

- 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  $\text{CaCl}_2$ .  
 b Silver ions are  $1+$  and nitrate ions are  $1-$ , so the formula is  $\text{AgNO}_3$ .  
 c Copper(II) is  $\text{Cu}^{2+}$  and phosphate is  $\text{PO}_4^{3-}$ . 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  $\text{Al}_2\text{O}_3$ .  
 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  $\text{g mol}^{-1}$ . 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  $\text{P}_4\text{O}_{10}$ , 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  $\text{Fe}^{2+}$  ion on the left-hand side and one  $\text{Fe}^{3+}$  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  $2\text{C}_6\text{H}_{14} + 25\text{O}_2 \rightarrow 16\text{CO}_2 + 18\text{H}_2\text{O}$  or  $\text{C}_6\text{H}_{14} + 12\frac{1}{2}\text{O}_2 \rightarrow 8\text{CO}_2 + 9\text{H}_2\text{O}$   
 ⑥ The only source of oxygen on the left-hand side of the equation is  $\text{O}_2$ , so the number of oxygen atoms on the right-hand side must be an even number. Therefore, there must be an even number of  $\text{H}_2\text{O}$  molecules on the right. This requires two  $\text{C}_6\text{H}_{14}$  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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