Assess

The number of carbon-13 atoms is smaller than the number of carbon atoms, which it should be, given that the isotopic abundance of carbon-13 is just 1.108%. To avoid mistakes, every quantity should be clearly labeled with its appropriate unit so that units cancel properly. Two points made by this problem are, first, that the relatively small ball bearing contains a large number of carbon-13 atoms even though carbon-13's abundance is only 1.108% of all carbon atoms. Second, the size of any atom must be very small.

PRACTICE EXAMPLE A: Calculate the number of ^{63}Cu atoms in a cubic crystal of copper that measures exactly 25 nm on edge. The density of copper is 8.92 g/cm^3 and the percent isotopic abundance of ^{63}Cu is 69.17%.

PRACTICE EXAMPLE B: The United States Food and Drug Administration (USFDA) suggests a daily value of 18 mg Fe for adults and for children over four years of age. The label on a particular brand of cereal states that one serving (55 g) of dry cereal contains 45% of the daily value of Fe. Given that the percent isotopic abundance of ^{58}Fe is 0.282%, how many full servings of dry cereal must be eaten to consume exactly one mole of ^{58}Fe? The atomic weight of ^{58}Fe is 57.9333 u. Is it possible for a person to consume this much cereal in a lifetime, assuming that one full serving of cereal is eaten every day?

Exercises

Law of Conservation of Mass

1. When an iron object rusts, its mass increases. When a match burns, its mass decreases. Do these observations violate the law of conservation of mass? Explain.
2. When a strip of magnesium metal is burned in air (recall Figure 2-1), it produces a white powder that weighs more than the original metal. When a strip of magnesium is burned in a flashbulb, the bulb weighs the same before and after it is flashed. Explain the difference in these observations.
3. A 0.406 g sample of magnesium reacts with oxygen, producing 0.674 g of magnesium oxide as the only product. What mass of oxygen was consumed in the reaction?
4. A 1.446 g sample of potassium reacts with 8.178 g of chlorine to produce potassium chloride as the only product. After the reaction, 6.867 g of chlorine remains unreacted. What mass of potassium chloride was formed?
5. When a solid mixture consisting of 10.500 g calcium hydroxide and 11.125 g ammonium chloride is strongly heated, gaseous products are evolved and 14.336 g of a solid residue remains. The gases are passed into 62.316 g water, and the mass of the resulting solution is 69.605 g. Within the limits of experimental error, show that these data conform to the law of conservation of mass.
6. Within the limits of experimental error, show that the law of conservation of mass was obeyed in the following experiment: 10.00 g calcium carbonate (found in limestone) was dissolved in 100.0 mL hydrochloric acid (d = 1.148 g/mL). The products were 120.40 g solution (a mixture of hydrochloric acid and calcium chloride) and 2.22 L carbon dioxide gas (d = 1.9769 g/L).

Law of Constant Composition

7. In Example 2-1, we established that the mass ratio of magnesium to magnesium oxide is 0.455 g magnesium/0.755 g magnesium oxide.
   (a) What is the ratio of oxygen to magnesium oxide, by mass?
   (b) What is the mass ratio of oxygen to magnesium in magnesium oxide?
   (c) What is the percent by mass of magnesium in magnesium oxide?
8. Samples of pure carbon weighing 3.62, 5.91, and 7.07 g were burned in an excess of air. The masses of carbon dioxide obtained (the sole product in each case) were 13.26, 21.66, and 25.91 g, respectively.
   (a) Do these data establish that carbon dioxide has a fixed composition?
   (b) What is the composition of carbon dioxide, expressed in % C and % O, by mass?
9. In one experiment, 2.18 g sodium was allowed to react with 16.12 g chlorine. All the sodium was used up, and 5.54 g sodium chloride (salt) was produced. In a second experiment, 2.10 g chlorine was allowed to react with 10.00 g sodium. All the chlorine was used up, and 3.46 g sodium chloride was produced. Show that these results are consistent with the law of constant composition.
10. When 3.06 g hydrogen was allowed to react with an excess of oxygen, 27.35 g water was obtained. In a second experiment, a sample of water was decomposed by electrolysis, resulting in 1.45 g hydrogen and 11.51 g oxygen. Are these results consistent with the law of constant composition? Demonstrate why or why not.
11. In one experiment, the burning of 0.312 g sulfur produced 0.623 g sulfur dioxide as the sole product of the reaction. In a second experiment, 0.842 g sulfur dioxide was obtained. What mass of sulfur must have been burned in the second experiment?
12. In one experiment, the reaction of 1.00 g mercury and an excess of sulfur yielded 1.16 g of a sulfide of mercury as the sole product. In a second experiment, the same sulfide was produced in the reaction of 1.50 g mercury and 1.00 g sulfur.
    (a) What mass of the sulfide of mercury was produced in the second experiment?
    (b) What mass of which element (mercury or sulfur) remained *unreacted* in the second experiment?

