CHEM 116

Solutions

FSG is
Mondays 11:00 am
and
Tuesdays 3:30 pm
both in
S-1-89

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

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:
1. 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
7. 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

As you add more solute to a solution

<table>
<thead>
<tr>
<th>Below saturation</th>
<th>At saturation point</th>
<th>Beyond saturation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Solvent is not yet holding the maximum amount of solute it can at that temperature</td>
<td>Solvent is holding the maximum amount of solute it can at that temperature</td>
<td>Solvent is holding more solute than it is able to at that temperature – situation is unstable</td>
</tr>
<tr>
<td>More solute could still dissolve – if you add more, it will mix in</td>
<td>If you add any more solute, it will not dissolve (will fall to the bottom – precipitate)</td>
<td>If you add anything to the solution, the excess (beyond saturation point) will crystallize out</td>
</tr>
</tbody>
</table>

**Solubility** = often measured in grams of solute per 100 mL of solvent

**How could you make a supersaturated solution?**
Solubility = Maximum amount that can dissolve (at a given temperature)

\[ [(\text{Cation})^+ (\text{Anion})^-] (s) \rightarrow (\text{Cation})^+ (aq) + (\text{Anion})^- (aq) \]

\[ \text{e.g., } \text{KCl} (s) \rightarrow \text{K}^+ (aq) + \text{Cl}^- (aq) \]

- Positive slope
- Negative slope

How to read a solubility curve

1. How much KCl 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 KClO₃ are dissolved in 100 g of water. Is the solution undersaturated, saturated or supersaturated?
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?

Interesting note: Most important thing is how many particles dissolve \( \rightarrow \) van’t Hoff factor...coming soon

Energetics of dissolving process (at \( \infty \) dilution) or Enthalpy of solution

\[ \Delta H_1: \text{Separation of solute molecules} \]

\[ \Delta H_2: \text{Separation of solvent molecules} \]

\[ \Delta H_3: \text{Formation of solute-solvent interactions} \]

Enthalpy state of solute ions (aq) \( \Delta H_j > 0 \)

Enthalpy state of solute as crystal (s)

Enthalpy state of separated solvent molecules (aq) \( \Delta H_j > 0 \)

Enthalpy state of pure solvent (l)

Enthalpy state with separated solute ions and solvent molecules \( \Delta H_j < 0 \)

Enthalpy state with solute-solvent attractions
Is a particular solution process endothermic or exothermic?

**Endothermic**
Ex: NH₄NO₃ (s) → NH₄NO₃ (aq)

- Magnitude of $\Delta H_3$
  - Energy released when attractions form between solvent molecules and solute particles
- Magnitude of $\Delta H_1 + \Delta H_2$
  - Energy input needed to separate solute particles
  + Energy input needed to separate solvent particles

**Exothermic**
Ex: NaOH (s) → NaOH (aq)

- Magnitude of $\Delta H_3$
  - Energy released when attractions form between solvent molecules and solute particles
- Magnitude of $\Delta H_1 + \Delta H_2$
  - Energy input needed to separate solute particles
  + Energy input needed to separate solvent particles

Remember the solubility rules?
(from p. 127 in text)

<table>
<thead>
<tr>
<th>TABLE 4.1 Solubility Guidelines for Common Ionic Compounds in Water</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Soluble Ionic Compounds</strong></td>
</tr>
<tr>
<td>Compounds containing NO₃⁻</td>
</tr>
<tr>
<td>C₂H₅O₂⁻</td>
</tr>
<tr>
<td>Cl⁻</td>
</tr>
<tr>
<td>Br⁻</td>
</tr>
<tr>
<td>I⁻</td>
</tr>
<tr>
<td>SO₄²⁻</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th><strong>Insoluble Ionic Compounds</strong></th>
<th><strong>Important Exceptions</strong></th>
</tr>
</thead>
<tbody>
<tr>
<td>Compounds containing S²⁻</td>
<td>Compounds of NH₄⁺, the alkali metal cations, and Ca²⁺, Sr²⁺, and Ba²⁺</td>
</tr>
<tr>
<td>CO₃²⁻</td>
<td>Compounds of NH₄⁺ and the alkali metal cations</td>
</tr>
<tr>
<td>PO₄³⁻</td>
<td>Compounds of NH₄⁺ and the alkali metal cations</td>
</tr>
<tr>
<td>OH⁻</td>
<td>Compounds of the alkali metal cations, and NH₄⁺, Ca²⁺, Sr²⁺, and Ba²⁺</td>
</tr>
</tbody>
</table>
Solubility of Some Common Sulfate Compounds (at 25°C)

<table>
<thead>
<tr>
<th>Formula</th>
<th>Solubility (g/100 g H₂O)</th>
<th>Solubility (mol/L)</th>
</tr>
</thead>
<tbody>
<tr>
<td>BaSO₄</td>
<td>0.00246</td>
<td>0.00010</td>
</tr>
<tr>
<td>PbSO₄</td>
<td>0.00425</td>
<td>0.00014</td>
</tr>
<tr>
<td>SrSO₄</td>
<td>0.0113</td>
<td>0.00060</td>
</tr>
<tr>
<td>Hg₂SO₄</td>
<td>0.060</td>
<td>0.0012</td>
</tr>
<tr>
<td>CaSO₄</td>
<td>0.209</td>
<td>0.0150</td>
</tr>
<tr>
<td>Ag₂SO₄</td>
<td>0.57</td>
<td>0.018</td>
</tr>
<tr>
<td>Na₂SO₄</td>
<td>4.76</td>
<td>0.335</td>
</tr>
<tr>
<td>MgSO₄</td>
<td>26.0</td>
<td>2.16</td>
</tr>
</tbody>
</table>

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

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