PARTIAL SOLUTIONS GUIDE

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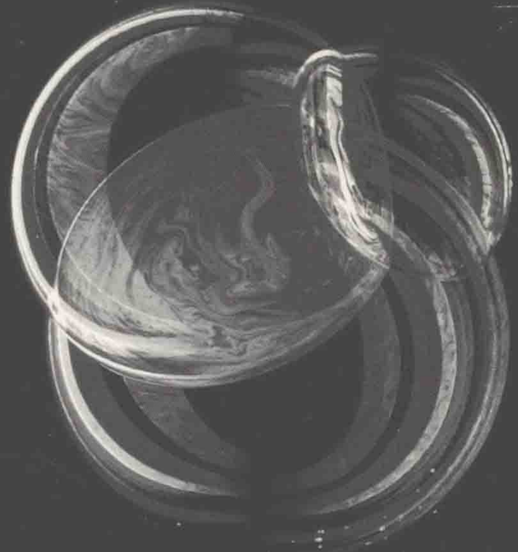

C H E M I S T R Y

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Partial Solutions Guide

Chemistry

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TO THE STUDENT: HOW TO USE THIS GUIDE

Solutions to all of the odd numbered end of chapter questions and exercises are in this manual. This "Solutions Guide" can be very valuable if you use it properly. The way NOT to use it is to look at an exercise in the book and then immediately check the solution, often saying to yourself, "That's easy, I can do it." Chemistry is easy once you get the hang of it, but it takes practice. Don't look up a solution to a problem until you have tried to work it on your own. If you are completely stuck, see if you can find a similar problem in the Sample Exercises in the chapter. Only look up the solution as a last resort. If you do this for a problem, look for a similar problem in the end of chapter exercises and try working it. The more problems you do, the easier chemistry becomes. It is also in your self interest to try to work as many problems as possible. Most exams that you will take in chemistry will involve a lot of problem solving. If you have worked several problems similar to the ones on an exam, you will do much better than if the exam is the first time you try to solve a particular type of problem. No matter how much you read and study the text, or how well you think you understand the material, you don't really understand it until you have taken the information in the text and applied the principles to problem solving. You will make mistakes, but the good students learn from their mistakes.

In this manual we have worked problems as in the textbook. We have shown intermediate answers to the correct number of significant figures and used the rounded answer in later calculations. Thus, some of your answers may differ slightly from ours. When we have not followed this convention, we have usually noted this in the solution.

We are grateful to Delores Wyatt for her outstanding effort in preparing the manuscript of this manual. We also thank Robert Pfaff for his careful and thorough accuracy review of the Solutions Guide.

TJH
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SSZ

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CHAPTER ONE

CHEMICAL FOUNDATIONS

Questions

13. No, it is useful whenever a systematic approach of observation and hypothesis testing can be used.

15. Precision is related to how many significant figures one can associate with a measurement. Consider weighing an object on three different balances with the following results: 11 g; 11.25 g; 11.2456 g. Since the assumed uncertainty in all measurements is ± 1 in the last digit, then the uncertainty of the three balances are: ± 1 g; ± 0.01 g; ± 0.0001 g, respectively. The balance with the smallest uncertainty is the balance with the most significant figures associated with the measurement, which is also the balance that is assumed most precise.

17. Chemical changes involve the making and breaking of chemical forces (bonds). Physical changes do not. The identity of a substance changes after a chemical change, but not after a physical change.

Exercises

Significant Figures and Unit Conversions

19. a. inexact b. exact c. exact

For c, $\frac{36\text{ in}}{\text{yd}} \times \frac{2.54\text{ cm}}{\text{in}} \times \frac{1\text{ m}}{100\text{ cm}} = \frac{0.9144\text{ m}}{\text{yd}}$ (All conversion factors used are exact.)

d. inexact; Although this number appears to be exact, it probably isn't. The announced attendance may be tickets sold but not the number who were actually in the stadium. Some people who paid may not have gone, some may leave early, or arrive late, some may sneak in without paying, etc.