Law of Multiple Proportions

13. Sulfur forms two compounds with oxygen. In the first compound, 1.000 g sulfur is combined with 0.998 g oxygen, and in the second, 1.000 g sulfur is combined with 1.497 g oxygen. Show that these results are consistent with Dalton's law of multiple proportions.

14. Phosphorus forms two compounds with chlorine. In the first compound, 1.000 g of phosphorus is combined with 3.433 g chlorine, and in the second, 2.500 g phosphorus is combined with 14.308 g chlorine. Show that these results are consistent with Dalton's law of multiple proportions.

15. The following data were obtained for compounds of nitrogen and hydrogen:

Compound	Mass of Nitrogen, g	Mass of Hydrogen, g
A	0.500	0.108
B	1.000	0.0720
C	0.750	0.108

(a) Show that these data are consistent with the law of multiple proportions.
(b) If the formula of compound B is N_2H_2, what are the formulas of compounds A and C?

16. The following data were obtained for compounds of iodine and fluorine:

Compound	Mass of Iodine, g	Mass of Fluorine, g
A	1.000	0.1497
B	0.500	0.2246
C	0.750	0.5614
D	1.000	1.0480

(a) Show that these data are consistent with the law of multiple proportions.
(b) If the formula for compound A is IF, what are the formulas for compounds B, C, and D?

17. There are two oxides of copper. One oxide has 20% oxygen, by mass. The second oxide has a *smaller* percent of oxygen than the first. What is the probable percent of oxygen in the second oxide?

18. The two oxides of carbon described on page 38 were CO and CO_2. Another oxide of carbon has 1.106 g of oxygen in a 2.350 g sample. In what ratio are carbon and oxygen atoms combined in molecules of this third oxide? Explain.

Fundamental Charges and Mass-to-Charge Ratios

19. The following observations were made for a series of five oil drops in an experiment similar to Millikan's (see Figure 2-8). Drop 1 carried a charge of 1.28×10^{-18} C; drops 2 and 3 each carried $\frac{1}{2}$ the charge of drop 1; drop 4 carried $\frac{1}{3}$ the charge of drop 1; drop 5 had a charge four times that of drop 1. Are these data consistent with the value of the electronic charge given in the text? Could Millikan have inferred the charge on the electron from this particular series of data? Explain.

20. In an experiment similar to that described in Exercise 19, drop 1 carried a charge of 6.41×10^{-19} C; drop 2 had $\frac{1}{2}$ the charge of drop 1; drop 3 had twice the charge of drop 1; drop 4 had a charge of 1.44×10^{-18} C; and drop 5 had $\frac{1}{3}$ the charge of drop 4. Are these data consistent with the value of the electronic charge given in the text? Could Millikan have inferred the charge on the electron from this particular series of data? Explain.

21. Use data from Table 2.1 to verify that
(a) the mass of electrons is about 1/2000 that of H atoms;
(b) the mass-to-charge ratio (m/e) for positive ions is considerably larger than that for electrons.

22. Determine the approximate value of m/e in grams per coulomb for the ions ${}^{127}_{53}I^-$ and ${}^{32}_{16}S^{2-}$. Why are these values only approximate?

Atomic Number, Mass Number, and Isotopes

23. The following radioactive isotopes have applications in medicine. Write their symbols in the form ${}^{A}_{Z}E$. **(a)** cobalt-60; **(b)** phosphorus-32; **(c)** iron-59; **(d)** radium-226.

24. For the isotope ${}^{202}Hg$, express the percentage of the fundamental particles in the nucleus that are neutrons.

25. Complete the following table. What minimum amount of information is required to completely characterize an atom or ion?
[*Hint:* Not all rows can be completed.]

Name	Symbol	Number Protons	Number Electrons	Number Neutrons	Mass Number
Sodium	${}^{23}_{11}Na$	11	11	12	23
Silicon	—	—	—	14	—
—	—	37	—	—	85
—	${}^{40}K$	—	—	—	—
—	—	—	33	42	—
—	${}^{20}Ne^{2+}$	—	—	—	—
—	—	—	—	—	80
—	—	—	—	126	—

26. Arrange the following species in order of increasing **(a)** number of electrons; **(b)** number of neutrons; **(c)** mass.

