CHEM 116 Solutions

FSG is Mondays 11:00 am and Tuesdays 3:30 pm both in S-1-89 Lecture 9 Prof. Sevian



Exam 1 is postponed to Thursday, October 2

Today's agenda

Note: If you printed out all of the lecture slides that were posted prior to last Thursday's lecture, then you already have what is in today's notes and you don't need to print them again.... I have moved those pages from what is now posted as Lecture 8 and placed them instead in these Lecture 9 slides. So if you didn't print Lecture 8 before, then print this Lecture 9 and bring it to class on 9/30/08.

Today we will cover material from chapter 13. This material will not be on Exam 1. It will be on Exam 2.

- · Solutions: what they are composed of
- · Solubility and dissolving
- Solubility curves



The extra credit is due at the beginning of class (2:00). Complicated explanations

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Mixed intermolecular forces

- Ion dipole
 - The ion causes the dipoles to orient in ways that increase attractions
 - This explains solvation of an ionic compound in water or other polar solvents (e.g., Na⁺ and Cl⁻ ions in water)
- Dipole non-polar molecule
 - The dipole causes the non-polar molecules to form more temporary dipoles
 - This causes the non-polar molecules to orient in ways that increase attractions
 - This explains why it is possible to dissolve tiny amounts of non-polar chemicals in a polar solvent (*e.g.*, oxygen gas in water)



About Exam 1 (Thurs Oct 2 in class)

- Chapters 10 and 11 will be covered on Exam 1
- Please read the information about exams on the course website
- Key points:
 - You will need a calculator devices that are capable of communication (e.g.,PDA devices, cell phones, some graphing calculators) are not allowed
 - Bring a pencil and eraser
 - Scratch paper will be provided you are not permitted to use your own
 - The exam starts at the beginning of class (2:00) and ends at the end of class (3:15) – please be on time so that you will have the complete time
 - Please sit every other seat in every other row

Beginning chapter 13 now

- From here forward, this material will not be on Exam 1
- Do not neglect the homework on chapter 13 in favor of studying for the exam
 - If you don't do the chapter 13 homework, you will get behind
 - Doing the chapter 13 homework will help you solidify your understanding of chapter 11

What is a solution?

Book definition: a solution is formed when one substance disperses uniformly throughout another (p. 530)

Terms you need to understand to make sense of this definition:

- substance
- disperses
- uniformly

What does "disperses" mean? Actions in chemistry that people often confuse:

- Dissolve
- Melt
- React

Which is which? And how can you tell the difference between them?

Categorize these as dissolving, melting or reacting, or a combination of these:

- Ice changes to liquid water when you leave it in a dish on the counter
- 2. Sugar crystals disappear when you mix them into hot tea
- 3. The contents inside a helium balloon escape into the air when you pop the balloon
- 4. An Alka-Seltzer tablet disappears when you plop it in water
- 5. When you add some blue salt crystals to a fish aquarium, the crystals disappear and the water turns bluish
- 6. When cooking an egg, sunny-side up, the clear liquid albumin turns white and solid
- Solid lipstick becomes liquid-like when you apply it to your warm lips
- 8. Saran[™] wrap becomes soft and gives off an unpleasant odor when accidentally placed in the oven

Recognizing solutions

- How did you decide which scenarios were examples of dissolving?
- Define solution in your own words.

Now compare to the book's definition

 a solution is formed when one substance disperses
 uniformly throughout another

The solutions we usually study are aqueous

An aqueous solution is:

- Pure water, and
- A pure substance (which was either a gas, liquid or solid before it got dissolved in the water) that has been dissolved in it.

Examples:

- Saline solution solid sodium chloride (NaCl) dissolved in water
- Seltzer water gaseous carbon dioxide (CO₂) dissolved in water
- Wine liquid ethanol (CH₃CH₂OH) dissolved in water

More generally, a solution contains:

- A pure solvent, and
- A pure solute that has been dissolved in it (or possibly more than one solute)

What does aqueous mean? (What does pure mean?)

Why are solutions interesting?

- Biology is all about (aqueous) solutions
 - Cells
 - Blood
 - Generally, how chemicals get from one place to another in living systems
- Medicine is all about (mostly aqueous) solutions
 - Medication
 - Assays
- Daily life is all about (aqueous) solutions
 - What you drink and eat
 - Cooking
 - Cleaning things
- Chemistry mostly happens in (aqueous and not) solutions

Solubility, saturation and supersaturation

Macroscopic

— As you add more solute to a solution

Below saturation

- Solvent is not yet holding the maximum amount of solute it can at that temperature
- More solute could still dissolve – if you add more, it will mix in



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At saturation point Beyond saturation

- Solvent is holding the maximum amount of solute it can at that temperature
- If you add any more solute, it will not dissolve (will fall to the bottom – precipitate)
 Solubility = often measured in grams of solute per 100

mL of solvent

more solute than it is able to at that

•

temperature – situation is unstable

Solvent is holding

 If you add anything to the solution, the excess (beyond saturation point) will crystallize out

How could you make a supersaturated solution?