e. exact f. inexact

21. a. 0.00<u>12</u>; 2 S.F., 1.2×10^{-3} b. <u>437</u>,000; 3 S.F., 4.37×10^5

c. <u>900.0</u>; 4 S.F., 9.000×10^2 d. <u>106</u>; 3 S.F.; 1.06×10^2

e. <u>125,904</u>,000; 6 S.F., 1.25904×10^8 f. <u>1.0012</u>; 5 S.F., 1.0012×10^0

g. 2006; 4 S.F., 2.006×10^3 h. 3050; 3 S.F., 3.05×10^3

i. 0.001060; 4 S.F., 1.060×10^{-3}

23. a. 6×10^8 b. 5.8×10^8 c. 5.82×10^8

d. 5.8200×10^8 e. 5.820000×10^8

25. a. 467; The difference of 25.27 - 24.16 = 1.11 has only three significant figures, so the answer will only have 3 significant figures. For this problem and for subsequent problems, the addition/subtraction rule is applied separately from the multiplication/division rule.

b. 0.24; The difference of 8.925 - 8.904 = 0.021 has only 2 significant figures.

c. $(9.04 - 8.23 + 21.954 + 81.0) \div 3.1416 = 103.8 \div 3.1416 = 33.04$

We will generally round off at intermediate steps in order to show the correct number of significant figures. However, you should round off at the end of all the mathematical operations in order to avoid round off error. Make sure you keep track of the correct number of significant figures during intermediate steps, but round off at the end.

d. $$\frac{9.2 \times 100.65}{8.321 + 4.026} = \frac{9.2 \times 100.65}{12.347} = 75$$

e. $0.1654 + 2.07 - 2.114 = 0.12$

Uncertainty begins to appear in the second decimal place. Numbers were added as written and the answer was rounded off to 2 decimal places at the end. If you round to 2 decimal places and add you get 0.13. Always round off at the end of the operation to avoid round off error.

f. $8.27(4.987 - 4.962) = 8.27(0.025) = 0.21$

g. $$\frac{9.5 + 4.1 + 2.8 + 3.175}{4} = \frac{19.6}{4} = 4.90 = 4.9$$

Uncertainty appears in the first decimal place. The average of several numbers can only be as precise as the least precise number. Averages can be exceptions to the significant figure rules.

h. $$\frac{9.025 - 9.024}{9.025} \times 100 = \frac{0.001}{9.025} \times 100 = 0.01$$

27. a. $1\ \text{km} = 10^3\ \text{m} = 10^6\ \text{mm} = 10^{15}\ \text{pm}$

b. $1\ \text{g} = 10^{-3}\ \text{kg} = 10^3\ \text{mg} = 10^9\ \text{ng}$;   c. $1\ \text{mL} = 10^{-3}\ \text{L} = 10^{-3}\ \text{dm}^3 = 1\ \text{cm}^3$

d. $1\ \text{mg} = 10^{-6}\ \text{kg} = 10^{-3}\ \text{g} = 10^3\ \mu\text{g} = 10^6\ \text{ng} = 10^9\ \text{pg} = 10^{12}\ \text{fg}$

e. $1\ \text{s} = 10^3\ \text{ms} = 10^9\ \text{ns}$

29. $1\ \text{Å} \times \frac{10^{-8}\ \text{cm}}{\text{Å}} \times \frac{1\ \text{m}}{100\ \text{cm}} \times \frac{10^{9}\ \text{nm}}{\text{m}} = 1 \times 10^{-1}\ \text{nm}$

$1\ \text{Å} \times \frac{1 \times 10^{-1}\ \text{nm}}{\text{Å}} \times \frac{1\ \text{m}}{10^{9}\ \text{nm}} \times \frac{10^{12}\ \text{pm}}{\text{m}} = 1 \times 10^{2}\ \text{pm}$

31. a. Appropriate conversion factors are found on the back cover of the text. In general, the number of significant figures we use in the conversion factors will be one more than the number of significant figures from the numbers given in the problem. This is usually sufficient to avoid round off error.

$3.91\ \text{kg} \times \frac{1\ \text{lb}}{0.4536\ \text{kg}} = 8.62\ \text{lb};\quad 0.62\ \text{lb} \times \frac{16\ \text{oz}}{\text{lb}} = 9.9\ \text{oz}$

Baby's weight = 8 lb and 9.9 oz or to the nearest ounce, 8 lb and 10 oz.