$^{112}_{50}Sn$ $^{40}_{18}Ar$ $^{122}_{52}Te$ $^{59}_{29}Cu$ $^{120}_{48}Cd$ $^{58}_{27}Co$ $^{39}_{19}K$

27. For the atom ^{108}Pd with mass 107.90389 u, determine **(a)** the numbers of protons, neutrons, and electrons in the atom; **(b)** the ratio of the mass of this atom to that of an atom of $^{12}_{6}C$.

28. For the ion $^{228}Ra^{2+}$ with a mass of 228.030 u, determine **(a)** the numbers of protons, neutrons, and electrons in the ion; **(b)** the ratio of the mass of this ion to that of an atom of ^{16}O (refer to page 47).

29. An isotope of silver has a mass that is 6.68374 times that of oxygen-16. What is the mass in u of this isotope? (Refer to page 47.)

30. The ratio of the masses of the two naturally occurring isotopes of indium is 1.0177:1. The heavier of the two isotopes has 7.1838 times the mass of ^{16}O. What are the masses in u of the two isotopes? (Refer to page 47.)

31. The following data on isotopic masses are from a chemical handbook. What is the ratio of each of these masses to that of $^{12}_{6}C$? **(a)** $^{35}_{17}Cl$, 34.96885 u; **(b)** $^{26}_{12}Mg$, 25.98259 u; **(c)** $^{222}_{86}Rn$, 222.0175 u.

32. The following ratios of masses were obtained with a mass spectrometer: $^{19}_{9}F/^{12}_{6}C = 1.5832$; $^{35}_{17}Cl/^{19}_{9}F = 1.8406$; $^{81}_{35}Br/^{35}_{17}Cl = 2.3140$. Determine the mass of a $^{81}_{35}Br$ atom in amu.

33. Which of the following species has **(a)** equal numbers of neutrons and electrons; **(b)** protons, neutrons, and electrons in the ratio 9:11:8; **(c)** a number of neutrons equal to the number of protons plus one-half the number of electrons?

$^{24}Mg^{2+}$, ^{47}Cr, $^{60}Co^{3+}$, $^{35}Cl^{-}$, $^{124}Sn^{2+}$, ^{226}Th, ^{90}Sr

34. Given the same species as listed in Exercise 33, which has **(a)** equal numbers of neutrons and protons; **(b)** protons contributing more than 50% of the mass; **(c)** about 50% more neutrons than protons?

35. An isotope with mass number 44 has four more neutrons than protons. This is an isotope of what element?

36. Identify the isotope X that has one more neutron than protons and a mass number equal to nine times the charge on the ion X^{3+}.

37. Iodine has many radioactive isotopes. Iodine-123 is a radioactive isotope used for obtaining images of the thyroid gland. Iodine-123 is administered to patients in the form of sodium iodide capsules that contain $^{123}I^{-}$ ions. Determine the number of neutrons, protons, and electrons in a single $^{123}I^{-}$ ion.

38. Iodine-131 is a radioactive isotope that has important medical uses. Small doses of iodine-131 are used for treating hyperthyroidism (overactive thyroid) and larger doses are used for treating thyroid cancer. Iodine-131 is administered to patients in the form of sodium iodide capsules that contain $^{131}I^{-}$ ions. Determine the number of neutrons, protons, and electrons in a single $^{131}I^{-}$ ion.

39. Americium-241 is a radioactive isotope that is used in high-precision gas and smoke detectors. How many neutrons, protons, and electrons are there in an atom of americium-241?

40. Some foods are made safer to eat by being exposed to gamma rays from radioactive isotopes, such as cobalt-60. The energy from the gamma rays kills bacteria in the food. How many neutrons, protons, and electrons are there in an atom of cobalt-60?

Atomic Mass Units, Atomic Masses

41. Which statement is probably true concerning the masses of *individual* chlorine atoms: *All have*, *some have*, or *none has* a mass of 35.45 u? Explain.

42. The mass of a carbon-12 atom is taken to be exactly 12 u. Are there likely to be any other atoms with an *exact* integral (whole number) mass, expressed in u? Explain.

43. Magnesium has three naturally occurring isotopes. Their masses are 23.985042 u, 24.985837 u, and 25.982593 u. What is the weighted-average atomic mass of magnesium in a sample for which the percent isotopic abundances of these three isotopes are 78.99%, 10.00%, and 11.01%, respectively?

