

Chapter 3

Aqueous Reactions and Solutions Stoichiometry

Visualising Concepts

- 3.1 Reactant A = blue, reactant B = red
Overall, 4 blue A₂ molecules + 4 red B atoms → 4 A₂B molecules
Since 4 is a common factor, this equation reduces to equation (a).
- 3.2 (a) There are twice as many O atoms as N atoms, so the empirical formula of the original compound is NO₂.
(b) No, because we have no way of knowing whether the empirical and molecular formulae are the same. NO₂ represents the simplest ratio of atoms in a molecule but not the only possible molecular formula.
- 3.3 The box contains 4 C atoms (black) and 16 H atoms (white), so the empirical formula of the hydrocarbon is CH₄.
- 3.4 (a) Given the molecular model, write the molecular formula.
Use the colours of the atoms (spheres) in the model to determine the number of atoms of each element.
Observe 2 black C atoms, 5 white H atoms, 1 blue N atom, 2 red O atoms.
C₂H₅NO₂
(b) Follow the method in SAMPLE EXERCISE 3.5. Calculate formula mass in u and molar mass in grams.

2 C atoms	=	2 (12.0 u)	=	24.0 u
5 H atoms	=	5 (1.0 u)	=	5.0 u
1 N atom	=	1 (14.0 u)	=	14.0 u
2 O atoms	=	2 (16.0 u)	=	<u>32.0 u</u>
				75.0 u

Formula mass = 75.0 u, molar mass = 75.0 g mol⁻¹

(c) Use the definition of mass% and the results from parts (a) and (b) above to find mass% N in glycine.

$$\text{mass\% N} = \frac{\text{g N}}{\text{g C}_2\text{H}_5\text{NO}_2} \times 100$$

Assume 1 mol C₂H₅NO₂. From the molecular formula of glycine [part (a)], there is 1 mol N per mol glycine.

$$\text{mass\% N} = \frac{1 \times (\text{molar mass N})}{\text{molar mass glycine}} \times 100 = \frac{14.0 \text{ g}}{75.0 \text{ g}} \times 100 = 18.7\%$$

3.5 Given: 4.0 mol CH₄. Find: mol CO and mol H₂

Examine the boxes to determine the CH₄:CO mol ratio and CH₄:H₂O mole ratio.

There are 2 CH₄ molecules in the reactant box and 2 CO molecules in the product box. The mole ratio is 2:2 or 1:1. Therefore, 4.0 mol CH₄ can produce 4.0 mol CO.

There are 2 CH₄ molecules in the reactant box and 6 H₂ molecules in the product box. The mole ratio is 2:6 or 1:3. So, 4.0 mol CH₄ can produce 12.0 mol H₂.

Check. Use proportions. 2 mol CH₄/2 mol CO = 4 mol CH₄/4 mol CO;

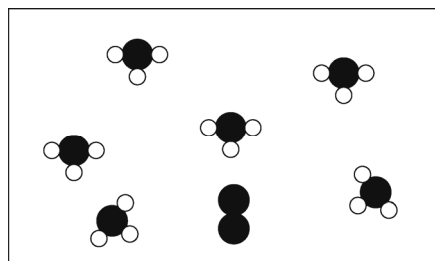
2 mol CH₄/6 mol H₂ = 4 mol CH₄/12 mol H₂.

3.6 Given a box diagram and formulae of reactants, draw a box diagram of products.

Write and balance the chemical equation. Determine combining ratios of elements and decide on limiting reactant. Draw a box diagram of products, containing the correct number of product molecules and only excess reactant.



Each N atom (1/2 of an N₂ molecule) reacts with 3 H atoms (1.5 H₂ molecules) to form an NH₃ molecule. Eight N atoms (4 N₂ molecules) require 24 H atoms (12 H₂ molecules) for a complete reaction. Only 9 H₂ molecules are available, so H₂ is the limiting reactant. Nine H₂ molecules (18 H atoms) determine that 6 NH₃ molecules are produced. One N₂ molecule is in excess.



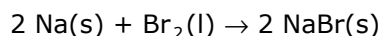
Check. Verify that mass is conserved in your solution, that the number and kinds of atoms are the same in reactant and product diagrams. In this example, there are 8 N atoms and 18 H atoms in both diagrams, so mass is conserved.

