Contents

Preface

This manual starts off with a section entitled "Lecture Schedule," which you may find helpful in adapting the text to your class schedule. Beyond that, the manual is organized by text chapters. For each chapter, we include three different features:

1. "Lecture Notes," which suggest the amount of time that we devote to each chapter and the topics we emphasize. Included are detailed lecture outlines (from our own lectures) that may serve as a guide for your lectures. At a minimum, they indicate how we cover topics and how successive topics can be integrated.

2. A list of demonstrations illustrating topics in the chapter. These are taken from three sources:

- The manual *Tested Demonstrations in Chemistry* (1994), Volumes I and II, compiled and edited by by George Gilbert by arrangement with the Journal of Chemical Education. These are coded as "GILB" with the experiment number (e.g., M 12)
- *Chemical Demonstrations* (1983-1992), Volumes 1-4, published by Bassam Shakashiri with many collaborators and contributors. These are listed as "SHAK" followed by the volume and page reference.
- Demonstrations described in the *Journal of Chemical Education* with the journal reference.

3. Answers and detailed solutions to:

- The summary problem at the end of each chapter
- Odd numbered text problems
- Challenge problems at the end of each problem set

Note that Appendix 6 has answers to all the even-numbered problems. Detailed solutions to many of these problems are in the *Student Solutions Manual* available from Cengage Learning, Brooks/Cole.

Lecture Schedule

Unlike other general chemistry texts, ours can be covered in its entirety in a one-year course. A reasonable schedule appears below. Further comments on the time you should devote to each chapter are in the body of this manual. The underlying assumption is that you are teaching 14-week semesters with two 50-minute lectures per week. If three class periods are devoted to examinations each semester, that leaves 25 for covering material. On that basis, you are able to complete Chapter 10 (Solutions) in the first semester. The seond semester will then start with Chapter 11 (Rate of Reaction).

FIRST SEMESTER SCHEDULE

Week	Lecture	Topic
1	1	Chapter 1 (Matter and Measurements)
	2	Chapter 1
2	3	Chapter 2 (Atoms, Molecules, and Ions)
	4	Chapter 2
3	5	Chapter 3 (Mass Relations in Chemistry; Stoichiometry)
	6	Chapter 3
4	7	Chapter 3
	8	EXAM I
5	9	Chapter 4 (Reactions in Aqueous Solution)
	10	Chapter 4
6	11	Chapter 4
	12	Chapter 5 (Gases)
7	13	Chapter 5
	14	Chapter 6 (Electronic Structure and the Periodic Table)
8	15	Chapter 6
	16	Chapter 6
9	17	EXAM II
	18	Chapter 7 (Covalent Bonding)
10	19	Chapter 7
	20	Chapter 7
11	21	Chapter 8 (Themochemistry)
	22	Chapter 8
12	23	Chapter 9 (Liquids and Solids)
	24	Chapter 9
13	25	Chapter 9
	26	EXAM III
14	27	Chapter 10 (Solutions)
	28	Chapter 10

SECOND SEMESTER SCHEDULE

Week	Lecture	Topic
1	1	Chapter 11 (Rate of Reaction)
	2	Chapter 11
2	3	Chapter 11
	4	Chapter 12 (Gaseous Chemical Equilibrium)
3	5	Chapter 12
	6	Chapter 13 (Acids and Bases)
4	7	Chapter 13
	8	Chapter 13
5	9	EXAM I
	10	Chapter 14 (Equilibria in Acid-Base Solutions)
6	11	Chapter 14
	12	Chapter 15 (Complex Ions)
7	13	Chapter 15
	14	Chapter 16 (Precipitation Equilibria)
8	15	Chapter 17 (Spontaneity of Reaction)
	16	Chapter 17
9	17	EXAM II
	18	Chapter 18 (Electrochemistry)
10	19	Chapter 18
	20	Chapter 18
11	21	Chapter 19 (Nuclear Chemistry)
	22	Chapter 20 (Chemistry of the Metals)
12	23	Chapter 20
	24	EXAM III
13	25	Chapter 21 (Chemistry of the Nonmetals)
	26	Chapter 21
14	27	Chapter 22 (Organic Chemistry)
	28	Chapter 23 (Organic Polymers: Natural and Synthetic)

If you want to use lecture time for review, for going over assigned problems, or for doing a large number of demonstrations, you will have trouble keeping up with this schedule. As you've almost certainly learned by now, the solution to this problem is not to talk faster. Judicious deletions work better. It's been said, and wisely, that the secret of giving a good lecture is knowing what to leave out. Possible candidates include:

- Introductory material on matter in Chapter 1 and atomic theory in Chapter 2. The chances are your students have been exposed to this material more than once in high school and understood it reasonably well the first time.

- Boyle's and Charles's laws in Chapter 5. We start the chapter by writing the ideal gas law and go on from there.

- The First Law discussion in Chapter 8. Quite frankly, this has very little to do with chemistry. Students will not be irreparably damaged if they are unaware of the distinction between H and E.

- The discussion of colligative properties in Chapter 10 could be shortened. Raoult's law could easily be omitted.

- Reaction mechanisms in Chapter 11. Students have a lot of trouble with this. We are not sure it is worth the effort.

- Polyprotic acids in Chapter 13.

- The Second Law discussion in Chapter 17.

Beyond these selective omissions, some instructors may want to delete one or another of the descriptive chapters at the end of the text (Chapters 20–22). If, in that way, you can squeeze out a couple of lectures, they can well be spent on Chapter 12 (three lectures instead of two) and Chapter 19 (two lectures instead of one).

Textbook authors sometimes tell you that chapters can be covered in almost any order, depending on your preference. This isn't really true for this textbook, or any other with structural integrity. It can be done, but only with very careful additions and deletions of material. Suppose, for example, you want to cover Precipitation Equilibria (Chapter 16) immediately after Acid-Base Equilibria (Chapter 14). Keep in mind that an understanding of formation constants (Complex Ions, Chapter 15) is assumed when methods of dissolving precipitates are considered in Section 16.2 of Chapter 16.