44. There are four naturally occurring isotopes of chromium. Their masses and percent isotopic abundances are 49.9461 u, 4.35%; 51.9405 u, 83.79%; 52.9407 u, 9.50%; and 53.9389 u, 2.36%. Calculate the weighted-average atomic mass of chromium.

45. The two naturally occurring isotopes of silver have the following abundances: ^{107}Ag, 51.84%; ^{109}Ag, 48.16%. The mass of ^{107}Ag is 106.905092 u. What is the mass of ^{109}Ag?

46. Gallium has two naturally occurring isotopes. One of them, gallium-69, has a mass of 68.925581 u and a percent isotopic abundance of 60.11%. What must be the mass and percent isotopic abundance of the other isotope, gallium-71?

47. The three naturally occurring isotopes of potassium are ^{39}K, 38.963707 u; ^{40}K, 39.963999 u; and ^{41}K. The percent isotopic abundances of ^{39}K and ^{41}K are 93.2581% and 6.7302%, respectively. Determine the isotopic mass of ^{41}K.

48. Use the conventional atomic mass of boron to estimate the fractional isotopic abundances of the two naturally occurring isotopes, ^{10}B and ^{11}B. These isotopes have masses of 10.012937 u and 11.009305 u, respectively.

Mass Spectrometry

49. A mass spectrum of germanium displayed peaks at mass numbers 70, 72, 73, 74, and 76, with relative heights of 20.5, 27.4, 7.8, 36.5, and 7.8, respectively.
(a) In the manner of Figure 2-14, sketch this mass spectrum.
(b) Estimate the weighted-average atomic mass of germanium, and state why this result is only approximately correct.

50. Hydrogen and chlorine atoms react to form simple diatomic molecules in a 1:1 ratio, that is, HCl. The percent isotopic abundances of the chlorine isotopes are ^{35}Cl and ^{37}Cl are estimated to be 75.77% and 24.23%, respectively. The percent isotopic abundances of ^{2}H and ^{3}H are estimated to be 0.015% and less than 0.001%, respectively.
(a) How many different HCl molecules are possible, and what are their mass numbers (that is, the sum of the mass numbers of the H and Cl atoms)?
(b) Which is the most abundant of the possible HCl molecules? Which is the second most abundant?

The Periodic Table

51. Refer to the periodic table inside the front cover and identify
(a) the element that is in group 14 and the fourth period
(b) one element similar to and one unlike sulfur
(c) the alkali metal in the fifth period
(d) the halogen element in the sixth period

52. Refer to the periodic table inside the front cover and identify
(a) the element that is in group 11 and the sixth period
(b) an element with atomic number greater than 50 that has properties similar to the element with atomic number 18
(c) the group number of an element E that forms an ion E^{2-}
(d) an element M that you would expect to form the ion M^{3+}

53. Assuming that the seventh period of the periodic table has 32 members, what should be the atomic number of **(a)** the noble gas following radon (Rn); **(b)** the alkali metal following francium (Fr)?

54. Find the several pairs of elements that are "out of order" in terms of increasing atomic mass and explain why the reverse order is necessary.

The Avogadro Constant and the Mole

55. What is the total number of atoms in **(a)** 15.8 mol Fe; **(b)** 0.000467 mol Ag; **(c)** 8.5×10^{-11} mol Na?

56. *Without doing detailed calculations*, indicate which of the following quantities contains the greatest number of atoms: 6.022×10^{23} Ni atoms, 25.0 g nitrogen, 52.0 g Cr, 10.0 cm^3 Fe ($d = 7.86$ g/cm^3). Explain your reasoning.

57. Determine
(a) the number of moles of Zn in a 415.0 g sample of zinc metal
(b) the number of Cr atoms in 147.4 kg chromium
(c) the mass of a one-trillion-atom (1.0×10^{12}) sample of metallic gold
(d) the average mass of a fluorine atom (in grams)

58. Determine
(a) the number of Kr atoms in a 5.25 mg sample of krypton
(b) the molar mass, M, and identity of an element if the mass of a 2.80×10^{22} atom sample of the element is 2.09 g
(c) the mass of a sample of phosphorus that contains the same number of atoms as 44.75 g of magnesium

59. How many Cu atoms are present in a piece of sterling-silver jewelry weighing 33.24 g? (Sterling silver is a silver–copper alloy containing 92.5% Ag by mass.)

60. How many atoms are present in a 50.0 cm^3 sample of plumber's solder, a lead–tin alloy containing 67% Pb by mass and having a density of 9.4 g/cm^3?