Balancing Chemical Equations

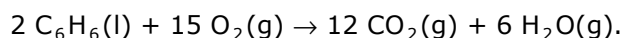
- 3.7 (a) In balancing chemical equations, the law of conservation of mass, that atoms are neither created nor destroyed during the course of a reaction, is observed. This means that the **number** and **kinds** of atoms on both sides of the chemical equation must be the same.
- (b) Subscripts in chemical formulae should not be changed when balancing equations, because changing the subscript changes the identity of the compound (law of constant composition).
- (c) gases, (g); liquids, (l); solids, (s); aqueous solutions, (aq)
- 3.8 (a) In a CO molecule, there is one O atom bound to C. 2 CO indicates that there are **two CO molecules**, each of which contains one C and one O atom. Adding a subscript 2 to CO to form CO₂ means that there are **two O atoms** bound to one C in a CO₂ molecule. The composition of the different molecules, CO₂ and CO, is different and the physical and chemical properties of the two compounds they constitute are very different. The subscript 2 changes molecular composition and thus properties of the compound. The prefix 2 indicates how many molecules (or moles) of the original compound are under consideration.
- (b) Yes. There are the same number and kinds of atoms on the reactants side and the products side of the equation.
- 3.9 (a) $2 \text{ CO(g)} + \text{O}_2\text{(g)} \rightarrow 2 \text{ CO}_2\text{(g)}$
- (b) $\text{N}_2\text{O}_5\text{(g)} + \text{H}_2\text{O(l)} \rightarrow 2 \text{ HNO}_3\text{(aq)}$
- (c) $\text{CH}_4\text{(g)} + 4 \text{ Cl}_2\text{(g)} \rightarrow \text{CCl}_4\text{(l)} + 4 \text{ HCl(g)}$
- (d) $\text{Al}_4\text{C}_3\text{(s)} + 12 \text{ H}_2\text{O(l)} \rightarrow 4 \text{ Al(OH)}_3\text{(s)} + 3 \text{ CH}_4\text{(g)}$
- (e) $2 \text{ C}_5\text{H}_{10}\text{O}_2\text{(l)} + 13 \text{ O}_2\text{(g)} \rightarrow 10 \text{ CO}_2\text{(g)} + 10 \text{ H}_2\text{O(g)}$
- (f) $2 \text{ Fe(OH)}_3\text{(s)} + 3 \text{ H}_2\text{SO}_4\text{(aq)} \rightarrow \text{Fe}_2\text{(SO}_4)_3\text{(aq)} + 6 \text{ H}_2\text{O(l)}$
- (g) $\text{Mg}_3\text{N}_2\text{(s)} + 4 \text{ H}_2\text{SO}_4\text{(aq)} \rightarrow 3 \text{ MgSO}_4\text{(aq)} + (\text{NH}_4)_2\text{SO}_4\text{(aq)}$
- 3.10 (a) $\text{CaC}_2\text{(s)} + 2 \text{ H}_2\text{O(l)} \rightarrow \text{Ca(OH)}_2\text{(aq)} + \text{C}_2\text{H}_2\text{(g)}$
- (b) $2 \text{ KClO}_3\text{(s)} \xrightarrow{\Delta} 2 \text{ KCl(s)} + 3 \text{ O}_2\text{(g)}$
- (c) $\text{Zn(s)} + \text{H}_2\text{SO}_4\text{(aq)} \rightarrow \text{H}_2\text{(g)} + \text{ZnSO}_4\text{(aq)}$
- (d) $\text{PCl}_3\text{(l)} + 3 \text{ H}_2\text{O(l)} \rightarrow \text{H}_3\text{PO}_3\text{(aq)} + 3 \text{ HCl(aq)}$
- (e) $3 \text{ H}_2\text{S(g)} + 2 \text{ Fe(OH)}_3\text{(s)} \rightarrow \text{Fe}_2\text{S}_3\text{(s)} + 6 \text{ H}_2\text{O(g)}$

Patterns of Chemical Reactivity

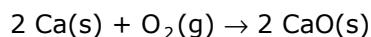
- 3.11 (a) When a metal reacts with a non-metal, an ionic compound forms. The combining ratio of the atoms is such that the total positive charge on the metal cation(s) is equal to the total negative charge on the non-metal anion(s). All simple ionic compounds are solids at room temperature.



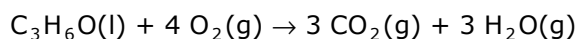
- (b) The second reactant is oxygen gas from the air, $\text{O}_2(\text{g})$. The products are $\text{CO}_2(\text{g})$ and $\text{H}_2\text{O}(\text{g})$.



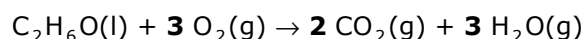
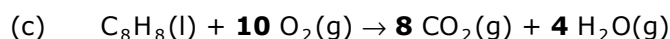
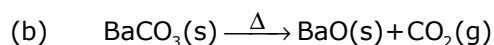
- 3.12 (a) Neutral Ca atom loses 2e^- to form Ca^{2+} . Neutral O_2 molecule gains 4e^- to form 2O^{2-} . The formula of the product will be CaO, because the cationic and anionic charges are opposite and equal.



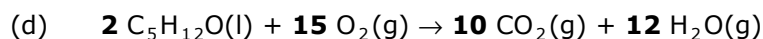
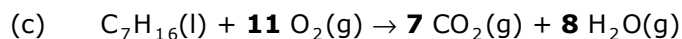
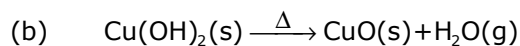
- (b) The products are $\text{CO}_2(\text{g})$ and $\text{H}_2\text{O}(\text{g})$.