$51.4\ \text{cm} \times \frac{1\ \text{in}}{2.54\ \text{cm}} = 20.2\ \text{in} \approx 20\ 1/4\ \text{in} = \text{baby's height}$

b. $25{,}000\ \text{mi} \times \frac{1.61\ \text{km}}{\text{mi}} = 4.0 \times 10^{4}\ \text{km};\ 4.0 \times 10^{4}\ \text{km} \times \frac{1000\ \text{m}}{\text{km}} = 4.0 \times 10^{7}\ \text{m}$

c. $V = 1 \times w \times h = 1.0\ \text{m} \times \left(5.6\ \text{cm} \times \frac{1\ \text{m}}{100\ \text{cm}}\right) \times \left(2.1\ \text{dm} \times \frac{1\ \text{m}}{10\ \text{dm}}\right) = 1.2 \times 10^{-2}\ \text{m}^3$

$1.2 \times 10^{-2}\ \text{m}^3 \times \left(\frac{10\ \text{dm}}{\text{m}}\right)^3 \times \frac{1\ \text{L}}{\text{dm}^3} = 12\ \text{L}$

$12\ \text{L} \times \frac{1000\ \text{cm}^3}{\text{L}} \times \left(\frac{1\ \text{in}}{2.54\ \text{cm}}\right)^3 = 730\ \text{in}^3;\ 730\ \text{in}^3 \times \left(\frac{1\ \text{ft}}{12\ \text{in}}\right)^3 = 0.42\ \text{ft}^3$

33. a. $928\ \text{mi} \times \frac{5280\ \text{ft}}{\text{mi}} \times \frac{1\ \text{fathom}}{6\ \text{ft}} \times \frac{1\ \text{cable length}}{100\ \text{fathoms}} \times \frac{1\ \text{nautical mi}}{10\ \text{cable lengths}}$

$\times \frac{1\ \text{league}}{3\ \text{nautical miles}} = 272\ \text{leagues}$

$928\ \text{mi} \times \frac{1.609\ \text{km}}{1\ \text{mi}} = 1.49 \times 10^{3}\ \text{km}$

b. $1.0\ \text{cable length} \times \frac{100\ \text{fathom}}{\text{cable length}} \times \frac{6\ \text{ft}}{\text{fathom}} \times \frac{1\ \text{yd}}{3\ \text{ft}} \times \frac{1\ \text{m}}{1.09\ \text{yd}} \times \frac{1\ \text{km}}{1000\ \text{m}} = 0.18\ \text{km}$

$1.0\ \text{cable length} = 0.18\ \text{km} \times \frac{1000\ \text{m}}{\text{km}} \times \frac{100\ \text{cm}}{\text{m}} = 1.8 \times 10^{4}\ \text{cm}$

c. $315\ \text{ft} \times \frac{12\ \text{in}}{\text{ft}} \times \frac{2.54\ \text{cm}}{\text{in}} \times \frac{1\ \text{m}}{100\ \text{cm}} = 96.0\ \text{m}$

$$37\text{ ft} \times \frac{12\text{ in}}{\text{ft}} \times \frac{2.54\text{ cm}}{\text{in}} \times \frac{1\text{ m}}{100\text{ cm}} = 11\text{ m}$$

$$315\text{ ft} \times \frac{1\text{ fathom}}{6\text{ ft}} \times \frac{1\text{ cable length}}{100\text{ fathoms}} = 5.25 \times 10^{-1}\text{ cable lengths}$$

$$37\text{ ft} \times \frac{1\text{ fathom}}{6\text{ ft}} = 6.2\text{ fathoms}$$

35. a. $1\text{ troy lb} \times \frac{12\text{ troy oz}}{\text{troy lb}} \times \frac{20\text{ pw}}{\text{troy oz}} \times \frac{24\text{ grains}}{\text{pw}} \times \frac{0.0648\text{ g}}{\text{grain}} \times \frac{1\text{ kg}}{1000\text{ g}} = 0.373\text{ kg}$

$1\text{ troy lb} = 0.373\text{ kg} \times \frac{2.205\text{ lb}}{\text{kg}} = 0.822\text{ lb}$

b. $1\text{ troy oz} \times \frac{20\text{ pw}}{\text{troy oz}} \times \frac{24\text{ grains}}{\text{pw}} \times \frac{0.0648\text{ g}}{\text{grain}} = 31.1\text{ g}$

$1\text{ troy oz} = 31.1\text{ g} \times \frac{1\text{ carat}}{0.200\text{ g}} = 156\text{ carats}$

c. $1\text{ troy lb} = 0.373\text{ kg};\ 0.373\text{ kg} \times \frac{1000\text{ g}}{\text{kg}} \times \frac{1\text{ cm}^3}{19.3\text{ g}} = 19.3\text{ cm}^3$

