

## Chemistry 12

### Unit 3-Solubility Equilibrium-Unit Outline

#### 1. Quick Demonstration of Conductivity of Some Liquids

Put a Table like the following on the overhead and test the conductivity of the liquids with the light bulb assembly (in drawer of demo table). BE CAREFUL WITH THE PRONGS. THEY COULD ELECTROCUTE YOU IF YOU TOUCH THEM. ALWAYS UNPLUG IT AS SOON AS YOU FINISH TESTING A SOLUTION. THEY CAN BE RINSED WITH WATER (UNPLUGGED) BETWEEN SOLUTIONS.

| Liquid                               | Type                  | Ionic or Molecular | Conductivity (good, poor or none) |
|--------------------------------------|-----------------------|--------------------|-----------------------------------|
| H <sub>2</sub> O <sub>(l)</sub>      | Pure liquid           | molecular          |                                   |
| HCl <sub>(aq)</sub>                  | Acid solution         | ionic              |                                   |
| NaCl <sub>(aq)</sub>                 | Salt solution         | ionic              |                                   |
| NaOH <sub>(aq)</sub>                 | Base solution         | Ionic              |                                   |
| CH <sub>3</sub> OH <sub>(l)</sub>    | Liquid alcohol        | molecular          |                                   |
| CH <sub>3</sub> OH <sub>(aq)</sub>   | Alcohol solution      | molecular          |                                   |
| CH <sub>3</sub> COOH <sub>(l)</sub>  | Pure organic acid     | molecular          |                                   |
| CH <sub>3</sub> COOH <sub>(aq)</sub> | Organic acid solution | Partially ionic    |                                   |

#### NOTES:

- Concentrations of solutions are not important.
  - Pure methyl alcohol can be found in the yellow cabinet in the chem. storage room.
  - In going from CH<sub>3</sub>OH<sub>(l)</sub> to CH<sub>3</sub>OH<sub>(aq)</sub>, just add some water. (it still won't conduct)
  - Pure CH<sub>3</sub>COOH (glacial acetic acid) can be found in yellow cabinet. Again in going from CH<sub>3</sub>COOH<sub>(l)</sub> to CH<sub>3</sub>COOH<sub>(aq)</sub>, just add some water. If you add enough water, the light bulb will glow dimly. This would be called "poor conductivity"
  - At the end, you can do a little summary of how molecular and ionic relate to conductivity.
  - REMEMBER, ONLY HAVE THE LIGHT BULB PLUGGED IN LONG ENOUGH TO DO EACH TEST. NEVER TOUCH THE BARE WIRES!
2. After the demonstration, assign **Tutorial 7**
  3. Go over p. 73-74 in SW. Have them do Ex. 1(a-j) and 2(a-d) on p. 74 of SW.
  4. Have them do **Tutorial 8**
  5. Go over p. 75-76 in SW. Make sure they know the differences between net ionic equations for:
 

Dissolving:  $\text{CaCl}_{2(s)} \rightarrow \text{Ca}^{2+}_{(aq)} + 2 \text{Cl}^{-}_{(aq)}$  (single arrow, solid on left)

Crystallization (Precipitation):  $\text{Al}^{3+}_{(aq)} + 3 \text{Br}^{-}_{(aq)} \rightarrow \text{AlBr}_{3(s)}$  (single arrow, solid on right)

Equilibrium:  $\text{CaCO}_{3(s)} \rightleftharpoons \text{Ca}^{2+}_{(aq)} + \text{CO}_3^{2-}_{(aq)}$  (double arrow, solid on left)
  6. Have them do Ex. 3-7 on p. 76. They can look up answers and discuss them if you need to. (Watch out for a salt with a deadly weapon!)

7. Review unit conversion method of converting solubility in g/L to molar solubility (mol/L) and vice versa.(See p. 77 in SW.)

A good way to convert g/100.0 mL to g/L is eg:

$$\frac{0.75\text{g}}{100.0\text{mL}} \times \frac{100.0\text{mL}}{0.1000\text{L}} = 7.5\text{ g/L}$$

8. Have them do Ex. 8, 9, 10 and 11 on p. 77 of SW.
9. Discuss the graph of Sol. Vs. Temp. on Ex. 17 on p. 79 of SW. There should be a transparency of this in the “Chem 12 Unit 3” binder (which hopefully can be found in the left big wooden cupboard). I usually get them to try questions a-h on p. 79 and go over them. (it’s a little confusing for some of them)

10. Demonstration on Experimentally Determining the Solubility of a Saturated Solution  
Steps:

Find bottle of Saturated Calcium Acetate solution (on white shelf behind demo. table in the room.