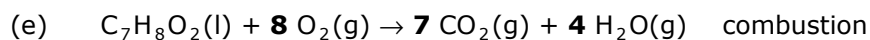
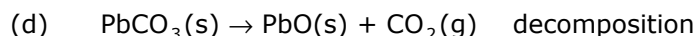
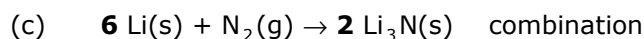
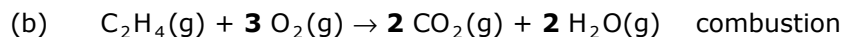
- 3.13 (a) $\text{Mg(s)} + \text{Cl}_2(\text{g}) \rightarrow \text{MgCl}_2(\text{s})$



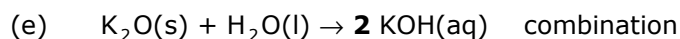
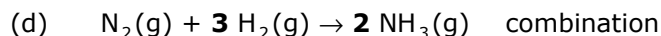
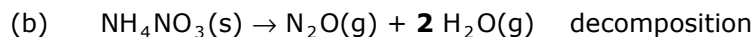
- 3.14 (a) $4 \text{Al(s)} + 3 \text{O}_2(\text{g}) \rightarrow 2 \text{Al}_2\text{O}_3(\text{s})$



- 3.15 (a) $2 \text{Al(s)} + 3 \text{Cl}_2(\text{g}) \rightarrow 2 \text{AlCl}_3(\text{s})$ combination



- 3.16 (a) $2 \text{C}_3\text{H}_6(\text{g}) + 9 \text{O}_2(\text{g}) \rightarrow 6 \text{CO}_2(\text{g}) + 6 \text{H}_2\text{O(g)}$ combustion



Formula Mass

3.17 Given molecular formula or name, calculate formula mass.

If a name is given, write the correct molecular formula. Then, follow the method in SAMPLE EXERCISE 3.5.

- (a) N_2O_5 : $2(14.0) + 5(16.0) = 108.0 \text{ u}$
- (b) CuSO_4 : $1(63.6) + 1(32.1) + 4(16.0) = 159.6 \text{ u}$
- (c) $(\text{NH}_4)_3\text{PO}_4$: $3(14.0) + 12(1.0) + 1(31.0) + 4(16.0) = 149.0 \text{ u}$
- (d) $\text{Ca}(\text{HCO}_3)_2$: $1(40.1) + 2(1.0) + 2(12.0) + 6(16.0) = 162.1 \text{ u}$
- (e) Al_2S_3 : $2(27.0) + 3(32.1) = 150.3 \text{ u}$
- (f) $\text{Fe}_2(\text{SO}_4)_3$: $2(55.8) + 3(32.1) + 12(16.0) = 399.9 \text{ u}$
- (g) Si_2Br_6 : $2(28.1) + 6(79.9) = 535.6 \text{ u}$

3.18 Formula mass in u to 1 decimal place.

- (a) N_2O : $2(14.0) + 1(16.0) = 44.0 \text{ u}$
- (b) $\text{C}_7\text{H}_6\text{O}_2$: $7(12.0) + 6(1.0) + 2(16.0) = 122.0 \text{ u}$
- (c) $\text{Mg}(\text{OH})_2$: $1(24.3) + 2(16.0) + 2(1.0) = 58.3 \text{ u}$
- (d) $(\text{NH}_2)_2\text{CO}$: $2(14.0) + 4(1.0) + 1(12.0) + 1(16.0) = 60.0 \text{ u}$
- (e) $\text{CH}_3\text{CO}_2\text{C}_5\text{H}_{11}$: $7(12.0) + 14(1.0) + 2(16.0) = 130.0 \text{ u}$

3.19 (a) C_2H_2 : FM = $2(12.0) + 2(1.0) = 26.0 \text{ u}$

$$\% \text{ C} = \frac{2(12.0) \text{ u}}{26.0 \text{ u}} \times 100 = 92.3\%$$

(b) $\text{HC}_6\text{H}_7\text{O}_6$: FM = $6(12.0) + 8(1.0) + 6(16.0) = 176.0 \text{ u}$

$$\% \text{ H} = \frac{8(1.0) \text{ u}}{176.0 \text{ u}} \times 100 = 4.5\%$$

(c) $(\text{NH}_4)_2\text{SO}_4$: FM = $2(14.0) + 8(1.0) + 1(32.1) + 4(16.0) = 132.1 \text{ u}$

$$\% \text{ H} = \frac{8(1.0) \text{ u}}{132.1 \text{ u}} \times 100 = 6.1\%$$

(d) $\text{PtCl}_2(\text{NH}_3)_2$: FM = $1(195.1) + 2(35.5) + 2(14.0) + 6(1.0) = 300.1 \text{ u}$

$$\% \text{ Pt} = \frac{1(195.1) \text{ u}}{300.1 \text{ u}} \times 100 = 65.01\%$$