61. How many ^{204}Pb atoms are present in a piece of lead weighing 215 mg? The percent isotopic abundance of ^{204}Pb is 1.4%.

62. A particular lead–cadmium alloy is 8.0% cadmium by mass. What mass of this alloy, in grams, must you weigh out to obtain a sample containing 7.25×10^{23} Cd atoms?

63. Medical experts generally believe a level of 30 μg Pb per deciliter of blood poses a significant health risk (1 dL = 0.1 L). Express this level **(a)** in the unit mol Pb/L blood; **(b)** as the number of Pb atoms per milliliter of blood.

64. During a severe episode of air pollution, the concentration of lead in the air was observed to be 3.11 μg Pb/m^3. How many Pb atoms would be present in a 0.500 L sample of this air (the volume of air displaced in the lungs between inhaling and exhaling)?

65. *Without doing detailed calculations*, determine which of the following samples has the greatest number of atoms:
(a) a cube of iron with a length of 10.0 cm ($d = 7.86$ g/cm^3)
(b) 1.00 kg of hydrogen contained in a 10,000 L balloon
(c) a mound of sulfur weighing 20.0 kg
(d) a 76 lb sample of liquid mercury ($d = 13.5$ g/mL)

66. *Without doing detailed calculations*, determine which of the following samples occupies the largest volume:
(a) 25.5 mol of sodium metal ($d = 0.971$ g/cm^3)
(b) 0.725 L of liquid bromine ($d = 3.12$ g/mL)
(c) 1.25×10^{25} atoms of chromium metal ($d = 9.4$ g/cm^3)
(d) 2.15 kg of plumber's solder ($d = 9.4$ g/cm^3), a lead–tin alloy with a 2:1 atom ratio of lead to tin

Integrative and Advanced Exercises

67. A solution was prepared by dissolving 2.50 g potassium chlorate (a substance used in fireworks and flares) in 100.0 mL of water at 40 °C. When the solution was cooled to 20 °C, its volume was still found to be 100.0 mL, but some of the potassium chlorate had crystallized (deposited from the solution as a solid). At 40 °C, the density of water is 0.9922 g/mL, and at 20 °C, the potassium chlorate solution had a density of 1.0085 g/mL.
(a) Estimate, to two significant figures, the mass of potassium chlorate that crystallized.
(b) Why can't the answer in **(a)** be given more precisely?

68. William Prout (1815) proposed that all other atoms are built up of hydrogen atoms, suggesting that all elements should have integral atomic masses based on an atomic mass of one for hydrogen. This hypothesis appeared discredited by the discovery of atomic masses, such as 24.3 u for magnesium and 35.5 u for chlorine. In terms of modern knowledge, explain why Prout's hypothesis is actually quite reasonable.

69. Fluorine has a single atomic species, ^{19}F. Determine the atomic mass of ^{19}F by summing the masses of its protons, neutrons, and electrons, and compare your results with the value listed on the inside front cover. Explain why the agreement is poor.

70. Use 1×10^{-13} cm as the approximate diameter of the spherical nucleus of the hydrogen-1 atom, together with data from Table 2.1, to estimate the density of matter in a proton.

71. Use fundamental definitions and statements from Chapters 1 and 2 to establish the fact that 6.022×10^{23} u = 1.000 g.

72. In each case, identify the element in question.
(a) The mass number of an atom is 234, and the atom has 60.0% more neutrons than protons.
(b) An ion with a +2 charge has 10.0% more protons than electrons.
(c) An ion with a mass number of 110 and a 2+ charge has 25.0% more neutrons than electrons.

73. Determine the only possible +2 ion for which the following two conditions are both satisfied:
- The net ionic charge is *one-tenth* the nuclear charge.
- The number of neutrons is *four* more than the number of electrons.

74. Determine the only possible isotope (E) for which the following conditions are met:
- The mass number of E is 2.50 times its atomic number.
- The atomic number of E is equal to the mass number of another isotope (Y). In turn, isotope Y has a neutron number that is 1.33 times the atomic number of Y and equal to the neutron number of selenium-82.

75. Suppose we redefined the atomic mass scale by arbitrarily assigning to the naturally occurring *mixture* of chlorine isotopes an atomic mass of 35.00000 u.
(a) What would be the atomic masses of helium, sodium, and iodine on this new atomic mass scale?
(b) Why do these three elements have nearly integral (whole-number) atomic masses based on carbon-12, but not based on naturally occurring chlorine?

76. The two naturally occurring isotopes of nitrogen have masses of 14.0031 and 15.0001 u, respectively. Use the conventional atomic mass of nitrogen to estimate the percentage of ^{15}N atoms in naturally occurring nitrogen.

