Solution Manual for Chemistry for Changing Times 14th Edition by Hill McCreary ISBN 0321972023 9780321972026

Full link download:

Test Bank:

Solution Manual:

Chapter 2: Atoms

Overview

Chapter 2 begins with the Greek concept of matter and moves chronologically forward through the development of the atomistic model of matter.

Lecture Outline

2.1 Atoms: Ideas from the Ancient Greeks

The Greek philosophers had two main ideas regarding matter: the atomistic view (Leucippus and Democritus) the continuous view (Aristotle)

2.2 Scientific Laws: Conservation of Mass and Definite Proportions

Lavoisier was the first great experimentalist. More than anything, he introduced experimental chemistry to Western Civilization. Through his experiments he formulated the law of conservation of mass, which states: Matter is neither created nor destroyed during a chemical reaction.

2.3 John Dalton and the Atomic Theory of Matter

John Dalton summarized the work of Lavoisier and Proust to form the famous atomic theory.

2.4 The Mole and Molar Mass

The mole and Avogadro's number
Molar mass, atomic mass
Mass and atom ratios
Moles

2.5 Mendeleev and the Periodic Table

The periodic table: Mendeleev and Meyer
John Dalton established relative atomic masses using hydrogen as a base. Mendeleev arranged a table of elements according to increasing atomic weights placing elements

with similar properties in the same column. This was the beginning of the modern periodic table!

2.5 Atoms and Molecules: Real and Relevant

Demonstrations

1. Electrolysis of water.

2. Fill three flasks with zinc (gray powder), sulfur (yellow powder), and zinc sulfide (white powder), and show them to the class to illustrate how the properties of a compound differ from those of the component elements. [Or use mercury (silver liquid), oxygen (colorless gas), and mercuric oxide (red powder).]

3. Use samples of metals and nonmetals to demonstrate the difference in their properties. (Aluminum foil, copper wire, nickel coin, zinc strip; capped bottle of liquid bromine, balloon filled with hydrogen, bottles of carbon, sulfur, etc.)

 Chapter 2: Atoms

4. Using the largest nuts and bolts you can find, start with unequal numbers of nuts and bolts (e.g., 5 and 7). Connect the pairs of nuts and bolts, with some remaining unpaired, to illustrate the law of definite proportions.

5. Heat a sample of HgO in a large test tube with a burner. Mercury droplets form on the inside of the tube. Test for oxygen with a glowing splint, which bursts into flame when inserted into the test tube.

Review Questions

1. (a) The atomic view of matter assumes that matter is made up of small unit particles that cannot be further subdivided and still be the same kind of matter, while the continuous view of matter is just that— continuous throughout with no smaller particle. (b) The Greeks believed that there were four basic elements: earth, air, fire, and water, while today we recognize the existence of elements that are pure chemical substances made of a single atom.

2. Democritus described the tiny ultimate particle as “atomos” (which means “cannot be cut”). Matter appears continuous, and without experimental chemistry there is no way to prove that matter is not continuous. The discoveries are summarized in the laws of definite proportions and multiple proportions.

3. discrete: a, c, e; continuous: b, d

4. Robert Boyle, in *The Sceptical Chymist* (1661), stated that if a substance could be broken down into simpler substances, it was not an element.

5. A compound contains elements in certain definite proportions and in no other combinations. All samples of water will have 2 atoms of hydrogen and 1 atom of oxygen. The ratio of atoms yields a constant 2:1 volume ratio.

6. Lavoisier is the “father” of modern chemistry. He was the first to carry out a series of experiments in which the masses of reactants and products were carefully recorded. These experiments led Lavoisier to formulate the law of conservation of mass.

7. The law of definite proportions.

8. (a) Dalton’s atomic theory states: “No atoms are created or broken apart in a chemical reaction.” This is a restatement of the law of conservation of mass. When 3 g of carbon is allowed to react with 8 g of oxygen, 11 g of carbon dioxide is formed. (b) One of the statements of Dalton’s atomic theory is that “compounds are formed when atoms of different elements combine in fixed proportions.” This is a restatement of the law of definite proportions. (c) One statement of Dalton’s atomic theory is that

“compounds are formed when atoms of different elements combine in fixed proportions.” This is a restatement of the law of definite proportions but may also explain the law of multiple proportions.