(e) $\text{C}_{18}\text{H}_{24}\text{O}_2$: FM = $18(12.0) + 24(1.0) + 2(16.0) = 272.0 \text{ u}$

$$\% \text{ O} = \frac{2(16.0) \text{ u}}{272.0 \text{ u}} \times 100 = 11.8\%$$

$$(f) \quad \text{C}_{18}\text{H}_{27}\text{NO}_3: \text{FM} = 18(12.0) + 27(1.0) + 1(14.0) + 3(16.0) = 305.0 \text{ u}$$

$$\% \text{ C} = \frac{18(12.0) \text{ u}}{305.0 \text{ u}} \times 100 = 70.8\%$$

3.20 Follow the logic for calculating mass% C given in SAMPLE EXERCISE 3.6.

$$(a) \quad \text{C}_7\text{H}_6\text{O}: \text{FM} = 7(12.0) + 6(1.0) + 1(16.0) = 106.0 \text{ u}$$

$$\% \text{ C} = \frac{7(12.0) \text{ u}}{106.0 \text{ u}} \times 100 = 79.2\%$$

$$(b) \quad \text{C}_8\text{H}_8\text{O}_3: \text{FM} = 8(12.0) + 8(1.0) + 3(16.0) = 152.0 \text{ u}$$

$$\% \text{ C} = \frac{8(12.0) \text{ u}}{152.0 \text{ u}} \times 100 = 63.2\%$$

$$(c) \quad \text{C}_7\text{H}_{14}\text{O}_2: \text{FM} = 7(12.0) + 14(1.0) + 2(16.0) = 130.0 \text{ u}$$

$$\% \text{ C} = \frac{7(12.0) \text{ u}}{130.0 \text{ u}} \times 100 = 64.6\%$$

Avogadro's Number and the Mole

3.21 (a) 6.022×10^{23} . This is the number of objects in a mole of anything.

(b) The formula mass of a substance in u has the same numerical value as the molar mass expressed in grams.

3.22 (a) exactly 12 g (b) 6.0221421×10^{23} , Avogadro's number

3.23 Since the mole is a counting unit, use it as a basis of comparison; determine the total moles of atoms in each given quantity.

23 g Na contains 1 mol atoms

0.5 mol H_2O contains $(3 \text{ atoms} \times 0.5 \text{ mol}) = 1.5 \text{ mol atoms}$

6.0×10^{23} N_2 molecules contains $(2 \text{ atoms} \times 1 \text{ mol}) = 2 \text{ mol atoms}$

3.24 3.0×10^{23} H_2O_2 molecules contains $(4 \text{ atoms} \times 0.5 \text{ mol}) = 2 \text{ mol atoms}$

32 g O_2 contains $(2 \text{ atoms} \times 1 \text{ mol}) = 2 \text{ mol atoms}$

2.0 mol CH_4 contains $(5 \text{ atoms} \times 2 \text{ mol}) = 10 \text{ mol atoms}$

3.25 (a) Given: 0.773 mol CaH_2 . Find: mass in g.

Use molar mass (g mol^{-1}) of CaH_2 to find g CaH_2

molar mass = $1(40.08 \text{ g mol}^{-1}) + 2(1.008 \text{ g mol}^{-1})$

= $42.096 \text{ g mol}^{-1} = 42.10 \text{ g mol}^{-1} \text{ CaH}_2$

$0.773 \text{ mol CaH}_2 \times 42.096 \text{ g mol}^{-1} = 32.5 \text{ g CaH}_2$

Check. $3/4(40) = 30 \text{ g}$. The calculated result is reasonable.

- (b) Given: mass. Find: moles. Use molar mass of $\text{Mg}(\text{NO}_3)_2$.
 molar mass = $1(24.31 \text{ g mol}^{-1}) + 2(14.01 \text{ g mol}^{-1}) + 6(16.00 \text{ g mol}^{-1}) = 148.33 \text{ g mol}^{-1} = 148.3 \text{ g mol}^{-1}$

$$\frac{5.35 \text{ g Mg}(\text{NO}_3)_2}{148.33 \text{ g mol}^{-1}} = 0.0361 \text{ mol Mg}(\text{NO}_3)_2$$

Check. $5/150 \approx 1/30 = 0.033 \text{ mol}$

- (c) $0.0305 \text{ mol CH}_3\text{OH} \times 6.022 \times 10^{23} \text{ molecules mol}^{-1}$
 $= 1.8367 \times 10^{22} \text{ CH}_3\text{OH molecules}$
 $= 1.84 \times 10^{22} \text{ CH}_3\text{OH molecules}$

$$\text{Check. } (0.03 \times 6 \times 10^{23}) = 0.18 \times 10^{23} = 1.8 \times 10^{22}$$