1 MATTER AND MEASUREMENTS

LECTURE NOTES

This material ordinarily requires two lectures (100 minutes), allowing for a 10–15 minute introduction to the course in the first lecture. If you're in a hurry, this can be cut to $1\frac{1}{2}$ lectures by discussing only quantitative material (significant figures, unit conversions, density, solubility).

A few points to keep in mind:

- Virtually all of your students will be familiar with the metric system and prefixes. It may be worth discussing the rationale for SI, but you don't have to dwell on it.
- Students readily learn the rules of significant figures, but typically ignore them after Chapter 1. It may help to emphasize that these are common-sense (albeit, approximate) rules for estimating experimental error.
- Many (typically, the weaker) students resist using conversion factors, preferring instead a rote method. It may be useful to point out that conversion factors will be a recurring tool throughout the text, so are well worth learning at this point.
- Students often have trouble with solubility calculations. The approach in the text involves conversions (Example 1.8). The solubility is considered to be a conversion factor relating grams of solute to grams of solvent.

Lecture 1

I. Types of Substances

A. Elements

Cannot be broken down into simpler substances. Examples: nitrogen, lead, sodium, arsenic. Symbols: N, Pb, Na, As.

B. Compounds

Contain two or more elements with fixed mass percents. Glucose: 40.00% C, 6.71% H, 53.29% O. Sodium chloride: 39.34% Na, 60.66% Cl.

C. Mixtures

Homogeneous (solutions) vs. heterogeneous. Separation by filtration, distillation.

II. Measured Quantities

A. Length

Base unit is the meter. 1 km = 10^3 m; 1 cm = 10^{-2} m; 1 mm = 10^{-3} m; 1 nm = 10^{-9} m. Dimensions of very tiny particles will be expressed in nanometers.

B. Volume

1 L = 10^3 mL = 10^3 cm^3 = 10^{-3} m^3. Buret, pipet, volumetric flask.

C. Mass

1 kg = 10^3 g; 1 mg = 10^{-3} g. Two different kinds of balances will be used in the lab. An analytical balance ($\pm$0.001 g) should be used only for accurate, quantitative work.

D. Temperature

$t_{\circ F} = 1.8\,t_{\circ C} + 32^\circ$; $T_K = t_{\circ C} + 273.15$.
Convert 68°F to °C and K: $t_{\circ C} = (68^\circ - 32^\circ)/1.8 = 20^\circ C$ $T_K = 293$

Lecture 2

III. Experimental Error; Significant Figures

Suppose an object is weighed on a crude balance to $\pm$0.1 g and the mass is found to be 23.6 g. This quantity contains three significant figures, that is, three experimentally significant digits. With an analytical balance, the mass might be 23.582 g (five significant figures).

A. Counting significant figures

1. Volume of liquid = 24.0 mL; three significant figures. Zeroes at the end of the measured quantity are significant when they follow nonzero digits.

2. Volume = 0.0240 L; three significant figures (note that 0.0240 L = 24.0 mL). Zeroes at the beginning of a measured quantity are not significant when they precede nonzero digits.

B. Multiplication and division

Keep only as many significant figures as there are in the least precise quantity. Density of a piece of metal weighing 36.123 g with a volume of 13.4 mL = ?

$$\text{density} = \frac{36.123\ \text{g}}{13.4\ \text{mL}} = 2.70\ \text{g/mL}$$

C. Addition and subtraction

Keep only as many digits after the decimal point as there are in the least precise quantity. Add 1.223 g of sugar to 154.5 g of coffee:

$$\text{total mass} = 1.2\ \text{g} + 154.5\ \text{g} = 155.7\ \text{g}$$

Note that the rule for addition and subtraction does not apply to significant figures. The number of significant figures may well decrease after subtraction.

mass beaker + sample =	52.169 g	(five significant figures)
mass empty beaker =	52.120 g	(five significant figures)
mass sample =	0.049 g	(two significant figures)

D. Exact numbers

One liter means 1.000000... L.

IV. Conversion Factors

A. One-step conversion

A rainbow trout is measured to be 16.2 inches long. What is its length in centimeters?

$$\text{length in cm} = 16.2\ \text{in} \times \frac{2.54\ \text{cm}}{1\ \text{in}} = 41.1\ \text{cm}$$

Note the cancellation of units. To convert from centimeters to inches, use the conversion factor 1 in = 2.54 cm. (Here, there are *exactly* 2.54 cm in one inch.)

B. Multiple conversion factors

A thrown baseball has speed 89.6 miles per hour. What is its speed in meters per second?

$$1\ \text{mile} = 1.609\ \text{km} = 1.609 \times 10^3\ \text{m}; \quad 1\ \text{h} = 3600\ \text{s}$$

$$\text{speed} = \left(89.6\ \frac{\text{mile}}{\text{h}}\right) \times \left(1.609 \times 10^3\ \frac{\text{m}}{\text{mile}}\right) \times \left(\frac{1\ \text{h}}{3600\ \text{s}}\right) = 40.0\ \text{m/s}$$

V. Properties of Substances

Distinguish between intensive and extensive, and between chemical and physical.

A. Density

An empty flask weighs 22.138 g. Pipet 5.00 mL of octane into the flask, producing a total mass of 25.598 g. What volume is occupied by ten grams of octane?

$$d = 3.460\ \text{g}/5.00\ \text{mL} = 0.692\ \text{g/mL}$$

$$V = 10.00\ \text{g} \times \frac{1\ \text{mL}}{0.692\ \text{g}} = 14.5\ \text{mL}$$

Note that for density calculations, 1 mL = 1 cm^3.

B. Solubility

This is often expressed as grams of solute per 100 g of solvent.

Solubility of sugar at 20°C = 210 g sugar/100 g water.