77. The masses of the naturally occurring mercury isotopes are ^{196}Hg, 195.9658 u; ^{198}Hg, 197.9668 u; ^{199}Hg, 198.9683 u; ^{200}Hg, 199.9683 u; ^{201}Hg, 200.9703 u; ^{202}Hg, 201.9706 u; and ^{204}Hg, 203.9735 u. Use these data, together with data from Figure 2-14, to calculate the weighted-average atomic mass of mercury.

78. Germanium has three major naturally occurring isotopes: ^{70}Ge (69.92425 u, 20.85%), ^{72}Ge (71.92208 u, 27.54%), ^{74}Ge (73.92118 u, 36.29%). There are also two minor isotopes: ^{73}Ge (72.92346 u) and ^{76}Ge (75.92140 u). Calculate the percent isotopic abundances of the two minor isotopes. Comment on the precision of these calculations.

79. From the densities of the lines in the mass spectrum of krypton gas, the following observations were made:
- Somewhat more than 50% of the atoms were krypton-84.
- The numbers of krypton-82 and krypton-83 atoms were essentially equal.
- The number of krypton-86 atoms was 1.50 times as great as the number of krypton-82 atoms.
- The number of krypton-80 atoms was 19.6% of the number of krypton-82 atoms.
- The number of krypton-78 atoms was 3.0% of the number of krypton-82 atoms.

The masses of the isotopes are
^{78}Kr, 77.9204 u ^{80}Kr, 79.9164 u ^{82}Kr, 81.9135 u
^{83}Kr, 82.9141 u ^{84}Kr, 83.9115 u ^{86}Kr, 85.9106 u
The weighted-average atomic mass of Kr is 83.80. Use these data to calculate the percent isotopic abundances of the krypton isotopes.

80. The two naturally occurring isotopes of chlorine are ^{35}Cl (34.9689 u, 75.77%) and ^{37}Cl (36.9658 u, 24.23%). The two naturally occurring isotopes of bromine are ^{79}Br (78.9183 u, 50.69%) and ^{81}Br (80.9163 u, 49.31%). Chlorine and bromine combine to form bromine monochloride, BrCl. Sketch a mass spectrum for BrCl with the relative number of molecules plotted against molecular mass (similar to Figure 2-14).

81. How many atoms are present in a 1.50 m length of 20-gauge copper wire? A 20-gauge wire has a diameter of 0.03196 in., and the density of copper is 8.92 g/cm^3.

82. Monel metal is a corrosion-resistant copper–nickel alloy used in the electronics industry. A particular alloy with a density of 8.80 g/cm^3 and containing 0.022% Si by mass is used to make a rectangular plate 15.0 cm long, 12.5 cm wide, 3.00 mm thick, and has a 2.50 cm diameter hole drilled through its center. How many silicon-30 atoms are found in this plate? The mass of a silicon-30 atom is 29.97376 u, and the percent isotopic abundance of silicon-30 is 3.10%.

83. Deuterium, ^{2}H (2.0140 u), is sometimes used to replace the principal hydrogen isotope ^{1}H in chemical studies. The percent isotopic abundance of deuterium is 0.015%. If it can be done with 100% efficiency, what mass of hydrogen gas would have to be processed to obtain a sample containing 2.50×10^{21} ^{2}H atoms?

84. An alloy that melts at about the boiling point of water has Bi, Pb, and Sn atoms in the ratio 10:6:5, respectively. What mass of alloy contains a total of one mole of atoms?

85. A particular silver solder (used in the electronics industry to join electrical components) is to have the *atom* ratio of 5.00 Ag/4.00 Cu/1.00 Zn. What masses of the three metals must be melted together to prepare 1.00 kg of the solder?

86. A low-melting Sn–Pb–Cd alloy called *eutectic alloy* is analyzed. The *mole* ratio of tin to lead is 2.73:1.00, and the *mass* ratio of lead to cadmium is 1.78:1.00. What is the mass percent composition of this alloy?

87. In an experiment, 125 cm^3 of zinc and 125 cm^3 of iodine are mixed together and the iodine is completely converted to 164 cm^3 of zinc iodide. What volume of zinc remains unreacted? The densities of zinc, iodine, and zinc iodide are 7.13 g/cm^3, 4.93 g/cm^3, and 4.74 g/cm^3, respectively.

88. Atoms are spherical and so when silver atoms pack together to form silver metal, they cannot fill all the available space. In a sample of silver metal, approximately 26.0% of the sample is empty space. Given that the density of silver metal is 10.5 g/cm^3, what is the radius of a silver atom? Express your answer in picometers.

