# Chapters 16 & 17. Cracolice, 2/e. Chemistry X11.

**Grade option forms**. After Test 2 is returned, I will ask you to choose a grade option for the course. You can choose to receive a letter grade, pass/no pass, or "no credit" (audit). Some employers require a letter grade in order to reimburse students; check your employer's policy. Those not familiar with the Extension grading system should look at the section on Grades in a current catalog. Feel free to talk with me about your situation. However, you can estimate your current course grade by averaging your letter grades on the first two tests.

If you choose the pass-no pass option, you need to take the tests and receive at least a C average to pass.

**WRITTEN ASSIGNMENT #4**, due as we finish Ch 16. Ch 16: #14, 17, 19, 20, 26, 42, 43, 47, 48. As usual, aim for clear set-up, with clear units. That is the secret to doing harder problems. And how about contributing one more "Part 2" question?

$\Rightarrow$ See attached page for some problems, at the basic level, on two topics that have inadequate problems in the book. Thus I consider these problems important.

$\Rightarrow$ A **practice quiz** for Ch 17 is at the web site. It is good practice with determining what happens when something dissolves. Strongly recommended.

## Chapter 16

Read Ch 16 Sect 1-6, 10-11. Ch 16 is the first of a series of chapters dealing with solutions. We will consider the <u>terminology</u> of solutions (e.g., solute, solvent) and <u>concentration</u> <u>units</u> (such as weight percent and molarity, Sect 5-6). Once we know what solutions are and how to express their concentration, we can extend our discussion of <u>stoichiometry</u> (Ch 9) to reactions in solutions (Sect 11). The key is to calculate moles from concentration; that's easy when concentration is in molarity. We will include such practical topics as how to <u>make</u> a solution -- from the solid solute, or by <u>dilution</u> from a more concentrated stock solution (Sect 10).

$\Rightarrow$ One common way to indicate the molarity of a solution is to use square brackets. That is, the statement "The concentration of sodium ions is 0.1 mol/L" can be written as "$[Na^+] = 0.1$ M". Cracolice introduces this in Ch 18, page 513.

> If you miss the class discussion of dilutions, where I emphasize an alternative
> to the dilution equation (Eq 16.11), you might want to look at my
> supplementary worksheet on dilutions, especially Sections A-D.

The skipped sections: Sect 7-8 present two additional concentration units. I suggest that you not worry about these until you are comfortable with percentage and molarity, the more common units -- or unless you need them. Sect 9 summarizes the various kinds of concentration units. It's good, but remember that we are skipping two of them. Sect 12-13 present titrations, a special case of solution stoichiometry (that is, an extension of Sect 11); the two sections use different units. Some of you may need to know about acid-base titrations. I

suggest that you first do Sect 11, then try Sect 12. Note that Sect 12-13, on acid-base reactions, may be more useful after Ch 18.

> [The symbol for molality (Sect 7) is generally written as italic *m* (or underlined m) (presumably to avoid confusion with meter).]

> [If you work in a lab where "normality" is used for concentration, then you'll need Sect 8 & 13. Normality is no longer a recommended unit, but it's still fairly common -- and convenient. Normality is very much like molarity, but deals with "equivalents" rather than real moles.]

Sect 14 discusses why anti-freeze works. More generally, it discusses several properties of solutions that depend on how many solute particles there are but not on what they are. These are called colligative properties. You might want to browse this for the idea.

If you need some of these skipped sections, feel free to talk to me about them, or try (and turn in) some of the problems.


**Chapter 17**

Chemistry in solutions. Electrolytes and non-electrolytes. Ions and ionic equations. Acids and bases (continues into Ch 18; and recall Ch 6 Sect 5-7). Single and double replacement reactions (recall Ch 8 Sect 8-10). Solubility "rules".

The logical (and safe) way to write a net ionic equation is to write the complete equation, then re-write all the aqueous strong electrolytes in ionic form, then cancel out the spectator ions. With experience, you may see through the intermediates, and just write down the net ionic equation. But, as usual, I don't encourage you to skip steps -- until you are comfortable.

The solubility rules guide you when you don't have specific data at hand. Most of the rules have exceptions; emphasize the "stronger" rules -- those without exceptions. Also note a general trend that salts with ions of +/- 1 charge tend to be more soluble than salts with ions of higher charge. The reason for this trend is presumably that salts of ions with higher charge are held together, as salts, more strongly. Of course, AgCl is an exception.

You should practice with the problems at the end of the chapter. Emphasize seeing what happens in reactions involving ionic chemicals. This includes being able to write chemicals in the ionic form, and then going on to write net ionic equations. Try to do some for each section; being able to go on and do some of the "ungrouped" questions is always the best way to test your understanding. As you do these questions, look up data (activity series, solubility) that you need. For most people, there is little value in memorizing these tables in an intro course. I do expect you to know some broad trends (e.g., more active metals towards the left; all nitrates are soluble), but it is most important to practice the ideas.

⇒ I strongly encourage you to show me some net ionic equations that you have tried to work out, showing your work. The time schedule is tight near the end for turning things in and getting feedback, so I am reluctant to write a formal assignment. But there is often

confusion at first about how to do these, so having me check some, even a few moments before or after class, may be very useful. This is an important topic, which will almost certainly be on the test.


