

- ④ To find the molecular formula, multiply the empirical formula by the ratio of molar mass divided by empirical mass. In this example it is $66.0/33.0$, which equals 2.
- 4 a Calcium ions are $2+$ and chloride ions are $1-$, so the formula is CaCl_2 .
 b Silver ions are $1+$ and nitrate ions are $1-$, so the formula is AgNO_3 .
 c Copper(II) is Cu^{2+} and phosphate is PO_4^{3-} . For a neutral compound you need three copper ions and two phosphate ions, making charges of $+6$ and -6 . Therefore, the formula is $\text{Cu}_3(\text{PO}_4)_2$.
 d Aluminium is $3+$ and oxide is $2-$, so aluminium oxide is Al_2O_3 .
 ⑤ There are two ways of working out formulae of ionic compounds. One way is to make sure that the total positive charge equals the total negative charge (see part c). The other way, if the charges are not the same, is to swap them round. Copper is $2+$ and phosphate $3-$, so there are three copper ions and two phosphate ions in the formula.
- 5 a As there is one manganese ion to each carbonate ion, the charges have the same numerical value. Carbonate ions are $2-$, so the manganese ions in this compound must be $2+$.
 b The charge on two vanadium ions must equal the charge on three sulfate ions. Each sulfate is -2 , so three sulfate ions have a total charge of -6 . This means that two vanadium ions have a total charge of $+6$, so each has a charge of $+3$.
- 6 a molar mass of $\text{Ca}(\text{OH})_2 = 40.1 + 2 \times (16.0 + 1.0) = 74.1 \text{ g mol}^{-1}$
 b molar mass of $\text{Al}_2(\text{SO}_4)_3 = (2 \times 27.0) + 3 \times [32.1 + (4 \times 16.0)] = 342.3 \text{ g mol}^{-1}$
 c molar mass of $\text{FeSO}_4 \cdot 7\text{H}_2\text{O} = 55.8 + 32.1 + (4 \times 16.0) + 7 \times (2.0 + 16.0) = 277.9 \text{ g mol}^{-1}$
 ⑥ In calculating molar mass, make sure that you use the relative atomic masses (the larger number), not the atomic numbers, from the periodic table. Molar mass has the units of g mol^{-1} . Relative molecular mass has no units. The numbers are the same.
- 7 The balanced equations are:
 a $\text{P}_4 + 5\text{O}_2 \rightarrow \text{P}_4\text{O}_{10}$ or $2\text{P}_2\text{O}_5$
 ⑥ Phosphorus(V) oxide exists as the dimer P_4O_{10} , but $2\text{P}_2\text{O}_5$ on the right of the equation would be acceptable.
 b $4\text{Al} + 3\text{O}_2 \rightarrow 2\text{Al}_2\text{O}_3$
 c $2\text{Mg}(\text{NO}_3)_2 \rightarrow 2\text{MgO} + 4\text{NO}_2 + \text{O}_2$
 ⑥ In part c, balance the oxygen atoms last. Remember that the number of oxygen atoms on the right-hand side of the equation must be an even number, because $\text{Mg}(\text{NO}_3)_2$ contains an even number of oxygen atoms.
 d $2\text{LiOH} + \text{H}_2\text{SO}_4 \rightarrow \text{Li}_2\text{SO}_4 + 2\text{H}_2\text{O}$
 e $2\text{Fe}^{2+} + 2\text{I}^- \rightarrow 2\text{Fe}^{3+} + \text{I}_2$
 ⑥ Make sure that this equation balances for charge. Having one Fe^{2+} ion on the left-hand side and one Fe^{3+} on the right-hand side balances the number of Fe ions, but the charge on the left would not equal the charge on the right.
 f $2\text{C}_8\text{H}_{18} + 25\text{O}_2 \rightarrow 16\text{CO}_2 + 18\text{H}_2\text{O}$ or $\text{C}_8\text{H}_{18} + 12\frac{1}{2}\text{O}_2 \rightarrow 8\text{CO}_2 + 9\text{H}_2\text{O}$
 ⑥ The only source of oxygen on the left-hand side of the equation is O_2 , so the number of oxygen atoms on the right-hand side must be an even number. Therefore, there must be an even number of H_2O molecules on the right. This requires two C_8H_{18} molecules on the left.
- 8 For equations a–c, the simplest way is to realise that the component ions of the compound combine to give the precipitate. All other ions are spectators.

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Table of Contents Teaching Transparency Answers Key For

1. Understanding the eBook Teaching Transparency Answers Key For
 - The Rise of Digital Reading Teaching Transparency Answers Key For
 - Advantages of eBooks Over Traditional Books
2. Identifying Teaching Transparency Answers Key For
 - 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 Teaching Transparency Answers Key For
 - User-Friendly Interface
4. Exploring eBook Recommendations from Teaching Transparency Answers Key For

- Personalized Recommendations
- Teaching Transparency Answers Key For User Reviews and Ratings
- Teaching Transparency Answers Key For and Bestseller Lists
- 5. Accessing Teaching Transparency Answers Key For Free and Paid eBooks
 - Teaching Transparency Answers Key For Public Domain eBooks
 - Teaching Transparency Answers Key For eBook Subscription Services
 - Teaching Transparency Answers Key For Budget-Friendly Options
- 6. Navigating Teaching Transparency Answers Key For eBook Formats
 - ePub, PDF, MOBI, and More
 - Teaching Transparency Answers Key For Compatibility with Devices
 - Teaching Transparency Answers Key For Enhanced eBook Features
- 7. Enhancing Your Reading Experience
 - Adjustable Fonts and Text Sizes of Teaching Transparency Answers Key For
 - Highlighting and Note-Taking Teaching Transparency Answers Key For
 - Interactive Elements Teaching Transparency Answers Key For
- 8. Staying Engaged with Teaching Transparency Answers Key For
 - Joining Online Reading Communities
 - Participating in Virtual Book Clubs
 - Following Authors and Publishers Teaching Transparency Answers Key For
- 9. Balancing eBooks and Physical Books Teaching Transparency Answers Key For
 - Benefits of a Digital Library
 - Creating a Diverse Reading Collection Teaching Transparency Answers Key For
- 10. Overcoming Reading Challenges
 - Dealing with Digital Eye Strain
 - Minimizing Distractions
 - Managing Screen Time
- 11. Cultivating a Reading Routine Teaching Transparency Answers Key For
 - Setting Reading Goals Teaching Transparency Answers Key For
 - Carving Out Dedicated Reading Time
- 12. Sourcing Reliable Information of Teaching Transparency Answers Key For

- Fact-Checking eBook Content of Teaching Transparency Answers Key For
- 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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