Feature Problems

89. The data Lavoisier obtained in the experiment described on page 35 are as follows:

Before heating: glass vessel + tin + air
= 13 onces, 2 gros, 2.50 grains

After heating: glass vessel + tin calx + remaining air
= 13 onces, 2 gros, 5.62 grains

How closely did Lavoisier's results conform to the law of conservation of mass? (1 livre = 16 onces; 1 once = 8 gros; 1 gros = 72 grains. In modern terms, 1 livre = 30.59 g.)

90. Some of Millikan's oil-drop data are shown below. The measured quantities were not actual charges on oil drops but were proportional to these charges. Show that these data are consistent with the idea of a fundamental electronic charge.

Observation	Measured Quantity	Observation	Measured Quantity
1	19.66	8	53.91
2	24.60	9	59.12
3	29.62	10	63.68
4	34.47	11	68.65
5	39.38	12	78.34
6	44.42	13	83.22
7	49.41		

91. Before 1961, the standard for atomic masses was the isotope ^{16}O, to which physicists assigned a value of exactly 16. At the same time, chemists assigned a value of exactly 16 to the naturally occurring mixture of the isotopes ^{16}O, ^{17}O, and ^{18}O. Would you expect atomic masses listed in a 60-year-old text to be the same, generally higher, or generally lower than in this text? Explain.

92. German chemist Fritz Haber proposed paying off the reparations imposed against Germany after World War I by extracting gold from seawater. Given that (1) the amount of the reparations was $28.8 billion dollars, (2) the value of gold at the time was about $21.25 per troy ounce (1 troy ounce = 31.103 g), and (3) gold occurs in seawater to the extent of 4.67×10^{17} atoms per ton of seawater (1 ton = 2000 lb), how many cubic kilometers of seawater would have had to be processed to obtain the required amount of gold? Assume that the density of seawater is 1.03 g/cm^3. (Haber's scheme proved to be commercially infeasible, and the reparations were never fully paid.)

93. Mass spectrometry is one of the most versatile and powerful tools in chemical analysis because of its capacity to discriminate between atoms of different masses. When a sample containing a mixture of isotopes is introduced into a mass spectrometer, the ratio of the peaks observed reflects the ratio of the percent isotopic abundances. This ratio provides an internal standard from which the amount of a certain isotope present in a sample can be determined. This is accomplished by deliberately introducing a known quantity of a particular isotope into the sample to be analyzed. A comparison of the new isotope ratio to the first ratio allows the determination of the amount of the isotope present in the original sample.

An analysis was done on a rock sample to determine its rubidium content. The rubidium content of a portion of rock weighing 0.350 g was extracted, and to the extracted sample was added an additional 29.45 μg of ^{87}Rb. The mass spectrum of this spiked sample showed a ^{87}Rb peak that was 1.12 times as high as the peak for ^{85}Rb. Assuming that the two isotopes react identically, what is the Rb content of the rock (expressed in parts per million by mass)? The isotopic abundances and isotopic masses are shown in the table.

Isotope	% Isotopic Abundance	Atomic Mass, u
^{87}Rb	27.83	86.909
^{85}Rb	72.17	84.912

Self-Assessment Exercises

94. In your own words, define or explain these terms or symbols: **(a)** $^{A}_{Z}E$; **(b)** β particle; **(c)** isotope; **(d)** ^{16}O; **(e)** molar mass.

95. Briefly describe
(a) the law of conservation of mass
(b) Rutherford's nuclear atom
(c) weighted-average atomic mass
(d) a mass spectrum

96. Explain the important distinctions between each pair of terms:
(a) cathode rays and X-rays
(b) protons and neutrons
(c) nuclear charge and ionic charge
(d) periods and groups of the periodic table
(e) metal and nonmetal
(f) the Avogadro constant and the mole

97. A certain element contains one atom of mass 10.013 u for every four atoms of mass 11.009 u. Compute the atomic weight of the element.

98. When 10.0 g zinc and 8.0 g sulfur are allowed to react, all the zinc is consumed, 14.9 g zinc sulfide is produced, and the mass of unreacted sulfur remaining is
(a) 2.0 g
(b) 3.1 g
(c) 4.9 g
(d) impossible to predict from this information alone

99. One oxide of rubidium has 0.187 g O per gram of Rb. A possible O:Rb mass ratio for a second oxide of rubidium is **(a)** 16:85.5; **(b)** 8:42.7; **(c)** 1:2.674; **(d)** any of these.