**Errata** (both chapters)

p 467, Table 16.2. For hexane, the atom at the left should be H not I.

p 498. In #5 and in #6, the second formula listed is not possible. (Same in answer.)

p 499, #29. The net ionic equation given in the answer is incorrect, since $PbSO_4$ is insoluble.


**Further reading**

G Shiflet, Materials science: The more elements, the merrier. Science 300:443, 4/18/03. News. Design of metal alloys (which are solid solutions) typically has been largely empirical. Here they discuss some recent advances in computational approaches to making better alloys. The story starts by noting the role of a new aluminum alloy in the original Wright brothers' plane; in December 2003 we commemorated the centennial of its initial successful flight.

D Erhardt, Materials science: The art of restoration. Nature 431:410, 9/23/04. News. Discusses new developments in cleaning old art work. Part of the secret is a careful consideration of the solvent properties of the cleaning agent, in view of the nature of both the "grime" to be removed and the underlying art that is not to be damaged.

G Michalski et al, Long term atmospheric deposition as the source of nitrate and other salts in the Atacama Desert, Chile: New evidence from mass-independent oxygen isotopic compositions. Geochimica et Cosmochimica Acta 68:4023, 10/15/04. Nitrate salts are typically very soluble, but deposits of nitrate salts can be found in (and mined from) ultra-dry desert regions. The northern Chilean Atacama desert is one of the driest places on earth, with 1-2 mm of rain per year. This paper explores the origins of such deposits. Their main conclusion -- for this desert -- is that the salts arise from atmospheric photochemical oxidation of nitrogen oxides, followed by particle deposition. The authors are from UC San Diego; they include the Mojave Desert in their discussion.

C M Dobson, Chemical space and biology. Nature 432:824, 12/16/04. Introduction to a feature section, on the types of molecules found in biology, and the implications for drug discovery.

S J Hawkes, Gravimetric titrations. J Chem Educ 81:1715, 12/04. (+ accompanying letter by R W Ramette.) As the book indicates, titrations are commonly done by measuring volumes of the solutions used. Here, Hawkes proposes that one should measure mass instead. Why? It has long been true that mass measurements have higher precision than volume measurements. But it used to be that mass measurements were slow. Now, with modern electronic balances, high speed and high precision mass measurements are routine.

R E Casey & F A Pittman, Intermolecular forces as a key to understanding the environmental fate of organic xenobiotics. J Chem Educ 82:260, 2/05. How do pollutants move through the ground? Well, that depends on their properties, such as how soluble they are in water. This article discusses issues such as this, and is probably suitable reading for students at X11 level.

T D Jickells et al, Global iron connections between desert dust, ocean biogeochemistry, and climate. Science 308:67, 4/1/05. Review. This relates to a story that has been in the news lately: how the transport of dust in the atmosphere may affect ocean productivity. A key part of that is transport of iron in the dust. Complicated, but even a quick browse may help you realize how several seemingly unrelated topics come together. For more about the problems due to the insolubility of Fe(III), see Abergel et al (2006).

S Weiner et al, Structural biology: Choosing the crystallization path less traveled. Science 309:1027, 8/12/05. News. A discussion of growth of crystals, especially bio-crystals, such as the calcium carbonate found in many types of skeletons.

R J Abergel et al, Anthrax pathogen evades the mammalian immune system through stealth siderophore production. PNAS 103:18499, 12/5/06. The insolubility of Fe(III) presents a serious problem for most aerobic organisms. Fe(III), the common form of Fe under oxidizing conditions, is generally essential, but is quite insoluble. Thus most organisms, from bacteria to humans, have complex systems for obtaining the Fe they need. Bacteria that invade humans have systems for obtaining Fe from the host -- and the hosts may have systems to defeat the bacterial acquisition of iron. This article discusses some recent work on this host-pathogen competition for iron. The specific case is that of the anthrax bacterium, and the work is, in part, from UC Berkeley. I have listed this paper, along with a news story about it in the student newspaper, on my web pages for this course and also for BITN, under the miscellaneous topic of anthrax. Also see Jickells et al (2005).


**<u>Computer resources & Supplementary materials</u>** (See web page for details and links.)

Three of the self-help worksheet files posted on my Chemistry Problems web page are relevant to Ch 16. One introduces molarity. One is on percentages; it discusses various concentration units that use percentage, including %(w/w) and %(w/v) and similar units such as ppm (parts per million). The third is a worksheet on dilutions. It includes discussion of the dilution equation. If you miss the class discussion of dilutions, where I emphasize an alternative to the dilution equation (Eq 16.11), you might want to look at this worksheet, especially Sections A-D. All these optional worksheets provide many practice problems, from basic to difficult, with answers.

As noted above, a practice quiz for Ch 17 is at the web site. It is good practice with determining what happens when something dissolves. Strongly recommended.

The common solubility rules -- in verse, and a cartoon showing what molarity means.

⇒ **Attached page**: Problems for Ch 16 Sections 5 & 10. These are topics where I think the problems in the book are inadequate. Thus I consider the attached problems important. Solutions included; you can do these on your own.