9. The law of multiple proportions.

 Chapter 2: Atoms

10. Elements can combine in more than one set of proportions. (They will still of course combine in small whole-number ratios.) (a) The ratio of the oxygen in ClO_2 to oxygen in ClO, the second compound, is 2:1. (b) The ratio of fluorine ClF_3 to fluorine in ClF 3 parts to 1 part by mass. (c) The ratio of oxygen in P_4O_6 to oxygen in P_4O_{10} is 6:10, which reduces to 3:5 by mass.

11. Rectangle C cannot represent the mixture after reaction. Rectangle C contains 15 oxygen atoms; there are 14 oxygen atoms in the initial mixture. The law of conservation of mass has been violated.

12. (a) Avogadro's number represents the number of particles contained in one mole. (b) 6.02×10^{23} oxygen atoms. (c) 6.02×10^{23} molecules (d) 28.02 g.

13. Dmitri Mendeleev linked the mass of elements to chemical properties and created periodic law and thus the first form of the periodic table.

14. Oersted's experiment proved that aluminum could not be broken down further; it was an element.

Problems

15. (a) The container would weigh the same regardless of how much time passed. (b) No. The mouse would exchange gases with the surroundings.

16. The iron atoms have not been destroyed; they are in solution.

17. c. Carbon dioxide gas will be generated and released as the anti-acid pill dissolves.

18. (a) The carbon, which forms the diamond, has been converted to carbon dioxide and escaped into the atmosphere. (b) The carbon in the wood is converted to carbon dioxide, but the residual ash is not. Burning the diamond in an enclosed environment would trap the carbon dioxide for further analysis.

19. $$\frac{2.02\text{ g H}}{24.02\text{ g C}} \times 78.5\text{ g C} = 6.60\text{ g H}$$

20. $$\frac{44.01\text{ g nitrous oxide}}{28.01\text{ g nitrogen}} \times 48.7\text{ g nitrogen} = 76.5\text{ g nitrous oxide}$$

21. $$\underline{36.03\text{ g C}} \quad = 72.06\text{ g C}$$
$$\times\ 88.20\text{ g C }H_3\ _8$$

 Chapter 2: Atoms

44.10 g C_3H_8

44.012.011g g C C O x$_2$ 300.06 g CO_2 = 81.88 g C

It takes 81.88 g of carbon to form 300.06 g of carbon dioxide, and you only have 72.06 g of carbon available from 88.20 g of propane. The answer violates the law of conservation of mass.

22. Mass before the reaction = 2.796 g of zinc + 2.414 g of sulfur = 5.210 g Mass after the reaction = 4.169 g of zinc sulfide + 1.041 g of sulfur = 5.210 g Yes, the experiment obeys the law of conservation of mass.

23. Mass before the reaction = 1.00 g of zinc + 0.80 g of sulfur = 1.80 g Mass after the reaction = 1.50 g zinc sulfide

Total mass before 1.80 g – mass after 1.50 = 0.30 g of sulfur (answer b)

24. None of the three methods will destroy the atoms. Incineration is a chemical change.

25. (a) $\frac{\text{2.02 g hydrogen}}{\text{18.02 g water}}$ x 775 g water = 86.9 g hydrogen

(b) 118.026.00 g oxygeng water x 775 g water = 688 g oxygen

26. $\frac{\text{2.02 g hydrogen}}{\text{18.02 g water}}$ x ? kg water = 125 kg hydrogen

? kg water = 125 kg hydrogen x $\frac{\text{18.02 g hydrogen}}{\text{2.02 g water}}$ = 1120 kg water

27. $\frac{\text{3.0 parts carbon}}{\text{3.0 + 8.0 parts carbon dioxide}}$ x 14 kg carbon dioxide = 3.8 kg carbon

28. $\frac{\text{3.0 + 8.0 parts carbon dioxide}}{\text{3.0 parts carbon}}$ x 0.4 g diamond = 2 g carbon dioxide

 Chapter 2: Atoms

29. Statement number two of Dalton's atomic theory says, "All atoms of a given element are alike, but atoms of one element differ from atoms of any other element." (a) The different masses of atoms of calcium and vanadium support Dalton's atomic theory. (b) The same mass of two different elements violates statement two of Dalton's atomic theory. (Dalton did not predict the existence of isotopes.)