A solution containing 210 g sugar/100 g water is saturated.

A solution containing less than 210 g sugar/100 g water is unsaturated.

A solution containing more than 210 g sugar/100 g water is supersaturated.

1. How much water is required to dissolve 52 g of sugar at 20°C?

$$52\ \text{g sugar} \times \frac{100\ \text{g water}}{210\ \text{g sugar}} = 25\ \text{g water}$$

2. A solution at 20°C contains 25 g sugar and 125 g water. Is it unsaturated, saturated or supersaturated?

$$\text{mass sugar/100 g water} = \frac{25\text{ g sugar}}{125\text{ g water}} \times 100\text{ g water} = 20\text{ g sugar} \qquad \text{(unsaturated)}$$

DEMONSTRATIONS

1. Scientific method: GILB H 29
2. Decomposition of mercury(II) oxide: GILB A 8
3. Separation of a mixture: GILB A 14
4. Reaction of sodium with chlorine: GILB A 24, A 25; SHAK 1 61; J. Chem. Educ. 73 539 (1996)
5. Chromatography: GILB Q 3, Q 13
6. Significant figures: J. Chem. Educ. 69 497 (1992)
7. Density of liquids: GILB C 13; SHAK 3 229
8. Supersaturation: GILB F 11; SHAK 1 27

SUMMARY PROBLEM

(a) K, Mn, O

(b) density, melting point, solubility, color

(c) $\text{mass} = 2.703\ \frac{\text{g}}{\text{cm}^3} \times 48.7\text{cm}^3 = 132\text{ g}$

(d) $2.703\ \frac{\text{g}}{\text{cm}^3} \times \frac{1\text{ lb}}{454\text{ g}} \times \frac{(2.54)^3\text{ cm}^3}{1^3\text{ in}^3} \times \frac{(12)^3\text{ in}^3}{1\text{ ft}} = 169\text{ lb/ft}^3$

(e) $°\text{F} = \frac{9}{5}(°\text{C}) + 32 = \frac{9}{5}(2.40 \times 10^2) + 32 = 464°\text{F}$

$(2.40 \times 10^2) + 273 = 513\text{ K}$

(f) $38.5\text{ g } H_2O \times \frac{6.38\text{ g } KMnO_4}{100\text{ g } H_2O} = 2.46\text{ g } KMnO_4$

(g) At 60°C: $65.0\text{ g } H_2O \times \frac{25\text{ g } KMnO_4}{100\text{ g } H_2O} = 16\text{ g } KMnO_4$

The solution is unsaturated.

At 20°C: $65.0\text{ g } H_2O \times \frac{6.38\text{ g } KMnO_4}{100\text{ g } H_2O} = 4.15\text{ g } KMnO_4$ can be dissolved.

The solution is supersaturated.

(h) $55.0 \text{ g } H_2O \times \frac{25 \text{ g } KMnO_4}{100 \text{ g } H_2O} = 14 \text{ g } KMnO_4$ can be dissolved at 60°C.

Yes, all the $KMnO_4$ added will dissolve.

$55.0 \text{ g } H_2O \times \frac{6.38 \text{ g } KMnO_4}{100 \text{ g } H_2O} = 3.51 \text{ g } KMnO_4$ can be dissolved at 20°C.

No, not all the $KMnO_4$ will dissolve.

10.0 g - 3.51 g = 6.5 g $KMnO_4$ will remain undissolved.

PROBLEMS

1. (a) mixture (b) element (c) mixture (d) compound

3. (a) solution (b) solution (c) heterogeneous mixture

5. (a) distillation (b) filtration (c) gas chromatography

7. (a) Ti (b) P (c) K (d) Mg

9. (a) mercury (b) silicon (c) sodium (d) iodine

11. (a) balance (b) thermometer (c) graduated cylinder

13. $t_{°F} = 1.8(52°) + 32° = 126°F$; $t_K = 52 + 273.15 = 325 \text{ K}$

15. $t_{°C} = (85.0 - 32) \times \frac{5}{9} = 29.4°C$; solid

17. (a) 3 (b) ambiguous (c) 4
(d) exact (e) 5

19. (a) 7.49 g (b) 298.69 cm (c) 1×10^1 lb (d) 12.0 oz

21. (a) 1.325×10^2 cm (b) 8.83×10^{-4} km (c) 6.432×10^9 nm

23. (c)

25. 10,000: ambiguous $1.71 \times 10^5 \text{ ft}^2$: 3 \$22.00: exact 20%: ambiguous

27. (a) 80.0 (b) 0.7615 (c) 14.712
(d) 0.03 (e) 1.5×10^{-22}

29. $\dfrac{4\pi(4.30\text{ cm})^3}{3} = 333\text{ cm}^3$; $\dfrac{4\pi(4.33\text{ cm})^3}{3} = 3.40 \times 10^2\text{ cm}^3$; 7 cm^3

31. (a) $303\text{ m} = 0.303\text{ km} < 303 \times 10^3\text{ km}$ (b) 500 g = 0.500 kg

(c) $1.50\text{ cm}^3 = 1.50 \times 10^{21}\text{ nm}^3 > 1.50 \times 10^3\text{ nm}^3$

33. (a) $22.3\text{ mL} \times \dfrac{1\text{ L}}{10^3\text{ mL}} = 2.23 \times 10^{-2}\text{ L}$ (b) $22.3\text{ cm}^3 \times \dfrac{1\text{ in}^3}{(2.54\text{ cm})^3} = 1.36\text{ in}^3$

(c) $22.3\text{ mL} \times \dfrac{1\text{ L}}{10^3\text{ mL}} \times \dfrac{1.057\text{ qt}}{1\text{ L}} = 0.0236\text{ qt}$

35. (a) $19.2\text{ hands} \times \dfrac{\frac{1}{3}\text{ ft}}{1\text{ hand}} = 6.40\text{ ft}$