- a. Set up a data table like the following (on the overhead)

|   |         |
|---|---------|
| Mass of empty evaporating dish  | g       |
| Mass of evaporating dish and dry $\text{Ca}(\text{CH}_3\text{COOH})_2$        | g       |
| Volume of saturated $\text{Ca}(\text{CH}_3\text{COOH})_2$ solution evaporated | 10.0 mL |

- b. Get an empty evaporating dish and weigh it on the electronic balance. Record mass in table.
- c. Pipette 10.0 mL of saturated Calcium Acetate ( $\text{Ca}(\text{CH}_3\text{COOH})_2$ ) solution from the bottle into the evaporating dish.
- d. Put the evaporating dish on a hot plate (fairly low) and let it evaporate while you do something else.
- e. After it’s evaporated, let it cool and weigh it again. Record the mass in the data table.
- f. From the information, have the students calculate the solubility in g/10.0mL, g/L and mol/L.
- g. Tell them to make sure they have this in their notes as there could be a question relating to this on the test!
11. Go over how to find individual ion concentration in a solution of an ionic compound. There is an example on the next page...

You have a 2.0M solution of  $\text{Fe}(\text{NO}_3)_3$ . Find the  $[\text{Fe}^{3+}]$  and the  $[\text{NO}_3^-]$ .

Write dissociation equation:  $\text{Fe}(\text{NO}_3)_3(\text{s}) \rightarrow \text{Fe}^{3+}(\text{aq}) + 3 \text{NO}_3^-(\text{aq})$

$$[\text{Fe}^{3+}] = 2.0\text{M Fe}(\text{NO}_3)_3 \times \frac{1\text{MFe}^{3+}}{1\text{MFe}(\text{NO}_3)_3} = \mathbf{2.0\text{M}}$$

$$[\text{NO}_3^-] = 2.0\text{M Fe}(\text{NO}_3)_3 \times \frac{3\text{MNO}_3^-}{1\text{MFe}(\text{NO}_3)_3} = \mathbf{6.0\text{M}}$$

NOTE: Sometimes questions try to confuse you by giving you information you don't need.

Coefficient ratios can apply to moles or M. Also remember that  $M = \frac{\text{mol}}{\text{L}}$

eg.) The  $[\text{Na}^+]$  in 2.0 L of a solution of  $\text{Na}_3\text{PO}_4$  is 1.5M. Find the  $[\text{PO}_4^{3-}]$

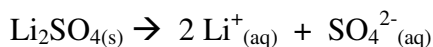
Since  $[\text{Na}^+]$  and  $[\text{PO}_4^{3-}]$  are both in Molarity (M), the 2.0L volume is NOT needed in the calculation.



$$[\text{PO}_4^{3-}] = 1.5\text{M Na}^+ \times \frac{1\text{MPO}_4^{3-}}{3\text{MNa}^+} = \mathbf{0.50\text{M}}$$

But sometimes you DO need to use the volume!

eg.) 2.0 mol of  $\text{Li}_2\text{SO}_4$  are dissolved in 8.0 L of solution. Find the  $[\text{Li}^+]$ .



$$[\text{Li}_2\text{SO}_4] = \frac{2.0\text{mol}}{8.0\text{L}} = 0.25\text{M}$$

$$[\text{Li}^+] = 0.25\text{M Li}_2\text{SO}_4 \times \frac{2\text{MLi}^+}{1\text{MLi}_2\text{SO}_4} = \mathbf{0.50\text{M}}$$

12. Have students do Ex. 18 (a,b,c) on p. 81 of SW.

13. Review the **Dilution Formula**:

The SW uses:  $[\text{SUBSTANCE}]_{\text{diluted}} = [\text{SUBSTANCE}]_{\text{old}} \times \frac{\text{OLD VOLUME}}{\text{DILUTED VOLUME}}$

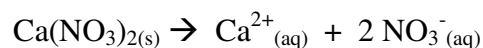
In Chem. 11, I taught it as:  $FC \times FV = IC \times IV$

$$\text{Or } FC = IC \times \frac{IV}{FV}$$

Where:  
 F = final  
 I = initial  
 C = concentration  
 V = volume  
 FV = final (total) volume

eg.) 250.0 mL of water are added to 350.0 mL of 0.40M  $\text{Ca}(\text{NO}_3)_2$ . Find the final concentrations of both ions.