37. $\frac{4.4 \times 10^9\text{ mi}}{2.0\text{ yr}} \times \frac{1.61\text{ km}}{\text{mi}} \times \frac{1\text{ yr}}{365\text{ d}} \times \frac{1\text{ d}}{24\text{ hr}} \times \frac{1\text{ hr}}{60\text{ min}} \times \frac{1\text{ min}}{60\text{ s}} = 110\text{ km/s}$

39. $\frac{14\text{ km}}{\text{L}} \times \frac{1\text{ mi}}{1.61\text{ km}} \times \frac{3.79\text{ L}}{\text{gal}} = 33\text{ mi/gal}$; The spouse's car has the better gas mileage.

Temperature

41. $T_C = \frac{5}{9}(T_F - 32) = \frac{5}{9}(102.5 - 32) = 39.2°C;\ T_K = T_C + 273.2 = 312.4\text{ K}$ (Note: 32 is exact)

43. $T_F = \frac{9}{5} \times T_C + 32 = \frac{9}{5} \times 25 + 32 = 77°F;\ T_K = 25 + 273 = 298\text{ K}$

45. We can do this two ways. First, we calculate the high and low temperature and get the uncertainty from the range. 20.6°C ± 0.1°C means the temperature can range from 20.5°C to 20.7°C.

$T_F = \frac{9}{5} \times T_c + 32 \leftarrow$ (exact); $T_F = \frac{9}{5} \times 20.6 + 32 = 69.1°F$

$T_F(\text{min}) = \frac{9}{5} \times 20.5 + 32 = 68.9°F;\ T_F(\text{max}) = \frac{9}{5} \times 20.7 + 32 = 69.3°F$

So the temperature ranges from 68.9°F to 69.3°F which we can express as 69.1 ± 0.2°F.

An alternative way is to treat the uncertainty and the temperature in °C separately.

$$T_F = \frac{9}{5} \times T_C + 32 = \frac{9}{5} \times 20.6 + 32 = 69.1°F; \quad \pm 0.1°C \times \frac{9°F}{5°C} = \pm 0.18°F \approx \pm 0.2°F$$

Combining the two calculations: $T_F = 69.1 \pm 0.2°F$

Density

47. $$\frac{2.70 \text{ g}}{\text{cm}^3} \times \frac{1 \text{ kg}}{1000 \text{ g}} \times \left(\frac{100 \text{ cm}}{\text{m}}\right)^3 = \frac{2.70 \times 10^3 \text{ kg}}{\text{m}^3}$$

$$\frac{2.70 \text{ g}}{\text{cm}^3} \times \frac{1 \text{ lb}}{453.6 \text{ g}} \times \left(\frac{2.54 \text{ cm}}{\text{in}}\right)^3 \times \left(\frac{12 \text{ in}}{\text{ft}}\right)^3 = \frac{169 \text{ lb}}{\text{ft}^3}$$

49. $d = \text{density} = \frac{\text{mass}}{\text{volume}}$; mass = 1.67×10^{-24} g; radius = diameter/2 = 5.0×10^{-4} pm

$$V = \frac{4}{3}\pi r^3 = \frac{4}{3} \times 3.14 \times \left(5.0 \times 10^{-4} \text{ pm} \times \frac{1 \text{ m}}{10^{12} \text{ pm}} \times \frac{100 \text{ cm}}{\text{m}}\right)^3 = 5.2 \times 10^{-40} \text{ cm}^3$$

$$d = \frac{1.67 \times 10^{-24} \text{ g}}{5.2 \times 10^{-40} \text{ cm}^3} = \frac{3.2 \times 10^{15} \text{ g}}{\text{cm}^3}$$

51. $$5.0 \text{ carat} \times \frac{0.200 \text{ g}}{\text{carat}} \times \frac{1 \text{ cm}^3}{3.51 \text{ g}} = 0.28 \text{ cm}^3$$

53. a. Both have the same mass.

b. 1.0 mL of mercury; Mercury has a larger density than water.