(b) $17.8\text{ hands} \times \dfrac{\frac{1}{3}\text{ ft}}{1\text{ hand}} \times \dfrac{12\text{ in}}{1\text{ ft}} \times \dfrac{1\text{ m}}{39.37\text{ in}} = 1.81\text{ m}$

(c) $20.5\text{ hands} \times \dfrac{\frac{1}{3}\text{ ft}}{1\text{ hand}} + 3.0\text{ ft} = 9.8\text{ ft}$

37. $2.0\text{ acre} \times \dfrac{4.356 \times 10^4\text{ ft}^2}{1\text{ acre}} \times \dfrac{(12)^2\text{ in}^2}{1\text{ ft}^2} \times \dfrac{1\text{ m}^2}{(39.37)^2\text{ in}^2} \times \dfrac{1\text{ hectare}}{10^4\text{ m}^2} = 0.81\text{ hectare}$

39. $\dfrac{5.0\text{ mi}}{1\text{ mile}} \times \dfrac{0.25\text{ min}}{1\text{ Eng lap}} = 1.2\text{ min}$

$0.50\text{ km} \times \dfrac{1\text{ mi}}{1.609\text{ km}} \times \dfrac{5.0\text{ min}}{1\text{ mi}} = 1.6\text{ min}$

41. (a) $3.0\text{ qt plasma} \times \dfrac{1\text{ L}}{1.057\text{ qt}} \times \dfrac{1000\text{ mL}}{1\text{ L}} \times \dfrac{0.080\text{ mL alcohol}}{100\text{ mL plasma}} = 2.3\text{ mL}$

(b) $3.0\text{ qt plasma} \times \dfrac{1\text{ L}}{1.057\text{ qt}} \times \dfrac{1000\text{ mL}}{1\text{ L}} \times \dfrac{0.10\text{ mL alcohol}}{100\text{ mL plasma}} = 2.8\text{ mL}$

(c) $2.8\text{ mL} - 2.3\text{ mL} = 0.5\text{ mL}$

43. $\dfrac{235\text{ kJ}}{250\text{ mL}} \times \dfrac{10^3\text{ J}}{1\text{ kJ}} \times \dfrac{1\text{ cal}}{4.18\text{ J}} \times \dfrac{1\text{ kcal}}{10^3\text{ cal}} \times \dfrac{1000\text{ mL}}{1.057\text{ qt}} \times \dfrac{1\text{ qt}}{4\text{ cups}} = 53.2\text{ kcal/cup}$

45. $\dfrac{252\text{ g}}{0.750 \times 225\text{ mL}} = 1.49\text{ g/mL}$

47. $V_{methanol} = 43.7\text{ g} \times \dfrac{1\text{ mL}}{0.791\text{ g}} = 55.2\text{ mL}$, $V_{slug} = 59.7\text{ mL} - 55.2\text{ mL} = 4.5\text{ mL}$

Therefore, $d_{slug} = \dfrac{25.17\text{ g}}{4.5\text{ mL}} = 5.6\text{ g/mL}$

49. $(8.0 \times 7.0 \times 0.75)\text{ ft}^3 \times \dfrac{(12)^3\text{ in}^3}{1\text{ ft}^3} \times \dfrac{(2.54)^3\text{ cm}^3}{1\text{ in}^3} \times \dfrac{1.00\text{ g}}{1\text{ cm}^3} \times \dfrac{1\text{ kg}}{1000\text{ g}} = 1.2 \times 10^3\text{ kg}$

51. Volume of air = volume of room

$V = 55\text{ kg } O_2 \times \dfrac{1000\text{ g}}{1\text{ kg}} \times \dfrac{1\text{ L}}{1.31\text{ g}} \times \dfrac{100\text{ L}}{21\text{ L}} = 2.0 \times 10^5\text{ L}$

53. At 30°C: maximum amount of $MgSO_4$ that can be dissolved = $25.0\text{ g } H_2O \times \dfrac{38.9\text{ g } MgSO_4}{100\text{ g } H_2O} = 9.72\text{ g}$

The solution is unsaturated.

(9.50 + 1.00 g) - 9.72 = 0.78 g will precipitate out

55. (a) $46\text{ g } H_2O \times \dfrac{16\text{ g } NaHCO_3}{100\text{ g } H_2O} = 7.4\text{ g } NaHCO_3$ can be dissolved

9.2 g in the mixture > 7.4 g, thus the solution is not homogeneous.

9.2 g – 7.4 g = 1.8 g $NaHCO_3$ are undissolved.

(b) $9.2\text{ g } NaHCO_3 \times \dfrac{100\text{ g } H_2O}{9.6\text{ g } NaHCO_3} = 96\text{ g } H_2O$ needed to dissolve

$96\text{ g} - 46\text{ g} = 5.0 \times 10^1\text{ g } H_2O$ needs to be added.

57. 57.0 g − 25.0 g = 32.0 g of $Pb(NO_3)_2$ dissolves in 64.0 g H_2O at 10°C. Solubility is

$$\frac{32.0\text{ g } Pb(NO_3)_2}{64.0\text{ g } H_2O} = \frac{1.00\text{ g } Pb(NO_3)_2}{2.00\text{ g } H_2O} = \frac{50.0\text{ g } Pb(NO_3)_2}{100.0\text{ g } H_2O}$$

59. (a) physical (b) physical (c) physical (d) chemical

61. $V = 35\text{ ft} \times 43\text{ ft} \times 28\text{ in} \times \dfrac{1\text{ ft}}{12\text{ in}} = 3.5 \times 10^3\text{ ft}^3$

$3.5 \times 10^3\text{ ft}^3 \times \dfrac{(12\text{ in})^3}{(1\text{ ft})^3} \times \dfrac{(2.54\text{ cm})^3}{(1\text{ in})^3} = 9.9 \times 10^7\text{ cm}^3$