- (d) $\text{mol C}_4\text{H}_{10} \rightarrow \text{mol C} \rightarrow \text{number of C atoms}$

$$0.585 \text{ mol C}_4\text{H}_{10} \times \frac{4 \text{ mol C}}{1 \text{ mol C}_4\text{H}_{10}} \times 6.022 \times 10^{23} \text{ atoms mol}^{-1}$$

$$= 1.41 \times 10^{24} \text{ C atoms}$$

$$\text{Check. } (0.6 \times 4 \times 6 \times 10^{23}) = 14 \times 10^{23} = 1.4 \times 10^{24}.$$

- 3.26 (a) molar mass = $1(137.33 \text{ g mol}^{-1}) + 2(126.904 \text{ g mol}^{-1})$
 $= 391.14 \text{ g mol}^{-1}$

$$1.906 \times 10^{-2} \text{ mol BaI}_2 \times 391.14 \text{ g mol}^{-1} = 7.455 \text{ g BaI}_2$$

- (b) molar mass = $1(14.01 \text{ g mol}^{-1}) + 4(1.008 \text{ g mol}^{-1}) + 1(35.45 \text{ g mol}^{-1}) = 53.49 \text{ g mol}^{-1}$

$$\frac{48.3 \text{ g NH}_4\text{Cl}}{53.49 \text{ g mol}^{-1}} = 0.903 \text{ mol NH}_4\text{Cl}$$

- (c) $0.05752 \text{ mol HCHO}_2 \times 6.02214 \times 10^{23} \text{ molecules mol}^{-1}$
 $= 3.464 \times 10^{22} \text{ HCHO}_2 \text{ molecules}$

- (d) $4.88 \times 10^{-3} \text{ mol Al}(\text{NO}_3)_3 \times \frac{9 \text{ mol O}}{1 \text{ mol Al}(\text{NO}_3)_3} \times 6.022 \times 10^{23} \text{ atoms mol}^{-1}$
 $= 2.64 \times 10^{22} \text{ O atoms}$

- 3.27 (a) $\text{C}_6\text{H}_{10}\text{OS}_2$

$$\text{molar mass} = 6(12.01 \text{ g mol}^{-1}) + 10(1.008 \text{ g mol}^{-1}) + 1(16.00 \text{ g mol}^{-1}) + 2(32.07 \text{ g mol}^{-1}) = 162.28 \text{ g mol}^{-1} = 162.3 \text{ g mol}^{-1}$$

- (b) $\text{mg} \rightarrow \text{g} \rightarrow \text{mol}$

$$\frac{5.00 \times 10^{-3} \text{ g allicin}}{162.3 \text{ g mol}^{-1}} = 3.081 \times 10^{-5} \text{ mol} = 3.08 \times 10^{-5} \text{ mol allicin}$$

Check. 5.00 mg is a small mass, so the small answer is reasonable.

$$(5 \times 10^{-3})/200 = 2.5 \times 10^{-5}$$

- (c) Use mol from part (b) and Avogadro's number to calculate molecules.

$$3.081 \times 10^{-5} \text{ mol allixin} \times 6.022 \times 10^{23} \text{ molecules mol}^{-1} \\ = 1.855 \times 10^{19} \text{ molecules} = 1.86 \times 10^{19} \text{ allixin molecules}$$

$$\text{Check. } (3 \times 10^{-5})(6 \times 10^{23}) = 18 \times 10^{18} = 1.8 \times 10^{19}$$

- (d) Use molecules from part (c) and molecular formula to calculate S atoms.

$$1.855 \times 10^{19} \text{ allixin molecules} \times \frac{2 \text{ S atoms}}{1 \text{ allixin molecule}} = 3.71 \times 10^{19} \text{ S atoms}$$

- 3.28 (a) Use molecular formula to determine number of H atoms that are present with 1.250×10^{21} C atoms.

$$\frac{12 \text{ H atoms}}{6 \text{ C atoms}} = \frac{2 \text{ H}}{1 \text{ C}} \times 1.250 \times 10^{21} \text{ C atoms} = 2.500 \times 10^{21} \text{ H atoms}$$

$$\text{Check. } (2 \times 1 \times 10^{21}) = 2 \times 10^{21}$$

- (b) Use molecular formula to find the number of glucose molecules that contain 1.250×10^{21} C atoms.

$$\frac{1 \text{ C}_6\text{H}_{12}\text{O}_6 \text{ molecule}}{6 \text{ C atoms}} \times 1.250 \times 10^{21} \text{ C atoms} \\ = 2.0833 \times 10^{20} \text{ C}_6\text{H}_{12}\text{O}_6 \text{ molecules} = 2.083 \times 10^{20} \text{ C}_6\text{H}_{12}\text{O}_6 \text{ molecules}$$

$$\text{Check. } (12 \times 10^{20}/6) = 2 \times 10^{20}$$

- (c) Use Avogadro's number to change molecules \rightarrow mol.

$$\frac{2.0833 \times 10^{20} \text{ C}_6\text{H}_{12}\text{O}_6 \text{ molecules}}{6.022 \times 10^{23} \text{ molecules mol}^{-1}} = 3.4595 \times 10^{-4} \text{ mol} \\ = 3.460 \times 10^{-4} \text{ mol C}_6\text{H}_{12}\text{O}_6$$

$$\text{Check. } (2 \times 10^{20})/(6 \times 10^{23}) = 0.33 \times 10^{-3} = 3.3 \times 10^{-4}$$