$1.0 \text{ mL} \times \frac{13.6 \text{ g}}{\text{mL}} = 13.6$ g of mercury; $1.0 \text{ mL} \times \frac{0.998 \text{ g}}{\text{mL}} = 1.0$ g of water

c. Same; Both represent 19.3 g of substance.

d. 1.0 L of benzene (880 g vs 670 g)

Classification and Separation of Matter

55. Solid: own volume, own shape, does not flow; Liquid: own volume, takes shape of container, flows; Gas: takes volume and shape of container, flows

57. a. pure b. mixture c. mixture d. pure e. mixture (copper and zinc)

f. pure g. mixture h. mixture i. pure

Iron and uranium are elements. Water and table salt are compounds. Water is H_2O and table salt is NaCl. Compounds are composed of two or more elements.

Additional Exercises

59. a. 8.41 (2.16 has only three significant figures.)

b. 16.1 (Uncertainty appears in the first decimal place for 8.1.)

c. 52.5 (All numbers have 3 significant figures.)

d. 5 (2 contains one significant figure.)

e. 0.009

f. 429.59 (Uncertainty appears in 2nd decimal place for 2.17 and 4.32.)

61. $$1.5 \text{ teaspoons} \times \frac{80. \text{ mg acet}}{0.50 \text{ teaspoons}} = 240 \text{ mg acetaminophen}$$

$$\frac{240 \text{ mg acet}}{24 \text{ lb}} \times \frac{1 \text{ lb}}{0.454 \text{ kg}} = 22 \text{ mg acetaminophen/kg}$$

$$\frac{240 \text{ mg acet}}{35 \text{ lb}} \times \frac{1 \text{ lb}}{0.454 \text{ kg}} = 15 \text{ mg acetaminophen/kg}$$

The range is from 15 mg to 22 mg acetaminophen per kg of body weight.

63. $$126 \text{ gal} \times \frac{4 \text{ qt}}{\text{gal}} \times \frac{1 \text{ L}}{1.057 \text{ qt}} = 477 \text{ L}$$

65. $$\text{Total volume} = \left(200. \text{ m} \times \frac{100 \text{ cm}}{\text{m}}\right) \times \left(300. \text{ m} \times \frac{100 \text{ cm}}{\text{m}}\right) \times 4.0 \text{ cm} = 2.4 \times 10^9 \text{ cm}^3$$

$$\text{Vol. covered by 1 bag of topsoil} = \left[10 \text{ ft}^2 \times \left(\frac{12 \text{ in}}{\text{ft}}\right)^2 \times \left(\frac{2.54 \text{ cm}}{\text{in}}\right)^2\right] \times \left(1.0 \text{ in} \times \frac{2.54 \text{ cm}}{\text{in}}\right) = 2.4 \times 10^4 \text{ cm}^3$$

$$2.4 \times 10^9 \text{ cm}^3 \times \frac{1 \text{ bag}}{2.4 \times 10^4 \text{ cm}^3} = 1.0 \times 10^5 \text{ bags topsoil}$$

67. $$T_c = (68°\text{F} - 32°\text{F}) \times \frac{5°\text{C}}{9°\text{F}} = 20.°\text{C}$$

Gallium is a solid at 20.°C. From the melting point, gallium doesn't convert to the liquid state until 29.8°C.

69. $$\text{Volume of lake} = 100 \text{ mi}^2 \times \left(\frac{5280 \text{ ft}}{\text{mi}}\right)^2 \times 20 \text{ ft} = 6 \times 10^{10} \text{ ft}^3$$

$$6 \times 10^{10} \text{ ft}^3 \times \left(\frac{12 \text{ in}}{\text{ft}} \times \frac{2.54 \text{ cm}}{\text{in}}\right)^3 \times \frac{1 \text{ mL}}{\text{cm}^3} \times \frac{0.4 \text{ µg}}{\text{mL}} = 7 \times 10^{14} \text{ µg}$$

$$7 \times 10^{14}\ \mu g \times \frac{1\ g}{10^6\ \mu g} \times \frac{1\ kg}{10^3\ g} = 7 \times 10^5\ \text{kg of mercury}$$

71. $V = 5.00\ cm \times 4.00\ cm \times 2.50\ cm = 50.0\ cm^3$; $50.0\ cm^3 \times \frac{22.57\ g}{cm^3} = 1130\ g$

$$1.00\ kg \times \frac{1000\ g}{kg} \times \frac{1\ cm^3}{22.57\ g} = 44.3\ cm^3$$

