

13 These calculations have used $L = 6.02 \times 10^{23}$

(a) 7.22×10^{22} (b) 3.01×10^{23}

(c) 1.20×10^{23}

14 0.53 mol H

15 0.250 mol

16 (a) 262.87 g mol⁻¹ (b) 176.14 g mol⁻¹

(c) 164.10 g mol⁻¹ (d) 248.22 g mol⁻¹

17 189.1 g

18 1.5 mol

19 0.0074 mol Cl⁻

20 1.83×10^{23} C atoms

21 171 g (integer value because no calculator)

22 10.0 g H₂O

23 2.0 mol N₂ > 3.0 mol NH₃ > 25.0 mol H₂ > 1.0 mol N₂H₄

24 (a) CH (b) CH₂O

(c) C₁₀H₁₂O₁₁ (d) C₃H₆

(e) C₃H₆ (f) CH₂O

25 Na₂S₂O₅

26 CoSO₄·7H₂O

27 C₁₁H₁₅N

28 NH₃

29 6.94 L

30 CdS

31 empirical formula CH; molecular formula C₇H₈

32 empirical formula H₂PO₃; molecular formula H₄P₂O₆

33 C₁₀H₁₂N₂P₂O₁₁ for both empirical and molecular formulas

34 C₂H₄O

35 Let y = mass of chalk in grams.

moles of chalk used = $\frac{\text{mass used}}{M_r(\text{CaCO}_3)} \times$

$y \frac{\text{g}}{100.09 \text{ g mol}^{-1}}$

This is the same as the number of moles of carbon atoms used.

Therefore the number of carbon atoms used =

moles of chalk $\times$ (6.02 $\times$ 10²³ mol⁻¹) =

$\frac{6.02 \times 10^{23} y}{100.09}$

36 (a) 2.50 mol

(b) 5.63 mol

(c) 665.5 g

37 (a) $2\text{C}_2\text{H}_6 + 13\text{O}_2 \rightarrow 8\text{CO}_2 + 10\text{H}_2\text{O}$

(b) 1.59 g

38 4.355 kg

39 (a) $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$

(b) 92.8%

(c) CaCO₃ is the only source of CO₂; all the CaCO₃ undergoes complete decomposition; all CO₂ released is captured; heating does not cause any change in the mass of the other minerals present.

40 (a) 65.2 g

(b) 1.3 g H₂

41 5.23 g C₂H₅Cl₂

42 253.2 g theoretical CaSO₄; 77.9%

43 3.16 g ester

44 107 g of C₂H₆ needed

45 (a) 2.40 mol

(b) 0.0110 mol

(c) 44 mol

46 (a) 35.65 dm³

(b) 5.7 dm³

47 0.652 dm³

48 0.138 mol Br₂ and 0.156 mol Cl₂, so more molecules of Cl₂

49 0.113 dm³

50 0.28 dm³

51 90 kPa

52 16 °C

53 3.0 dm³

54 2.8 dm³

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Wade Hampton Frost



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1. Understanding the eBook Answers For Pearson Standardized Preparation Workbook Chemistry
 - The Rise of Digital Reading Answers For Pearson Standardized Preparation Workbook Chemistry
 - Advantages of eBooks Over Traditional Books
2. Identifying Answers For Pearson Standardized Preparation Workbook Chemistry
 - Exploring Different Genres
 - Considering Fiction vs. Non-Fiction
 - Determining Your Reading Goals
3. Choosing the Right eBook Platform
 - Popular eBook Platforms
 - Features to Look for in an Answers For Pearson Standardized Preparation Workbook Chemistry
 - User-Friendly Interface
4. Exploring eBook Recommendations from Answers For Pearson Standardized Preparation Workbook Chemistry
 - Personalized Recommendations
 - Answers For Pearson Standardized Preparation Workbook Chemistry User Reviews and Ratings
 - Answers For Pearson Standardized Preparation Workbook Chemistry and Bestseller Lists
5. Accessing Answers For Pearson Standardized Preparation Workbook Chemistry Free and Paid eBooks
 - Answers For Pearson Standardized Preparation Workbook Chemistry Public Domain eBooks
 - Answers For Pearson Standardized Preparation Workbook Chemistry eBook Subscription Services
 - Answers For Pearson Standardized Preparation Workbook Chemistry Budget-Friendly Options

6. Navigating Answers For Pearson Standardized Preparation Workbook Chemistry eBook Formats
 - ePub, PDF, MOBI, and More
 - Answers For Pearson Standardized Preparation Workbook Chemistry Compatibility with Devices
 - Answers For Pearson Standardized Preparation Workbook Chemistry Enhanced eBook Features
7. Enhancing Your Reading Experience
 - Adjustable Fonts and Text Sizes of Answers For Pearson Standardized Preparation Workbook Chemistry
 - Highlighting and Note-Taking Answers For Pearson Standardized Preparation Workbook Chemistry
 - Interactive Elements Answers For Pearson Standardized Preparation Workbook Chemistry
8. Staying Engaged with Answers For Pearson Standardized Preparation Workbook Chemistry
 - Joining Online Reading Communities
 - Participating in Virtual Book Clubs
 - Following Authors and Publishers Answers For Pearson Standardized Preparation Workbook Chemistry
9. Balancing eBooks and Physical Books Answers For Pearson Standardized Preparation Workbook Chemistry
 - Benefits of a Digital Library
 - Creating a Diverse Reading Collection Answers For Pearson Standardized Preparation Workbook Chemistry
10. Overcoming Reading Challenges
 - Dealing with Digital Eye Strain
 - Minimizing Distractions
 - Managing Screen Time
11. Cultivating a Reading Routine Answers For Pearson Standardized Preparation Workbook Chemistry
 - Setting Reading Goals Answers For Pearson Standardized Preparation Workbook Chemistry
 - Carving Out Dedicated Reading Time
12. Sourcing Reliable Information of Answers For Pearson Standardized Preparation Workbook Chemistry
 - Fact-Checking eBook Content of Answers For Pearson Standardized Preparation Workbook Chemistry
 - Distinguishing Credible Sources
13. Promoting Lifelong Learning
 - Utilizing eBooks for Skill Development
 - Exploring Educational eBooks
14. Embracing eBook Trends
 - Integration of Multimedia Elements

- Interactive and Gamified eBooks

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