$$FC = IC \times \frac{IV}{FV} \quad \text{Final } [\text{Ca}(\text{NO}_3)_2] = 0.40\text{M} \times \frac{350.0\text{mL}}{600.0\text{mL}} = 0.233\text{M}$$



$$[\text{Ca}^{2+}] = 0.233\text{M } \text{Ca}(\text{NO}_3)_2 \times \frac{1\text{M}\text{Ca}^{2+}}{1\text{M}\text{Ca}(\text{NO}_3)_2} = \mathbf{0.23\text{M}}$$

$$[\text{NO}_3^-] = 0.233\text{M } \text{Ca}(\text{NO}_3)_2 \times \frac{2\text{M}\text{NO}_3^-}{1\text{M}\text{Ca}(\text{NO}_3)_2} = \mathbf{0.47\text{M}}$$

**Mixing Solutions Without Common Ions**

When two solutions are mixed, they don't react with each other, and there are no common ions, each solution "see's" the other solution as if it was water.

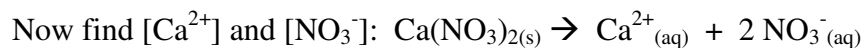
eg.) 20.0 mL of 0.20M NaCl is mixed with 30.0 mL of 0.15M  $\text{Ca}(\text{NO}_3)_2$ . Calculate the final concentrations of all four ions.

First calculate final [NaCl] (*consider the  $\text{Ca}(\text{NO}_3)_2$  solution like adding water.*)

$$\text{Final } [\text{NaCl}] = 0.20\text{M} \times \frac{20.0\text{mL}}{50.0\text{mL}} = 0.080\text{M}$$

Find the  $[\text{Na}^+]$  and  $[\text{Cl}^-]$ :  $\text{NaCl}_{(\text{s})} \rightarrow \text{Na}^+_{(\text{aq})} + \text{Cl}^-_{(\text{aq})}$   
 Since all coefficients are 1,  $[\text{Na}^+] = \mathbf{0.080\text{M}}$   
 $[\text{Cl}^-] = \mathbf{0.080\text{M}}$

$$\text{Now, find final } [\text{Ca}(\text{NO}_3)_2] = 0.15\text{M} \times \frac{30.0\text{mL}}{50.0\text{mL}} = 0.090\text{M}$$



$$[\text{Ca}^{2+}] = [\text{Ca}(\text{NO}_3)_2] = \mathbf{0.090\text{M}}$$

$$[\text{NO}_3^-] = 0.090\text{M Ca}(\text{NO}_3)_2 \times \frac{2\text{MNO}_3^-}{1\text{M Ca}(\text{NO}_3)_2} = \mathbf{0.18\text{M}}$$

$$\begin{aligned} \text{So, in summary: } [\text{Na}^+] &= \mathbf{0.080\text{M}} \\ [\text{Cl}^-] &= \mathbf{0.080\text{M}} \\ [\text{Ca}^{2+}] &= \mathbf{0.090\text{M}} \\ [\text{NO}_3^-] &= \mathbf{0.18\text{M}} \end{aligned}$$

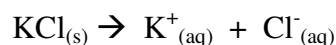

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14. Have students do Ex. 20 (d & e) on p. 81 of SW.  
 15. Here is an example of what do you when you have mixtures of solutions with a common ion. (an ion that is common to both solutions)

### Mixtures With Common Ions

50.0 mL of 0.20M KCl is mixed with 60.0 mL of 0.30M MgCl<sub>2</sub>. Find the final concentrations of all three ions.

$$\text{Final } [\text{KCl}] = 0.20\text{M} \times \frac{50.0\text{mL}}{110.0\text{mL}} = 0.0909\text{M}$$

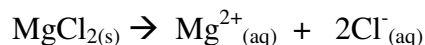


$$[\text{K}^+] = [\text{KCl}] = 0.091\text{M}$$

$$[\text{Cl}^-] \text{ (from KCl)} = [\text{KCl}] = 0.0909\text{M}$$


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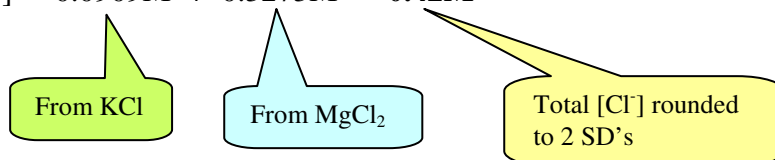
$$\text{Final } [\text{MgCl}_2] = 0.30\text{M} \times \frac{60.0\text{mL}}{110.0\text{mL}} = 0.1636\text{M}$$



$$[\text{Mg}^{2+}] = [\text{MgCl}_2] = 0.16\text{M}$$

$$[\text{Cl}^-] \text{ (from MgCl}_2) = 0.1636\text{M MgCl}_2 \times \frac{2\text{MCl}^-}{1\text{M MgCl}_2} = 0.3273\text{M} \text{ (continued over...)}$$

