rightarrow NaCl_{(aq)}$ (2 particles because sodium chloride is ionic and soluble)
- $CaCl_{2\,(s)} \rightarrow CaCl_{2\,(aq)}$ (3 particles because calcium chloride is ionic and soluble)

Ionic substances maximize colligative properties – each ion is a separate particle
The ions also conduct electricity and such compounds are called *electrolytes*
Use Table F to identify soluble electrolytes

Example: which sample, when dissolved in 1.0 L of water, produces a solution with the highest boiling point?

(1) 0.1 mole KI (3) 0.1 mole $MgCl_2$
(2) 0.2 mole KI (4) 0.2 mole $MgCl_2$

Answer: (4) 0.2 mole $MgCl_2$