- (d) Use molar mass to change mol \rightarrow g.

$$1 \text{ mol of C}_6\text{H}_{12}\text{O}_6 \text{ weighs } 180.0 \text{ g (SAMPLE EXERCISE 3.9)}$$

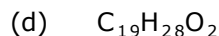
$$3.4595 \times 10^{-4} \text{ mol C}_6\text{H}_{12}\text{O}_6 \times 180.0 \text{ g mol}^{-1} = 0.06227 \text{ g C}_6\text{H}_{12}\text{O}_6$$

$$\text{Check. } 3.5 \times 180 = 630; 630 \times 10^{-4} = 0.063$$

- 3.29 (a) 7.08×10^{20} H atoms $\times \frac{19 \text{ C atoms}}{28 \text{ H atoms}} = 4.80 \times 10^{20}$ C atoms

- (b) 7.08×10^{20} H atoms $\times \frac{1 \text{ C}_{19}\text{H}_{28}\text{O}_2 \text{ molecule}}{28 \text{ H atoms}} = 2.529 \times 10^{19}$ molecules
 $= 2.53 \times 10^{19} \text{ C}_{19}\text{H}_{28}\text{O}_2 \text{ molecules}$

- (c) $\frac{2.529 \times 10^{19} \text{ C}_{19}\text{H}_{28}\text{O}_2 \text{ molecules}}{6.022 \times 10^{23} \text{ molecules mol}^{-1}} = 4.199 \times 10^{-5} \text{ mol}$
 $= 4.20 \times 10^{-5} \text{ mol C}_{19}\text{H}_{28}\text{O}_2$



$$\text{molar mass} = 19(12.01 \text{ g mol}^{-1}) + 28(1.008 \text{ g mol}^{-1}) + 2(16.00 \text{ g mol}^{-1}) = 288.41 \text{ g mol}^{-1} = 288.4 \text{ g mol}^{-1}$$

$$4.199 \times 10^{-5} \text{ mol C}_{19}\text{H}_{28}\text{O}_2 \times 288.4 \text{ g mol}^{-1} = 0.0121 \text{ g C}_{19}\text{H}_{28}\text{O}_2$$

- 3.30 The volume is constant throughout the problem, so we can ignore it. Use molar mass for $\text{g} \rightarrow \text{mol}$, Avogadro's number for $\text{mol} \rightarrow \text{molecules}$.

$$\frac{2.0 \times 10^{-6} \text{ g C}_2\text{H}_3\text{Cl}}{62.50 \text{ g mol}^{-1}} = 3.20 \times 10^{-8} \text{ mol}$$

$$= 3.2 \times 10^{-8} \text{ mol C}_2\text{H}_3\text{Cl} / \text{dm}^3$$

$$\frac{3.20 \times 10^{-8} \text{ mol C}_2\text{H}_3\text{Cl}}{1 \text{ dm}^3} \times 6.022 \times 10^{23} \text{ molecules mol}^{-1}$$

$$= 1.9 \times 10^{16} \text{ molecules dm}^{-3}$$

Check. $(200 \times 10^{-8}) / 60 = 2.5 \times 10^{-8}$

$$(2.5 \times 10^{-8}) \times (6 \times 10^{23}) = 15 \times 10^{15} = 1.5 \times 10^{16}$$

3.31 $\frac{25 \times 10^{-6} \text{ g C}_{21}\text{H}_{30}\text{O}_2}{314.5 \text{ g mol}^{-1}} = 7.95 \times 10^{-8} \text{ mol} = 8.0 \times 10^{-8} \text{ mol C}_{21}\text{H}_{30}\text{O}_2$

$$7.95 \times 10^{-8} \text{ mol C}_{21}\text{H}_{30}\text{O}_2 \times 6.022 \times 10^{23} \text{ molecules mol}^{-1}$$

$$= 4.8 \times 10^{16} \text{ C}_{21}\text{H}_{30}\text{O}_2 \text{ molecules}$$

Empirical Formulae

- 3.32 (a) Find the **simplest ratio of moles** by dividing by the smallest number of moles present.

$$0.0130 \text{ mol C} / 0.0065 \text{ mol} = 2$$

$$0.039 \text{ mol H} / 0.0065 \text{ mol} = 6$$

$$0.0065 \text{ mol O} / 0.0065 \text{ mol} = 1$$

The empirical formula is $\text{C}_2\text{H}_6\text{O}$.

Check. The subscripts are simple integers.

- (b) Calculate the moles of each element present, then the simplest ratio of moles.

$$\frac{11.66 \text{ g Fe}}{55.85 \text{ g mol}^{-1}} = 0.2088 \text{ mol Fe}; \quad 0.2088 / 0.2088 = 1$$

$$\frac{5.01 \text{ g O}}{16.00 \text{ g mol}^{-1}} = 0.3131 \text{ mol O}; \quad 0.3131 / 0.2088 \approx 1.5$$

Multiplying by two, the integer ratio is 2 Fe : 3 O; the empirical formula is Fe_2O_3 .