Solubility = Maximum amount that can dissolve (at a given temperature)

 $[(Cation)^+(Anion)^-] (s)$ $\rightarrow (Cation)^+ (aq) + (Anion)^- (aq)$

e.g., KCl (s) \rightarrow K⁺ (aq) + Cl⁻ (aq)

- Positive slope
- Negative slope



How to read a solubility curve

- How much KCI would be able to dissolve in 100 g of water at 50°C?
- 2. At 40°C, exactly 64 g of an unknown salt dissolved in 100 g of water. What is the likely identity of the unknown?
- 3. At what temperature could you fully dissolve 80 g of NaCl in 200 g of water?
- 4. At 60°C, 30 g of KCIO₃ are dissolved in 100 g of water. Is the solution undersaturated, saturated or supersaturated?

Particle level

The process of dissolving

- Both ionic and molecular solids can dissolve in water (or other solvents)
- The difference is in the intermolecular attractions that hold the solute particles in the water
- Which one (ionic or molecular solute) is each of these pictures representing?



important thing is how many particles dissolve \rightarrow van't Hoff factor...coming soon



Enthalpy state of **Energetics of** solute ions (aq) dissolving $\Delta H_1 > 0$ process Enthalpy state of solute as crystal (s) (at ∞ dilution) ΔH_1 : Separation of solute molecules or Enthalpy state of Enthalpy of separated solvent molecules (aq) solution $\Delta H_2 > 0$ Enthalpy state of ΔH_2 : Separation of solvent molecules pure solvent (1) Enthalpy state of separated solute ions and solvent molecules $\Delta H_3 < 0$

 ΔH_3 : Formation of solute-solvent interactions Copyright © 2006 Pearson Prentice Hall, Inc.

Enthalpy state with solute-solvent attractions







Remember the solubility rules?

(from p. 127 in text)

TABLE 4.1 Solubility Guidelines for Common Ionic Compounds in Water				
Soluble Ionic Compounds		Important Exceptions		
Compounds containing	NO_3^-	None		
	$C_2H_3O_2^-$	None		
	Cl ⁻	Compounds of Ag^+ , Hg_2^{2+} , and Pb^{2+}		
	Br^{-}	Compounds of Ag ⁺ , Hg ₂ ²⁺ , and Pb ²⁺		
	I-	Compounds of Ag ⁺ , Hg ₂ ²⁺ , and Pb ²⁺		
	SO_4^{2-}	Compounds of Sr ²⁺ , Ba ²⁺ , Hg ₂ ²⁺ , and Pb ²⁺		
Insoluble Ionic Compounds		Important Exceptions		
Compounds containing	S ²⁻	Compounds of NH4 ⁺ , the alkali metal cations, and Ca ²⁺ , Sr ²⁺ , and Ba ²⁺		
	CO3 ²⁻	Compounds of NH4 ⁺ and the alkali metal cations		
	PO4 ³⁻	Compounds of $\mathrm{NH_4}^+$ and the alkali metal cations		
	OH-	Compounds of the alkali metal cations, and NH_4^+ , Ca^{2+} , Sr^{2+} , and Ba^{2+}		

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Solubility of Some Common Sulfate Compounds (at 25°C)

	Solubility	Solubility
Formula	(g/100 g H ₂ O)	(mol/L)
BaSO ₄	0.00246	0.00010
PbSO ₄	0.00425	0.00014
SrSO ₄	0.0113	0.00060
Hg ₂ SO ₄	0.060	0.0012
CaSO ₄	0.209	0.0150
Ag ₂ SO ₄	0.57	0.018
Na ₂ SO ₄	4.76	0.335
MgSO ₄	26.0	2.16

What affects solubility?

- Of an ionic compound dissolving in water
- Of a solute whose molecules have polar parts dissolving in a polar solvent (like water)
- Of a nonpolar solute dissolving in a polar liquid (like water)
- Of a nonpolar solute dissolving in a nonpolar solvent

In all these cases, it's a balancing act.

- Energy given off (becoming dispersed) vs. energy input (becoming more organized)
- Organization of particles (solutions are always more dispersed than the original pure solute plus pure solvent)

Spontaneity: Which is more favorable?

Particles in their undissolved state



solvent (I)



Particles in their dissolved state



Key points about solutions

- Solutions are mixtures (two or more pure substances mixed together uniformly)
- Solvent is the medium in which a solute gets dissolved
- Saturation is the point at which no more solute can dissolve in the solvent
- "Solubility" is the maximum amount of solute that can dissolve in a given amount of solvent (usually measured in grams of solute per 100 g of solvent)
- Particle-level model/understanding: Solutions form when attractions between solute and solvent are stronger than solute-solute and solvent-solvent attractions

What solutions can do, practical uses

- Boil at higher temperature than the pure solvent
 - Cooking
 - Keeps things in the liquid phase
- Freeze at lower temperature than the pure solvent
 - Making ice cream
 - Melting ice on sidewalks and streets in winter
 - · Keeps your car radiator liquid from freezing in winter
 - Keeps things in the liquid phase
- Create osmotic pressure
 - Dialysis
 - Cell assays
- Nature figured out a long time ago how to take advantage of these properties

What's significant about how particles are arranged inside a solution?

In the bulk

- Solute particles surrounded by solvent particles
- Solute-solvent attraction is stronger than a solvent-solvent attraction



At the surface

- Phase changes occur at surfaces
- Only solvent particles change phase (solute particles remain dissolved)
- As solvent particles leave the liquid phase, the solution becomes more concentrated in solute