30. Yes, it agrees with statement number two of Dalton's atomic theory, which says, "All atoms of a given element are alike, but atoms of one element differ from atoms of any other element."

31. The atoms are neither being created nor destroyed; they are being rearranged.

32. a

$$\frac{8 \text{ units O}}{1 \text{ units H}} \times \frac{2 \text{ atoms H}}{1 \text{ atom O}} = 16$$

33. d;

34. a; $\frac{57 \text{ unit F}}{14 \text{ units N}} \times \frac{1 \text{ atom N}}{3 \text{ atoms F}} = \frac{19}{14}$ by mass

35. Assume that there is 1 g of carbon suboxide and that the mass ratio is 0.5296:0.4704
Divide both by 0.5296 to get a mass ratio of 1.0000:0.8882 in carbon suboxide. The ratio of oxygen in all three compounds is 1:2:3.

36. 8.0:85.5

37. Calculate the mass ratio of Sn to O in SnO. 0.742 g Sn / 0.100 g O = 7.4:1. Calculate the mass ratio of Sn to O in the unknown. 0.555 g Sn / 0.150 g O = 3.7:1 or 7.4:2. Compare the oxygen mass ratios of the two oxides. 1:2; thus the second oxide has 2 oxygen atoms for every oxygen atom in the first oxide. The formula of the second oxide is SnO_2.

38. Divide all oxygen and nitrogen ratios by the smallest oxygen to nitrogen ratio to obtain X 4.00:1.75, Y 2.00:1.75, and Z 1.00:1.75.

39. Calculate the ratio of Fe in hematite to the ratio of Fe in wusitite: 0.558g/0.558 g =1 Calculate the ratio of O in hematite to the ratio of O in wusitite: 0.240g/0.160 g =1.5 Multiply both by 2 to get whole numbers and you obtain Fe_2O_3 for hematite.
Calculate the ratio of Fe in magnetite to the ratio of Fe in wusitite: 3.35g/0.558 g =6 Calculate the ratio of O in magnetite to the ratio of O in wusitite: 1.28g/0.160 g =8
Reduce to the smallest whole numbers and you obtain Fe_3O_4 for magnetite. Yes.

40. Divide both carbon and hydrogen percentages by the carbon percent in each compound to obtain C : H ratios V 1.00:0.25 and W 1.00:0.20. The dividing by the smallest hydrogen amount yields hydrogen ratios of 1.00:1.25.

41. 1.000 g C x $\frac{1 \text{ molC}}{12.01 \text{ g C}}$ = 0.08326 mol C x $\frac{6.02\text{x}10^{23} \text{ atomsC}}{1 \text{ mol C}}$ = $5.01\text{x}10^{22}$ atoms C

42. 16.00 g/mol O + 2(1.01g/mol H) = 18.02 g/mol H_2O
6.02×10^{23} atoms of O and 2(6.02×10^{23} atoms of H) = 1.20×10^{24} atoms of H

43. (5 mol)(12.00 g/mol) = 60.00 g C-12. (5 mol)(12.01 g/mol) = 60.05 g C.
(To one significant figure, the correct number of significant figures. To use, both answers are 60 g.)

44. 1.01/12.00 and 1.01/12.01

45. Sample1 : $\frac{0.937 \text{ g C}}{1.000 \text{ g}}$ ×100% = 93.7% C; $\frac{0.0629 \text{ gH}}{1.000 \text{ g}}$ ×100% = 6.29% H

Sample 2 : $\frac{0.229 \text{ g C}}{0.244 \text{ g}}$ ×100% = 93.8% C; $\frac{0.0153 \text{ g H}}{0.244 \text{ g}}$ ×100% = 6.27% H

Sample 3 : $\frac{0.094 \text{ g C}}{0.100 \text{ g}}$ ×100% = 94% C; $\frac{0.0063 \text{ g H}}{0.100 \text{ g}}$ ×100% = 6.3% H

All ratios are constant to two significant figures. The solid is a pure compound!