100. An attempt was made to determine the atomic mass of element X. If X forms a compound with oxygen that contains 46.7% X by mass and has the formula XO, what is the atomic mass of X?

101. Cathode rays
(a) may be positively or negatively charged
(b) are a form of electromagnetic radiation similar to visible light
(c) have properties identical to β particles
(d) have masses that depend on the cathode that emits them

102. The scattering of α particles by thin metal foils established that
(a) the mass of an atom is concentrated in a positively charged nucleus
(b) electrons are fundamental particles of all matter
(c) all electrons carry the same charge
(d) atoms are electrically neutral

103. Which of the following have the same charge and approximately the same mass?
(a) an electron and a proton; **(b)** a proton and a neutron; **(c)** a hydrogen atom and a proton; **(d)** a neutron and a hydrogen atom; **(e)** an electron and an H^- ion

104. Which of the following is *not* a fundamental particle?
(a) proton; **(b)** neutron; **(c)** beta particle; **(d)** alpha particle; **(e)** all are fundamental particles

105. Which of the following scientists did *not* contribute to determining the *structure* of the atom?
(a) Thomson; **(b)** Rutherford; **(c)** Millikan; **(d)** Dalton; **(e)** Becquerel

106. A subatomic particle that has about the same mass as the hydrogen atom and a negative charge is called
(a) a proton; **(b)** a neutron; **(c)** an electron; **(d)** an isotope; (e) none of these.

107. What is the correct symbol for the species that contains 18 neutrons, 17 protons, and 16 electrons?

108. The properties of magnesium will most resemble those of which of the following? **(a)** cesium; **(b)** sodium; **(c)** aluminum; **(d)** calcium; **(e)** manganese.

109. Which group in the main group of elements contains **(a)** no metals or metalloids? **(b)** only one metal or metalloid? **(c)** only one nonmetal? **(d)** only nonmetals?

110. The two species that have the same number of electrons as ^{32}S are **(a)** ^{32}Cl; **(b)** $^{34}S^+$; **(c)** $^{33}P^+$; **(d)** $^{28}Si^{2-}$; **(e)** $^{35}S^{2-}$; **(f)** $^{40}Ar^{2+}$; **(g)** $^{40}Ca^{2+}$.

111. To four significant figures, all of the following masses are possible for an individual titanium atom except one. The exception is **(a)** 45.95 u; **(b)** 46.95 u; **(c)** 47.87 u; **(d)** 47.95 u; **(e)** 48.95 u; **(f)** 49.94 u.

112. The mass of the isotope $^{84}_{36}Xe$ is 83.9115 u. If the atomic mass scale were redefined so that $^{84}_{36}Xe = 84$ u, *exactly*, the mass of the $^{12}_{6}C$ isotope would be **(a)** 11.9115 u; **(b)** 11.9874 u; **(c)** 12 u exactly; **(d)** 12.0127 u; **(e)** 12.0885 u.

113. A 5.585-kg sample of iron (Fe) contains
(a) 10.0 mol Fe
(b) twice as many atoms as does 600.6 g C
(c) 10 times as many atoms as does 52.00 g Cr
(d) 6.022×10^{24} atoms

114. A 91.84 g sample of Ti contains **(a)** 4.175 mol of Ti; **(b)** 6.022×10^{23} Ti atoms; **(c)** 1.155×10^{24} protons; **(d)** 2.542×10^{25} electrons; **(e)** none of these.

115. There are three common iron–oxygen compounds. The one with the greatest proportion of iron has one Fe atom for every O atom and the formula FeO. A second compound has 2.327 g Fe per 1.000 g O, and the third has 2.618 g Fe per 1.000 g O. What are the formulas of these other two iron–oxygen compounds?

116. The four naturally occurring isotopes of strontium have the atomic masses 83.9134 u; 85.9093 u; 86.9089 u; and 87.9056 u. The percent isotopic abundance of the lightest isotope is 0.56% and of the heaviest, 82.58%. Estimate the percent isotopic abundances of the other two. Why is this result only a rough approximation?

117. Gold is present in seawater to the extent of 0.15 mg/ton. Assume the density of the seawater is 1.03 g/mL and determine how many Au atoms could conceivably be extracted from 0.250 L of seawater (1 ton = 2.000×10^3 lb; 1 kg = 2.205 lb).

118. Appendix E describes a useful study aid known as concept mapping. Using the method presented in Appendix E, construct a concept map illustrating the different concepts in Sections 2-7 and 2-8.