$\text{mass} = 9.9 \times 10^7\text{ cm}^3 \times \dfrac{(0.35\text{ g})}{(1\text{ cm})^3} \times \dfrac{1\text{ lb}}{453.6\text{ g}} = 7.7 \times 10^4\text{ lbs}$

63. $108 \text{ carats} \times \dfrac{0.200 \text{ g}}{1 \text{ carat}} \times \dfrac{1 \text{ lb}}{454 \text{ g}} = 0.0476 \text{ lb}$

$0.0476 \text{ lb} \times \dfrac{1 \text{ cm}^3}{3.51 \text{ g}} \times \dfrac{454 \text{ g}}{1 \text{ lb}} \times \dfrac{1 \text{ in}^3}{(2.54)^3 \text{ cm}^3} = 0.376 \text{ in}^3$

65. $153.2 \text{ g} \times \dfrac{1 \text{ cm}^3}{4.55 \text{ g}} = 33.7 \text{ cm}^3 = V$; $33.7 = \pi r^2(7.75)$; $r = 1.18$ cm; $d = 2.35$ cm

67. (a) Chemical properties show the behavior of the species in a reaction; physical properties are intrinsic qualities.

(b) Distillation vaporizes the liquid; filtration removes the solid.

(c) The solute is a component of the solution.

69. The bottom layer is Hg; the middle layer is Pb; the top layer is ethyl alcohol.

71. (a) $\approx$ 115 g; supersaturated (b) $\approx$ 30 g; unsaturated

(c) Dissolve 30 g of compound in 100 g H_2O.

73. (a) See Figure in the answers to the problems in Appendix 5.

(b) $\dfrac{\Delta y}{\Delta x} = \dfrac{100°J - 0°J}{78°C - (-117.0°C)} = 0.512$

(c) 60°J

(d) °J = 0.51(°C) + 60

74. $31.5 \text{ gal} \times \dfrac{4 \text{ qt}}{1 \text{ gal}} \times \dfrac{1 \text{ L}}{1.057 \text{ qt}} \times \dfrac{10^{-3} \text{ m}^3}{1 \text{ L}} \times \dfrac{1 \text{ km}^3}{10^9 \text{ m}^3} = 1.2 \times 10^{-10} \text{ km}^3$

$\text{area} = \dfrac{1.2 \times 10^{-10} \text{ km}^3}{100 \text{ nm}} \times \dfrac{1 \times 10^{12} \text{ nm}}{1 \text{ km}} = 1.2 \text{ km}^2$

75. $V = 12.0 \text{ g} \times \dfrac{1 \text{ cm}^3}{2.70 \text{ g}} = \pi(0.254 \text{ cm})^2 \ell$; $\ell = 21.9$ cm

76. $\dfrac{8.50 \times 10^3 \text{ L}}{1 \text{ d}} \times \dfrac{1 \text{ m}^3}{10^3 \text{ L}} \times \dfrac{7.0 \times 10^{-6} \text{ g Pb}}{1 \text{ m}^3} \times 0.75 \times 0.50 \times \dfrac{365 \text{ d}}{1 \text{ yr}} = 8.1 \times 10^{-3} \text{ g Pb}$

78. mass of Hg in cylinder B = 145.20 g

mass of Hg + metal in cylinder A = 92.60 g

mass of metal = 145.20 g − 92.60 g = 52.60 g

volume of cylinder A = volume of cylinder B = volume of Hg in cylinder B
= 145.2 g ÷ 13.6 g/mL = 10.7 mL

mass of Hg in cylinder A = 92.60 g – 52.60 g = 40.0 g

volume of Hg in cylinder A = 40.0 g ÷ 13.6 g/mL = 2.94 mL

volume of metal = volume of cylinder A - volume of Hg in cylinder A = 10.7 mL − 2.94 mL = 7.76 mL

density of metal = 52.60 g/7.76 mL = 6.78 g/mL

2 ATOMS, MOLECULES AND IONS

LECTURE NOTES

Students find this material relatively easy to assimilate; it's almost entirely qualitative. On the other hand, there's a lot of memorizing (sorry, learning) to do. This chapter is coverable in two lectures.

Some general observations:

- Material in Sections 2.1–2.3 is generally well covered in high-school chemistry courses; no need to dwell on it.
- Students need to know the molecular formulas of the elements (Figure 2.13), the charges of ions with noble-gas structures and the names and formulas of the common polyatomic ions (Table 2.2). The charges of transition-metal ions will be covered later, in Chapter 4.
- Naming compounds requires students to distinguish between ionic and molecular substances. It helps to point out that binary molecular compounds are composed of two nonmetals. Almost all ionic compounds contain a metal cation combined with a nonmetal anion or negatively charged polyatomic ion. The flow charts shown in Figures 2.18 and 2.19 should help visual learners.
- The periodic table will be discussed in greater detail later in the text (Chapter 6).

Lecture 1

I. Atomic Theory

A. Elements

Postulates: Elements consist of tiny particles called atoms, which retain their identity in reactions. In a compound, atoms of two or more elements combine in a fixed ratio of small whole numbers (e.g., 1:1, 2:1, etc.).

B. Components

	relative mass	relative charge	location
proton	1	+1	nucleus
neutron	1	0	nucleus
electron	0.0005	−1	outside

C. Atomic number

It is the number of protons in the nucleus or the number of electrons in a neutral atom. This is characteristic of a particular element: all H atoms have one proton, all He atoms have two protons, etc.

D. Mass number

1. It is the sum of the number of protons and the number of neutrons. Atoms of the same element can differ in mass number. Those are referred to as isotopes. For example:

	protons	neutrons	atomic no.	nuclear symbol	mass no.
carbon-12	6	6	6	$^{12}_{6}C$	12
carbon-14	6	8	6	$^{14}_{6}C$	14

2. Isotopes

Atoms of the same element (same atomic number) but differ in mass number.