73. Circumference = $c = 2\pi r$; $V = \frac{4\pi r^3}{3} = \frac{4\pi}{3}\left(\frac{c}{2\pi}\right)^3 = \frac{c^3}{6\pi^2}$

$$\text{Largest density} = \frac{5.25\ oz}{\frac{(9.00\ in)^3}{6\pi^2}} = \frac{5.25\ oz}{12.3\ in^3} = \frac{0.427\ oz}{in^3}$$

$$\text{Smallest density} = \frac{5.00\ oz}{\frac{(9.25\ in)^3}{6\pi^2}} = \frac{5.00\ oz}{13.4\ in^3} = \frac{0.373\ oz}{in^3}$$

Maximum range is: $\frac{(0.373 - 0.427)\ oz}{in^3}$ or $0.40 \pm 0.03\ oz/in^3$ (Uncertainty in 2nd decimal place.)

75. We need to calculate the maximum and minimum values of the density, given the uncertainty in each measurement. The maximum value is:

$$d_{max} = \frac{19.625\ g + 0.002\ g}{25.00\ cm^3 - 0.03\ cm^3} = \frac{19.627\ g}{24.97\ cm^3} = 0.7860\ g/cm^3$$

The minimum value of the density is:

$$d_{min} = \frac{19.625\ g - 0.002\ g}{25.00\ cm^3 + 0.03\ cm^3} = \frac{19.623\ g}{25.03\ cm^3} = 0.7840\ g/cm^3$$

The density of the liquid is between $0.7840\ g/cm^3$ and $0.7860\ g/cm^3$. These measurements are sufficiently precise to distinguish between ethanol ($d = 0.789\ g/cm^3$) and isopropyl alcohol ($d = 0.785\ g/cm^3$).

77. $V = V(\text{final}) - V(\text{initial})$; $d = \frac{28.90\ g}{9.8\ cm^3 - 6.4\ cm^3} = \frac{28.90\ g}{3.4\ cm^3} = 8.5\ g/cm^3$

$d_{max} = \frac{mass_{max}}{V_{min}}$, We get V_{min} from $9.7\ cm^3 - 6.5\ cm^3 = 3.2\ cm^3$

$$d_{max} = \frac{28.93\ g}{3.2\ cm^3} = \frac{9.0\ g}{cm^3};\ d_{min} = \frac{mass_{min}}{V_{max}} = \frac{28.87\ g}{9.9\ cm^3 - 6.3\ cm^3} = \frac{8.0\ g}{cm^3}$$

The density is: $8.5 \pm 0.5\ g/cm^3$.

Challenge Problems

79. In a subtraction, the result gets smaller but the uncertainties add. If the two numbers are very close together, the uncertainty may be larger than the result. For example, let us assume we want to take the difference of the following two measured quantities, 999,999 ± 2 and 999,996 ± 2. The difference is 3 ± 4. Because of the large uncertainty, subtracting two similar numbers is bad practice.

81. Heavy pennies (old): mean mass = 3.08 ± 0.05 g

Light pennies (new): mean mass = $\frac{(2.467 + 2.545 + 2.518)}{3}$ = 2.51 ± 0.04 g

Average density of old pennies: $d_{old} = \dfrac{\dfrac{95 \times 8.96\text{ g}}{\text{cm}^3} + \dfrac{5 \times 7.14\text{ g}}{\text{cm}^3}}{100} = \dfrac{8.9\text{ g}}{\text{cm}^3}$

Average density of new pennies: $d_{new} = \dfrac{\dfrac{2.4 \times 8.96\text{ g}}{\text{cm}^3} + \dfrac{97.6 \times 7.14\text{ g}}{\text{cm}^3}}{100} = \dfrac{7.18\text{ g}}{\text{cm}^3}$

Since $d = \dfrac{\text{mass}}{\text{volume}}$ and the volume of old and new pennies are the same, then:

$$\frac{d_{new}}{d_{old}} = \frac{\text{mass}_{new}}{\text{mass}_{old}};\quad \frac{d_{new}}{d_{old}} = \frac{7.18}{8.9} = 0.81;\quad \frac{\text{mass}_{new}}{\text{mass}_{old}} = \frac{2.51}{3.08} = 0.815$$

To the first two decimal places, the ratios are the same. We can reasonably conclude that yes, the difference in mass is accounted for by the difference in the alloy used.

83. a. One possibility is that rope B is not attached to anything and rope A and rope C are connected via a pair of pulleys and/or gears.

b. Try to pull rope B out of the box. Measure the distance moved by C for a given movement of A. Hold either A or C firmly while pulling on the other.