Check. The subscripts are simple integers.

(c) Assume 100 g sample, calculate moles of each element, find the simplest ratio of moles.

$$\frac{40.0 \text{ g C}}{12.01 \text{ g mol}^{-1}} = 3.33 \text{ mol C}; \quad 3.33/3.33 = 1$$

$$\frac{6.7 \text{ g H}}{1.008 \text{ g mol}^{-1}} = 6.65 \text{ mol H}; \quad 6.65/3.33 \approx 2$$

$$\frac{53.3 \text{ g O}}{16.00 \text{ g mol}^{-1}} = 3.33 \text{ mol O}; \quad 3.33/3.33 = 1$$

The empirical formula is CH_2O .

Check. The subscripts are simple integers.

3.33 (a) Calculate the simplest ratio of moles.

$$0.104 \text{ mol K}/0.052 \text{ mol} = 2$$

$$0.052 \text{ mol C}/0.052 \text{ mol} = 1$$

$$0.156 \text{ mol O}/0.052 \text{ mol} = 3$$

The empirical formula is K_2CO_3 .

(b) Calculate moles of each element present, then the simplest ratio of moles.

$$\frac{5.28 \text{ g Sn}}{118.7 \text{ g mol}^{-1}} = 0.04448 \text{ mol Sn}; \quad 0.04448/0.04448 = 1$$

$$\frac{3.37 \text{ g F}}{19.00 \text{ g mol}^{-1}} = 0.1774 \text{ mol F}; \quad 0.1774/0.04448 \approx 4$$

The integer ratio is 1 Sn : 4 F; the empirical formula is SnF_4 .

(c) Assume 100 g sample, calculate moles of each element, find the simplest ratio of moles.

$$87.5\% \text{ N} = \frac{87.5 \text{ g N}}{14.01 \text{ g mol}^{-1}} = 6.25 \text{ mol N}; \quad 6.25/6.25 = 1$$

$$12.5\% \text{ H} = \frac{12.5 \text{ g H}}{1.008 \text{ g mol}^{-1}} = 12.4 \text{ mol H}; \quad 12.4/6.25 \approx 2$$

The empirical formula is NH_2 .

3.34 The procedure in all these cases is to assume 100 g of sample, calculate the number of moles of each element present in that 100 g, then obtain the ratio of moles as smallest whole numbers.

(a) $\frac{10.4 \text{ g C}}{12.01 \text{ g mol}^{-1}} = 0.866 \text{ mol C}; \quad 0.866/0.866 = 1$

$$\frac{27.8 \text{ g S}}{32.07 \text{ g mol}^{-1}} = 0.867 \text{ mol S}; \quad 0.867/0.866 \approx 1$$

$$\frac{61.8 \text{ g Cl}}{35.45 \text{ g mol}^{-1}} = 1.74 \text{ mol Cl}; \quad 1.74/0.866 \approx 2$$

The empirical formula is CSCl_2 .

$$(b) \quad \frac{21.7 \text{ g C}}{12.01 \text{ g mol}^{-1}} = 1.81 \text{ mol C}; \quad 1.81 / 0.600 \approx 3$$

$$\frac{9.6 \text{ g O}}{16.00 \text{ g mol}^{-1}} = 0.600 \text{ mol O}; \quad 0.600 / 0.600 = 1$$

$$\frac{68.7 \text{ g F}}{19.00 \text{ g mol}^{-1}} = 3.62 \text{ mol F}; \quad 3.62 / 0.600 \approx 6$$

The empirical formula is C_3OF_6 .

$$(c) \quad \frac{21.7 \text{ g C}}{12.01 \text{ g mol}^{-1}} = 1.81 \text{ mol C}; \quad 1.81 / 0.600 \approx 3$$

$$\frac{9.6 \text{ g O}}{16.00 \text{ g mol}^{-1}} = 0.600 \text{ mol O}; \quad 0.600 / 0.600 = 1$$

$$\frac{68.7 \text{ g F}}{19.00 \text{ g mol}^{-1}} = 3.62 \text{ mol F}; \quad 3.62 / 0.600 \approx 6$$

The empirical formula is Na_3AlF_6 .