II. Atomic Masses

A. Meaning of atomic masses

They give the relative masses of atoms. Based on the C-12 scale; the most common isotope of carbon is assigned an atomic mass of exactly 12 amu.

element	B	Ca	Ni
atomic mass (amu)	10.81	40.08	58.69

A nickel atom is 58.69/40.08 times as heavy as a calcium atom. It is 58.69/10.81 = 5.429 times as heavy as a boron atom.

B. Atomic masses from isotopic composition

atomic mass = (atomic mass of isotope 1)(%/100) + (atomic mass of isotope 2)(%/100) + · · ·

Isotope	Atomic mass	Percent
Ne-20	20.00 amu	90.92
Ne-21	21.00 amu	0.26
Ne-22	22.00 amu	8.82

atomic mass of Ne = (20.00)(0.9092) + (21.00)(0.0026) + (22.00)(0.0882) = 20.18 amu

C. Masses of individual atoms

Since the atomic masses of H, Cl and Ni are, respectively, 1.008 amu, 35.45 amu and 58.69 amu, it follows that

$$1.008\,g\ H,\ 35.45\,g\ Cl,\ 58.69\,g\ Ni \text{ all contain the same number of atoms, } N_A.$$

$$N_A = \text{Avogadro's number} = 6.022 \times 10^{23}$$

1. Mass of a hydrogen atom?

$$1 \text{ atom H} \times \frac{1.008 \text{ g H}}{6.022 \times 10^{23} \text{ atom}} = 1.674 \times 10^{-24} \text{ g}$$

2. Number of atoms in one gram of nickel?

$$1.000 \text{ g Ni} \times \frac{6.022 \times 10^{23} \text{ atoms Ni}}{58.69 \text{ g Ni}} = 1.026 \times 10^{22} \text{ atoms}$$

III. Periodic Table

Periods and groups; numbering system for groups. Metals appear at the lower left, nonmetals at the upper right. Metalloids.

Lecture 2

IV. Molecules

A. Composition

Usually consist of nonmetal atoms; held together by covalent bonds.

B. Types of Formulas

Consider the compound ethyl alcohol:

Molecular formula: C_2H_6O

Structural formula:
$$\begin{array}{ccccccccc} & & H & & H & & & & \\ & & | & & | & & & & \\ H & - & C & - & C & - & O & - & H \\ & & | & & | & & & & \\ & & H & & H & & & & \end{array}$$

Condensed structural formula: CH_3CH_2OH

V. Ions

A. Formation of monatomic ions

Na atom $(11p^+, 11e^-) \rightarrow$ Na^+ ion $(11p^+, 10e^-) + e^-$

F atom $(9p^+, 9e^-) + e^- \rightarrow$ F^- ion $(9p^+, 10e^-)$

B. Charges of monatomic ions with noble-gas structures

Cations: Group 1 (+1); Group 2 (+2); Al^{3+}

Anions: Group 16 (−2); Group 17 (−1); N^{3-}

C. Polyatomic ions

Names and formulas (Table 2.2)

D. Formulas of compounds

Apply the principle of electroneutrality.

calcium fluoride:	Ca^{2+}, F^- ions:	CaF_2
aluminum nitrate:	Al^{3+}, NO_3^- ions:	$Al(NO_3)_3$
sodium dihydrogen phosphate:	Na^+, $H_2PO_4^-$ ions:	NaH_2PO_4

E. Ionic compounds

They can be distinguished from molecular substances by the conductivity of their water solutions. Solutions of NaCl, $Ca(OH)_2$, ... conduct electricity (electrolytes). Sugar is a nonelectrolyte.

VI. Names of Compounds

A. Ionic

Name cation, followed by anion. Note that with transition metal cations, charge is indicated by a Roman numeral.

Na_2SO_4	sodium sulfate	$Fe(NO_3)_3$	iron(III) nitrate

Systematic names of oxoanions (-ate, -ite, per-, hypo-)

Calcium hypochlorite $Ca(ClO)_2$

B. Binary molecular compounds

Use of Greek prefixes:	SF_6	sulfur hexafluoride
	N_2O_3	dinitrogen trioxide

C. Acids

Binary acids:	hydrochloric acid	
Oxo acids:	-ate salt → -ic acid	$HClO_4$, perchloric acid
	-ite salt → -ous acid	HClO, hypochlorous acid

DEMONSTRATIONS

1. Law of constant composition: GILB A 12
2. Law of conservation of mass: GILB A 16
3. Simulation of Rutherford's experiment: GILB L 7
4. Isotope effects (H_2O, D_2O): GILB M 18
5. Reaction of hydrogen with chlorine: GILB H 38
6. Conductivity of water solutions: SHAK 3 140
7. Breath alcohol detection: J. Chem. Educ. *67* 263 (1990); *71* 158 (1994)
8. Relative masses of atoms (analogy): GILB L 2

SUMMARY PROBLEM

(a) S_8 (b) 16 protons, 16 electrons (c) no; Al_2S_3 - aluminum sulfide

(d) yes (e) yes; S_2Cl_2 - disulfur dichloride (f) ${}^{34}_{16}S$

(g) group 1**6**, period 3 (h) 20 neutrons

(i) (31.97207)(0.9493) + (32.97146)(0.0076) + (33.96787)(0.0429) + (35.96708)(0.0002) = 32.07 amu

(j) $12.55\text{ g S} \times \dfrac{1\text{ mol S}}{32.07\text{ g}} \times \dfrac{6.022 \times 10^{23}\text{ atoms}}{1\text{ mol S}} = 2.357 \times 10^{23}\text{ atoms}$

(k) $1 \times 10^{9}\text{ S atoms} \times \dfrac{1\text{ mol S}}{6.022 \times 10^{23}\text{ atoms}} \times \dfrac{32.07\text{ g}}{1\text{ mol S}} = 5.325 \times 10^{-14}\text{ g}$