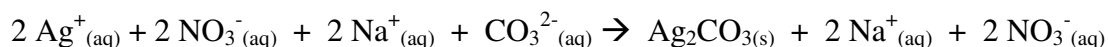
$$\begin{aligned} \text{Final [K}^+] &= \mathbf{0.91M} \\ \text{Final [Mg}^{2+}] &= \mathbf{0.16M} \\ \text{Final [Cl}^-] &= 0.0909M + 0.3273M = \mathbf{0.42M} \end{aligned}$$



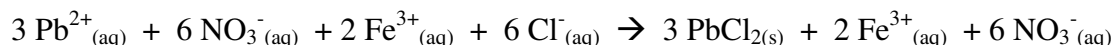
16. Have students do Ex. 20 (f & g) on p. 81.
17. Students can now be assigned the “**Worksheet on Solubility Calculations**” (called WorksheetSolCalc) on the Chem. 12 Webpage. The Key is also on there so they can check their answers. (Make this one a “show me”.)
18. Have students take out their Data Booklets and turn to the “Solubility of Common Compounds in Water” table (from now on we’ll call it the “Solubility Table” Have them go over the section on “Predicting the Solubility of Salts” from page 81-83 of SW.
19. Have students do Ex. 21 (a-j) and 22 (a-f) on page 83 of SW.
20. Carefully go over the section on “Writing Formula, Complete and Net-Ionic Equations” on pages 84-87 in SW.

**NOTE:** I have them include the “<sub>(aq)</sub>” after every individual ion in the Complete Ionic Equation (I notice the book doesn’t)

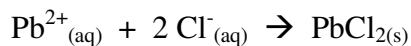
Eg.) The complete ionic equation on p. 85, should look like:



Near the bottom of page 86, the Complete Ionic Equation should read:



and the final Net-Ionic Equation should read:



Tell students it is very important that they know the differences between a Formula, CIE and NIE. If they are asked to write a specific one of these on a test, they will not get full marks for writing the wrong type.)

21. Have students do Ex. 25 (a,e,f &h) on page 87 of SW. (They can check their answers on p. 269-270. Notice he **does** include all of the “<sub>(aq)</sub>”s in his answers to these questions.)
22. Assign **Worksheet 3-1** (or “Homework Assignment 3-1”) from the Chem 12 Webpage. This will be a hand-in “marked” assignment.

23. Students can print off and do **Tutorial 9** from the Chem. 12 Webpage.
24. After doing Tutorial 9, students can read “Separation of Mixtures of Ions by Precipitation Methods from p. 88-89 in SW.
25. Have students do Ex. 27, 29, 31 & 33 on page 90 of SW.
26. Students can now print and do **Tutorial 10** from the Chem. 12 Webpage.
27. Have students carefully read the section “The Solubility Product” from page 91-95 in SW. It is really important that students do NOT get “Solubility” and “Solubility Product (Ksp)” mixed up. Also, for “AB<sub>2</sub>” compounds (when the 2 ions are in a 1:2 or 2:1 ratio), make sure the algebra is done correctly.

For “AB” compounds (like AgCl)  $K_{sp} = s^2$  or  $s = \sqrt{K_{sp}}$  (where “s” is molar solubility)

For “AB<sub>2</sub>” comp. (like PbCl<sub>2</sub>)  $K_{sp} = 4s^3$  or  $s = \sqrt[3]{\frac{K_{sp}}{4}}$

NOTE: These expressions must be derived by the methods shown in Tutorial 10.

Students cannot just memorize these equations and plug #'s in to get the answer. They are just a good little “check” to make sure they have done the right thing.