46. Sample1 : $\frac{0.862 \text{ g C}}{1.000 \text{ g}}$ ×100% = 86.2% C; $\frac{0.138 \text{ g H}}{1.000 \text{ g}}$ ×100% = 13.8% H

Sample 2 : $\frac{1.295 \text{ g C}}{1.549 \text{ g}}$ ×100% = 83.6% C; $\frac{0.254 \text{ g H}}{1.549 \text{ g}}$ ×100% = 16.4% H

Sample 3 : $\frac{0.826 \text{ g C}}{0.988 \text{ g}}$ ×100% = 83.6% C; $\frac{0.162 \text{ g H}}{0.988 \text{ g}}$ ×100% = 16.4% H

Sample 1 does not agree with samples 2 and 3. The liquid is a mixture.

 Chapter 2: Atoms

47. $$\frac{1 \text{ part by mass F}}{2.09 \text{ part by mass U}} \times \frac{238}{19} = \frac{6 \text{ F}}{1 \text{ U}}$$

48. From the first experiment we note that 3.06 g of hydrogen produced 27.35 g of water when allowed to react with oxygen. By applying the law of conservation of mass, 24.29 g of oxygen is required. Next, calculate the ratio by mass of hydrogen to oxygen in water. 3.06 g hydrogen / 24.29 g oxygen = 1 part hydrogen to 8 parts oxygen. Next, compare the ratio (1:8) with the masses of hydrogen and oxygen produced in the electrolysis experiment: 1.45 g hydrogen / 11.51 g oxygen = 1 part hydrogen to 8 parts oxygen. The results are consistent with the law of definite proportions.

49. From the first experiment we note that 0.312 g of sulfur produced 0.623 g of sulfur dioxide when allowed to react with oxygen. By applying the law of conservation of mass, 0.311 g of oxygen is required. Next, calculate the ratio by mass of sulfur to oxygen in sulfur dioxide: 0.312 g sulfur/0.311 g oxygen = 1 part sulfur to 1 part oxygen. The 1.305 g of sulfur in the second experiment requires 1.305 g of oxygen (1:1 by mass); thus, 2.610 g of sulfur dioxide is produced in the second experiment.

50. $$\frac{92.61 \text{ g mercury}}{100.00 \text{ g mercuric oxide}} = 0.9261\text{g mercury/g mercuric oxide}$$

0.9261 x (X g mercuric oxide) = 100.00 g mercury

$$\text{X g mercuric oxide} = \frac{100.00 \text{ g mercury}}{0.9261 \text{ g mercury/g mercuric oxide}}$$

X g mercuric oxide = 108.0 g mercuric oxide

51. Table 2.1 indicates 1 ratio of 1.000 g oxygen:0.4375 g nitrogen. The nitrogen ratio is 0.4375:0.5836 or 1.000:1.334 or in whole numbers 3:4.

 Chapter 2: Atoms

52. When gasoline is burned, it combines with oxygen from the atmosphere to form carbon dioxide.

53. The law of conservation of mass. The mass of the products is the same as the mass of the reactants.

54. a) $\frac{14.01\text{ g N}}{3.03\text{ g H}} = \frac{4.62\text{ g N}}{1.00\text{ g H}}$; $\frac{28.00\text{ g N}}{6.05\text{ g H}} = \frac{4.62\text{ g N}}{1.00\text{ g H}}$

The N and H are in the same ratios; thus 34.05 g of NH_3 will be produced.

b) $\frac{4.62\text{ g N}}{1.00\text{ g H}} = \frac{X\text{ g N}}{11.0\text{ g H}}$; $X = \frac{(4.62\text{ g N})(11.0\text{ g H})}{1.00\text{ g H}} = 50.8$ g N ; 61.8 g NH_3 produced

c) $\frac{4.62\text{ g N}}{1.00\text{ g H}} = \frac{X\text{ g N}}{2.0\text{ g H}}$; $X = \frac{(4.62\text{ g N})(2.0\text{ g H})}{1.00\text{ g H}} = 9.2$ g N required

Remaining N = 11.0 g – 9.2 g = 1.8 g N.

55. a) Using Mendeleev's masses for S and Te: (32+125)/2=78.5
b) Using values from the modern periodic table: (32.06+127.60)/2=79.83

56. a) neither b) rare c) hazardous d) neither e) hazardous

57. Titanium dioxide replacing lead in paints, metal salts replacing lead as stabilizers in plastics, and lead-free adhesives used in electronics.

58. Mercury is toxic and can leach from landfills into groundwater.