(l) SO_3 = sulfur trioxide; H_2SO_3 (aq) = sulfurous acid; SO_4^{2-} = sulfate ion; Na_2SO_3 = sodium sulfite

PROBLEMS

1. p. 29

3. (a) Conservation of mass (b) Constant composition (c) neither

5. J. J. Thompson; see p. 29

7. ${}^{80}_{34}Se$

9. no. of neutrons: ${}^{36}_{18}Ar$, ${}^{38}_{18}Ar$, ${}^{40}_{18}Ar$

11. (a) 92 (b) 143 (c) 92

13. (a) 14 p^+, 16 n, 14 e^-; R = Si

 (b) 39 p^+, 50 n, 39 e^-; T = Y

 (c) 55 p^+, 78 n, 55 e^-; X = Cs

15. (a) Ca-41, K-41, Ar-41 are isobars; Ca-40, Ca-41 are isotopes

 (b) atomic number = number of protons = 20

 (c) same mass number

17. (a) $\dfrac{79.90}{20.18} = 3.959$ (b) $\dfrac{79.90}{40.08} = 1.994$ (c) $\dfrac{79.90}{4.003} = 19.96$

19. Ce-140

21. 50%

23. 83.9134(0.0056) + 85.9094(0.0986) + 86.9089(0.0700) + 87.9056(0.8258) = 0.47 + 8.47 + 6.08 + 72.59

 average atomic mass = 87.61

25. 107.9 = 106.90509(0.5184) + 0.4816 x; x = 109 amu

27. Let x = abundance of the first isotope; abundance of second isotope = 0.9704 $-$ x

 $28.0855 = 27.9769\,x + (0.9704 - x)(28.9765) + (0.0296)(29.9738)$

 $= 27.9769\,x + 28.1188 - 28.9765\,x + 0.887$

 x = 0.921; abundance of first isotope is 92.1%

 $0.9704 - x = 0.9704 - 0.921 = 0.0494$; abundance of second isotope is 4.94%

29. Tall peak at mass 64; peak a little over 1/2 as high at mass 66; smallest peak is at mass 67, and the height of the peak at mass 64 is 2.5 times that of the peak at mass 68 .

31. $3 \times 10^{-7}\text{ g} \times \dfrac{1\text{ mol}}{207.2\text{ g}} \times \dfrac{6.022 \times 10^{23}\text{ atoms}}{1\text{ mol}} = 9 \times 10^{14}\text{ atoms}$

33. (a) $0.185\text{ g Pd} \times \dfrac{6.022 \times 10^{23}\text{ atoms}}{106.4\text{ g Pd}} = 1.05 \times 10^{21}\text{ atoms}$

 (b) $127\text{ protons} \times \dfrac{1\text{ atom}}{46\text{ protons}} \times \dfrac{106.4\text{ g}}{6.022 \times 10^{23}\text{ atoms}} = 4.88 \times 10^{-22}\text{ g}$

35. (a) $0.35744\,\text{mol} \times \dfrac{6.022 \times 10^{23}\ \text{atoms}}{1\ \text{mol}} = 2.152 \times 10^{23}\ \text{atoms}$

(b) $2.152 \times 10^{23}\ \text{atoms} \times \dfrac{(14\,p^{+} + 14\,n + 14\,e^{-})}{1\ \text{atom}} = 9.039 \times 10^{24}$

37. $V_{\text{cube}} = \left(1.25\ \text{in} \times \dfrac{2.54\ \text{cm}}{1\ \text{in}}\right)^3 = 32.0\ \text{cm}^3$

$32.0\,\text{cm}^3 \times \dfrac{0.968\,\text{g}}{1\ \text{cm}^3} \times \dfrac{6.022 \times 10^{23}\ \text{atoms}}{22.99\,\text{g}} = 8.12 \times 10^{23}\ \text{atoms}$

39. (a) K (b) Cd (c) Al (d) Sb (e) P

41. (a) main-group metal (b) transition metal (c) main-group metal (d) metalloid

(e) nonmetal

43. (a) 6 (b) 4 named, 1 not named (c) 0

45. (a) 13 (b) 2 (c) 17, 18

47. (a) C_2H_7N (b) C_3H_8O

49. (a) 14 p^{+}, 14e^{-} (b) 21 p^{+}, 22 e^{-} (c) 35 p^{+}, 34e^{-} (d) 70 p^{+}, 70 e^{-}

51.

$^{19}_{9}F$	0	9	10	9
$^{31}_{15}P$	0	15	16	15
$^{57}_{27}Co^{3+}$	+3	27	30	24
$^{32}_{16}S^{2-}$	−2	16	16	18

53. (a) electrolyte (b) nonelectrolyte (c) nonelectrolyte (d) electrolyte

55. (a) CH_4 (b) CI_4 (c) H_2O_2 (d) NO (e) SiO_2

57. (a) iodine trichloride (b) dinitrogen pentaoxide (c) phosphine

(d) carbon tetrabromide (e) sulfur trioxide

59. KCl, K_2S, $CaCl_2$, CaS

61. (a) $Fe(C_2H_3O_2)_3$ (b) $Ca(NO_3)_2$ (c) K_2O (d) $AuCl_3$ (e) Ba_3N_2

63. (a) potassium dichromate (b) copper(II) phosphate (c) barium acetate

    (d) aluminum nitride (e) cobalt(II) nitrate

65. (a) hydrochloric acid (b) chloric acid (c) iron(III) sulfite

    (d) barium nitrite (e) sodium hypochlorite

67. HNO_2, nickel(II) iodate, Au_2S_3, sulfurous acid, NF_3

69. (a) $Mn(NO_2)_3$; manganese(III) nitrite

    (b) BF_3; boron trifluoride

    (c) $Ca(HCO_3)_2$; calcium hydrogen carbonate

71. (a) In (b) Pb or Sn (c) K (d) Sb

73. (a) ... confirmed the presence of a dense nucleus with protons.