28. Have students do Ex. 43, 44, 46, 49, 52, 54 & 55 on p. 95 of SW. (Check answers on p.273-274) At this point, tell students to print out Tutorial 11 (long one)
29. Students can now do Experiment 16E “Qualitative Analysis” on p. 180 of Heath Lab Manual. (The solutions and reagents are in trays in the Prep. Room) Half the groups can start with Part 1 and the other half with Part 2, then they can switch when they are finished. Have them read through ALL the procedures on Part 2 before they do it and it is a little confusing and they might throw something away that they need to use in the next procedure!) NOTE: For Part 1, have them just include the “Unknown Cation” in Procedure 2 and add the stuff to it as well as the others in Procedures 3-6. Also for Part 2, have them include the “Unknown Anion” in Procedure 1 and add stuff to it as well as all the others in Procedures 2-8. NOTE: When they are adding HNO<sub>3</sub> or NH<sub>3</sub> in Part 2, they are observing whether they appear to DISSOLVE precipitates or not. The questions for the lab are just these: (They may be on a transparency in the Unit 3 binder?)
  1. In Part 1, according to our results, the unknown cation is most likely \_\_\_\_\_.
  2. Write a balanced net-ionic equation for each precipitate formed in Part 1.  
eg.  $\text{Ba}^{2+}_{(aq)} + \text{CrO}_4^{2-}_{(aq)} \rightarrow \text{BaCrO}_{4(s)}$  (leave some spaces)
  3. In Part 2, according to our results, the unknown anion is most likely \_\_\_\_\_.
  4. Write a balanced net-ionic equation for each precipitate formed in Part 2.  
eg.  $2\text{Ag}^{+}_{(aq)} + \text{SO}_4^{2-}_{(aq)} \rightarrow \text{Ag}_2\text{SO}_{4(s)}$   
(NOTE: Neither HNO<sub>3</sub> nor NH<sub>3</sub> form any precipitates, so there are no equations involving these!)

30. Have students print out and do **Worksheet 3-2** from the Chem. 12 Webpage.
31. Have students read (or go over with them) **Tutorial 11** pages 1-7 (from Chem. 12 Webpage).
32. After they have finished and checked the question on p. 7 of Tutorial 11, have them do Ex. 61 on page 98 of SW. (They may want to read p. 96-97 in SW.)
33. Have students read (or go over with them), p. 8-11 on Tutorial 11 and do the question on p. 11.

34. Have students read (or go over with them), Tutorial 11 pages 12-16, having them do and check the problems in that section.
  35. Have students do Ex. 67 on page 99 of SW.
  36. Have students read (or go over with them) from the bottom of p. 16 to page 19 of Tutorial 11. They must do and check Question 5 on the bottom of p. 19.
  37. Have students read (or go over with them) the section on “*Finding Which Precipitate Will Form First*” on pages 20-23 of Tutorial 11. Make sure they do question 6 on p. 23.
  38. Assign the Self-Test for Tutorial 11
  39. Have them do Ex. 68 on p. 99 of SW.
  40. Have students do **Tutorial 12** “Precipitation Titrations”. They may need some explanation of titrations and the calculations involved. Assign the Self-Test on Tutorial 12.
  41. Have students do Ex. 70 on p. 101 of SW.
  42. Do **Experiment 19D** (Part III ONLY!) on p. 225 of Heath Lab Manual. Use commercially prepared 0.100 M silver nitrate solution. There should be some in brown bottles in a tray in the chemical storage room. (You should find some dropper bottles of  $\text{Na}_2\text{CrO}_4$  (yellow solution) in the same tray.) I usually send a couple of students to the pool to get a big jug of pool water to analyze for chloride ( $\text{Cl}^-$ ). They can use a data table like the one for Part III on page 227 of Heath Lab Manual. Have them use 10.0 mL samples of pool water. They can use pipettes to measure the 10.0 mL precisely. Make sure they don’t forget to add 10-20 drops of  $\text{Na}_2\text{CrO}_4$  each time they get another sample. If they have unused and un-polluted  $\text{AgNO}_3$  solution left at the end, they can put it back in the bottle. Have them do the calculations for Part III on p. 228 of the Heath Lab Manual. A partner lab with one write up per group (with all names on it) would be fine.
  43. Have students print out **Tutorial 13** and read it. You may have to go over some of the concepts in it. Make sure they understand the explanations for why solubilities change. Assign the Self-Test on Tutorial 13
  44. Have students read p. 105-107 in SW. Have them do Ex. 82-85 on p. 108 of SW.
  45. Assign the “**ESP**” (Extra Solubility Problems) **Worksheet** found on the Chem 12 Webpage. (The Key is on there so they can check it) Let them know that **now** is the time to get help with any of the questions they don’t understand. There are many types of problems in Unit 3, and it’s easy to get them mixed up and not know which strategy to use for which problem. Practice is essential at this point.
  46. Give them a bit of time to do **Unit 3 Review** (on Chem 12 Webpage) I don’t usually give marks for doing reviews.
  47. Hopefully you can also find the folder with “**Provincial Questions for Unit 3**” on my shelves somewhere.
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