CHAPTER TWO

ATOMS, MOLECULES AND IONS

Questions

11. a. Atoms have mass and are neither destroyed nor created by chemical reactions. Therefore, mass is neither created nor destroyed by chemical reactions. Mass is conserved.

    b. The composition of a substance depends on the number and kinds of atoms that form it.

    c. Compounds of the same elements differ only in the numbers of atoms of the elements forming them, i.e., NO, N_2O, NO_2.

13. Deflection of cathode rays by magnetic and electric fields led to the conclusion that they were negatively charged. The cathode ray was produced at the negative electrode and repelled by the negative pole of the applied electric field.

15. The atomic number of an element is equal to the number of protons in the nucleus of an atom of that element. The mass number is the sum of the number of protons plus neutrons in the nucleus. The atomic mass is the actual mass of a particular isotope (including electrons). As we will see in chapter three, the average mass of an atom is taken from a measurement made on a large number of atoms. The average atomic mass value is listed in the periodic table.

17. A compound will always contain the same numbers (and types) of atoms. A given amount of hydrogen will react only with a specific amount of oxygen. Any excess oxygen will remain unreacted.

Exercises

Development of the Atomic Theory

19. a. The composition of a substance depends on the numbers of atoms of each element making up the compound (i.e., the formula of the compound) and not on the composition of the mixture from which it was formed.

    b. Avogadro's hypothesis implies that volume ratios are equal to molecule ratios at constant temperature and pressure. $H_2 + Cl_2 \rightarrow 2\ HCl$. From the balanced equation, the volume of HCl produced will be twice the volume of H_2 (or Cl_2) reacted.

21. $\frac{1.188}{1.188} = 1.000;\ \frac{2.375}{1.188} = 1.999;\ \frac{3.563}{1.188} = 2.999$

The masses of fluorine are simple ratios of whole numbers to each other, 1:2:3.

23. To get the atomic mass of H to be 1.00, we divide the mass of hydrogen that reacts with 1.00 g of oxygen by 0.126, i.e., $\frac{0.126}{0.126} = 1.00$. To get Na, Mg and O on the same scale, we do the same division.

Na: $\frac{2.875}{0.126} = 22.8$; Mg: $\frac{1.500}{0.126} = 11.9$; O: $\frac{1.00}{0.126} = 7.94$

	H	O	Na	Mg
Relative Value	1.00	7.94	22.8	11.9
Accepted Value	1.008	16.00	22.99	24.31

The atomic masses of O and Mg are incorrect. The atomic masses of H and Na are close. Something must be wrong about the assumed formulas of the compounds. It turns out the correct formulas are H_2O, Na_2O, and MgO. The smaller discrepancies result from the error in the atomic mass of H.

The Nature of the Atom

25. Density of hydrogen nucleus (contains one proton only):

$$V_{nucleus} = \frac{4}{3}\pi r^3 = \frac{4}{3}(3.14)(5 \times 10^{-14}\ \text{cm})^3 = 5 \times 10^{-40}\ \text{cm}^3$$

$$d = \frac{1.67 \times 10^{-24}\ \text{g}}{5 \times 10^{-40}\ \text{cm}^3} = 3 \times 10^{15}\ \text{g/cm}^3$$

Density of H-atom (contains one proton and one electron):

$$V_{atom} = \frac{4}{3}(3.14)(1 \times 10^{-8}\ \text{cm})^3 = 4 \times 10^{-24}\ \text{cm}^3$$

$$d = \frac{1.67 \times 10^{-24} + 9 \times 10^{-28}\ \text{g}}{4 \times 10^{-24}\ \text{cm}^3} = 0.4\ \text{g/cm}^3$$

27. $5.93 \times 10^{-18}\ \text{C} \times \frac{1\ \text{electron charge}}{1.602 \times 10^{-19}\ \text{C}} = 37$ negative (electron) charges on the oil drop

29. gold - Au; silver - Ag; mercury - Hg; potassium - K; iron - Fe; antimony - Sb; tungsten - W

31. fluorine - F; chlorine - Cl; bromine - Br; sulfur - S; oxygen - O; phosphorus - P

33. Sn - tin; Pt - platinum; Co - cobalt; Ni - nickel; Mg - magnesium; Ba - barium; K - potassium

35. The noble gases are He, Ne, Ar, Kr, Xe, and Rn (helium, neon, argon, krypton, xenon, and radon). Radon has only radioactive isotopes. In the periodic table the whole number enclosed in parenthesis is the mass number of the longest lived isotope of the element.