    (b) ... elements arranged according to increasing atomic number.

    (c) ... same number of protons.

    (d) Be_3N_2 is beryllium nitride.

75. $$6.00 \text{ oz salami} \times \frac{1 \text{ g}}{0.03527 \text{ oz}} \times \frac{0.090 \text{ g } NaC_7H_5O_2}{100 \text{ g salami}} \times \frac{6.022 \times 10^{23} \text{ molecules } NaC_7H_5O_2}{144.1 \text{ g } NaC_7H_5O_2}$$
$$\times \frac{1 \text{ atom Na}}{1 \text{ molecule } NaC_7H_5O_2} = 6.4 \times 10^{20} \text{ Na atoms}$$

77. (b), (d), (e)

79. 8 □○ molecules; 3 □□ molecules left

81. A square with four circles around it (several of them in a flask with a defined volume)

83. (a) 118 (b) 120 (c) 117 (d) 120 (e) 119

85. first experiment: $\%\,O = \dfrac{3.87}{52.30} \times 100 = 7.40$; $\%\,Hg = 92.60$

second experiment: $\%\,Hg = \dfrac{15.68}{16.93} \times 100 = 92.62$; $\%\,O = 7.38$

87. (a) K, Sr (b) O, F, Ar, S (c) S, K, Sr

(d) S (e) S, O or S, F or O, F (f) Sr, S or Sr, O or K, F

(g) Sr, F (h) K, O or K, S (i) Ar

(j) O, F, Ar

88. A: mass C/mass H = 11.9 ($\approx$ 12) B: mass C/mass H = 2.99 ($\approx$ 3)

ratio for (a) = 2.77 ratio for (b) = 4.67 ratio for (c) = 5.96

(c) is best choice

89. (a) ethane: 18.0 g C/4.53 g H = 3.97 g C/g H ethylene: 43.20 g C/7.25 g H = 5.96 g C/g H

5.96/3.97 = 1.50 = 3/2

(b) CH_2 and CH_3; C_2H_4 and C_2H_6

90. mass $= 13\left(1.6726 \times 10^{-4}\,g\right) + 13\left(9.1094 \times 10^{-28}\,g\right) + 14\left(1.6749 \times 10^{-24}\,g\right) = 4.5204 \times 10^{-23}\,g$

$V = \frac{4}{3}\pi\left(1.43 \times 10^{-8}\,cm\right)^3 = 1.22 \times 10^{-23}\,cm^3$

$d = 4.5204\,g/1.22\,cm^3 = 3.71\,g/cm^3$

Empty space between Al atoms.

91. $2.3440 \times 10^{-23}\,g \;+\; 3(9.1095 \times 10^{-28}\,g) \;=\; 2.3443 \times 10^{-23}\,g$

92. (a) $200 \text{ inhalations} \times \dfrac{500 \text{ mL}}{1 \text{ inhalation}} \times \dfrac{2.5 \times 10^{19} \text{ molecules}}{1 \text{ mL}} = 2.5 \times 10^{24} \text{ molecules}$

(b) $\dfrac{2.5 \times 10^{24}}{1.1 \times 10^{44}} = 2.3 \times 10^{-20}$

(c) $1 \text{ inhalation} \times \dfrac{500 \text{ mL}}{1 \text{ inhalation}} \times \dfrac{2.5 \times 10^{19} \text{ molecules}}{1 \text{ mL}} \times 2.3 \times 10^{-20} = 2.8 \times 10^{2} \text{ molecules}$

93. Total mass before reaction: 18.00 g + (25.00 x 1.025 g/mL) = 43.63 g

    After reaction following the law of conservation of mass: 43.63 g = 12 g + 30.95 g + mass of H_2

    mass of H_2 = 0.68 g; volume of H_2 = 0.68 g x $\frac{1\ \text{L}}{0.0824\ \text{g}}$ = 8.25 L

3 MASS RELATIONS IN CHEMISTRY; STOICHIOMETRY

LECTURE NOTES

This chapter is considerably more difficult and time-consuming than the two preceding ones. It contains a good deal of quantitative material that is fundamental for future chapters. We suggest that you devote three lectures to Chapter 3. The first lecture deals with the mole and the mole in solutions (molarity), the second with the quantitative aspects of chemical formulas (Section 3.2), and the third with mass relations in reactions (Section 3.3). Points to keep in mind include:

1. Note that mole-gram conversions and molarity calculations (Section 3.1) will be required in many later chapters, often as the first step in a more complex problem.
2. When dealing with formulas (Section 3.2), it is important to emphasize early on that the subscripts give not only the atom ratio but also the mole ratio. Students must realize this in order to follow the logic of obtaining simplest formulas from mass percents.
3. Students ordinarily have little trouble calculating formulas from mass percents. They are much less adept at obtaining formulas from analytical data such as that in Example 3.6.
4. It is important to get across the point (Section 3.3) that a chemical equation describes what happens when a reaction is carried out in the laboratory. Including the physical states of reactants and products in the equation helps to emphasize this point.
5. When you discuss mass relations in reactions, some students will revert to the infamous "ratio and proportion" method. The comments of Chapter 1 about conversion factors apply here, too.
6. There are many ways to find the limiting reactant and calculate the theoretical yield. We've tried most of them, and recommend the approach described in Section 3.3; students seem to grasp it readily.

Lecture 1

I. The Mole

A. Meaning

1 mol = 6.022×10^{23} items

1 mol H = 6.022×10^{23} H atoms; mass = 1.008 g

1 mol Cl = 6.022×10^{23} Cl atoms; mass = 35.45 g

1 mol Cl_2 = 6.022×10^{23} Cl_2 molecules; mass = 70.90 g

1 mol HCl = 6.022×10^{23} HCl molecules; mass = 36.46 g

B. Molar mass

Generalizing from the above examples, the molar mass, MM, is numerically equal to the sum of the atomic masses.