37. a. Eight; Li to Ne
b. Eight; Na to Ar
c. Eighteen; K to Kr
d. Five; N, P, As, Sb, Bi

39. a. $^{238}_{94}Pu$; 94 protons, 238 - 94 = 144 neutrons
b. $^{65}_{29}Cu$; 29 protons, 65 - 29 = 36 neutrons
c. $^{52}_{24}Cr$; 24 protons, 28 neutrons
d. $^{4}_{2}He$; 2 protons, 2 neutrons
e. $^{60}_{27}Co$; 27 protons, 33 neutrons
f. $^{54}_{24}Cr$; 24 protons, 30 neutrons

41. 9 protons means the atomic number is 9. The mass number is 9 + 10 = 19; Symbol: $^{19}_{9}F$

43. Atomic number = 63 (Eu); Charge = +63 - 60 = +3; Mass number = 63 + 88 = 151; Symbol: $^{151}_{63}Eu^{3+}$

45. Atomic number = 16 (S); Charge = +16 - 18 = -2; Mass number = 16 + 18 = 34; Symbol: $^{34}_{16}S^{2-}$

47.

Symbol	Number of protons in nucleus	Number of neutrons in nucleus	Number of electrons	Net charge
$^{75}_{33}As^{3+}$	33	42	30	3+
$^{128}_{52}Te^{2-}$	52	76	54	2-
$^{32}_{16}S$	16	16	16	0
$^{204}_{81}Tl^{+}$	81	123	80	1+
$^{195}_{78}Pt$	78	117	78	0

49. Metals: Mg, Ti, Au, Bi, Ge, Eu, Am; Nonmetals: Si, B, At, Rn, Br

51. a and d; A group is a vertical column of elements in the periodic table. Elements in the same family (group) have similar chemical properties.

53. Carbon is a nonmetal. Silicon and germanium are metalloids. Tin and lead are metals. Thus, metallic character increases as one goes down a family in the periodic table.

55. Metals lose electrons to form cations and nonmetals gain electrons to form anions. Group IA, IIA and IIIA metals form stable +1,+2 and +3 charged cations, repectively. Group VA, VIA and VIIA nonmetals form -3, -2 and -1 charged anions, respectively.

a. Lose 1 e^- to form Na^+. b. Lose 2 e^- to form Sr^{2+}. c. Lose two e^- to form Ba^{2+}.

d. Gain 1 e^- to form I^-. e. Lose 3 e^- to form Al^{3+}. f. Gain 2 e^- to form S^{2-}.

Nomenclature

57. a. sodium chloride b. rubidium oxide
c. calcium sulfide d. aluminum iodide

59. a. chromium(VI) oxide b. chromium(III) oxide c. aluminum oxide
d. sodium hydride e. calcium bromide
f. zinc chloride (Zinc only forms +2 ions so no roman numerals are needed for zinc compounds.)

61. a. potassium perchlorate b. calcium phosphate
c. aluminum sulfate d. lead(II) nitrate

63. a. nitrogen triiodide b. phosphorus trichloride
c. sulfur difluoride d. dinitrogen tetrafluoride

65. a. copper(I) iodide b. copper(II) iodide c. cobalt(II) iodide
d. sodium carbonate e. sodium hydrogen carbonate or sodium bicarbonate
f. tetrasulfur tetranitride g. sulfur hexafluoride h. sodium hypochlorite
i. barium chromate j. ammonium nitrate

67. a. CsBr b. $BaSO_4$ c. NH_4Cl d. ClO
e. $SiCl_4$ f. ClF_3 g. BeO h. MgF_2

69. a. NaOH b. $Al(OH)_3$ c. HCN
d. Na_2O_2 e. $Cu(C_2H_3O_2)_2$ f. CF_4
g. PbO h. PbO_2 i. $HC_2H_3O_2$
j. CuBr k. H_2SO_3 l. GaAs (Ga^{3+} and As^{3-} ions)

Additional Exercises

71. There should be no difference. The composition of insulin from both sources will be the same and therefore, it will have the same activity regardless of the source. As a practical note, trace contaminants in the two types of insulin may be different. These trace contaminates may be important.