3.35 See SOLUTION 3.34 for stepwise problem-solving approach.

$$(a) \quad \frac{55.3 \text{ g K}}{39.10 \text{ g mol}^{-1}} = 1.414 \text{ mol K}; \quad 1.414 / 0.4714 \approx 3$$

$$\frac{14.6 \text{ g P}}{30.97 \text{ g mol}^{-1}} = 0.4714 \text{ mol P}; \quad 0.4714 / 0.4714 = 1$$

$$\frac{30.1 \text{ g O}}{16.00 \text{ g mol}^{-1}} = 1.881 \text{ mol O}; \quad 1.881 / 0.4714 \approx 4$$

The empirical formula is K_3PO_4 .

$$(b) \quad \frac{24.5 \text{ g Na}}{22.99 \text{ g mol}^{-1}} = 1.066 \text{ mol Na}; \quad 1.066 / 0.5304 \approx 2$$

$$\frac{14.9 \text{ g Si}}{28.09 \text{ g mol}^{-1}} = 0.5304 \text{ mol Si}; \quad 0.5304 / 0.5304 = 1$$

$$\frac{60.6 \text{ g F}}{19.00 \text{ g mol}^{-1}} = 3.189 \text{ mol F}; \quad 3.189 / 0.5304 \approx 6$$

The empirical formula is Na_2SiF_6 .

$$(c) \quad \frac{62.1 \text{ g C}}{12.01 \text{ g mol}^{-1}} = 5.17 \text{ mol C}; \quad 5.17 / 0.864 \approx 6$$

$$\frac{5.21 \text{ g H}}{1.008 \text{ g mol}^{-1}} = 5.17 \text{ mol O}; \quad 5.17 / 0.864 \approx 6$$

$$\frac{12.1 \text{ g N}}{14.01 \text{ g mol}^{-1}} = 0.864 \text{ mol N}; \quad 0.864 / 0.864 \approx 1$$

$$\frac{20.7 \text{ g O}}{16.00 \text{ g mol}^{-1}} = 1.29 \text{ mol O}; \quad 1.29 / 0.864 \approx 1.5$$

Multiplying by two, the empirical formula is $\text{C}_{12}\text{H}_{12}\text{N}_2\text{O}_3$.

- 3.36 Calculate the empirical formula mass (FM); divide FM by molar mass (MM) to calculate the integer that relates the empirical and molecular formulae. Check if FM/MM is an integer, the result is reasonable.

$$(a) \quad \text{FM CH}_2 = 12 \text{ g mol}^{-1} + 2(1 \text{ g mol}^{-1}) = 14 \text{ g mol}^{-1} \quad \frac{\text{MM}}{\text{FM}} = \frac{84 \text{ g mol}^{-1}}{14 \text{ g mol}^{-1}} = 6$$

The subscripts in the empirical formula are multiplied by 6. The molecular formula is C_6H_{12} .

$$(b) \quad \text{FM NH}_2\text{Cl} = 14.01 \text{ g mol}^{-1} + 2(1.008 \text{ g mol}^{-1}) + 35.45 \text{ g mol}^{-1} = 51.48 \text{ g mol}^{-1}$$

$$\frac{\text{MM}}{\text{FM}} = \frac{51.5 \text{ g mol}^{-1}}{51.5 \text{ g mol}^{-1}} = 1$$

The empirical and molecular formulae are NH_2Cl .

3.37 (a) HCO_2 FM $12.01 + 1.008 + 2(16.00) = 45.0$ $\frac{\text{MM}}{\text{FM}} = \frac{90.0}{45.0} = 2$

The molecular formula is $\text{H}_2\text{C}_2\text{O}_4$.

$$(b) \quad \text{C}_2\text{H}_4\text{O FM} = 2(12) + 4(1) + 16 = 44. \quad \frac{\text{MM}}{\text{FM}} = \frac{88}{44} = 2$$

The molecular formula is $\text{C}_4\text{H}_8\text{O}_2$.

- 3.38 (a) Upon combustion, all $\text{C} \rightarrow \text{CO}_2$, all $\text{H} \rightarrow \text{H}_2\text{O}$.

$$\text{mg CO}_2 \rightarrow \text{g CO}_2 \rightarrow \text{mol C}; \text{mg H}_2\text{O} \rightarrow \text{g H}_2\text{O} \rightarrow \text{mol H}$$

Find simplest ratio of moles and empirical formula.

$$\frac{5.86 \times 10^{-3} \text{ g CO}_2}{44.01 \text{ g mol}^{-1}} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 1.33 \times 10^{-4} \text{ mol C}$$

$$\frac{1.37 \times 10^{-3} \text{ g H}_2\text{O}}{18.02 \text{ g mol}^{-1}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 1.52 \times 10^{-4} \text{ mol H}$$

Dividing both values by 1.33×10^{-4} gives C:H of 1:1.14. This is not 'close enough' to be considered 1:1. No obvious multipliers (2, 3, 4) produce an integer ratio. Testing other multipliers (trial and error!), the correct factor seems to be 7. The empirical formula is C_7H_8 .

Check. See discussion of C:H ratio above.

- (b) Calculate mol C and mol H in the sample.

$$\frac{0.2829 \text{ g CO}_2}{44.01 \text{ g mol}^{-1}} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 0.0064281 \text{ mol C} = 0.006428 \text{ mol C}$$

$$\frac{0.1159 \text{ g H}_2\text{O}}{18.02 \text{ g mol}^{-1}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 0.012863 \text{ mol H} = 0.01286 \text{ mol H}$$

Calculate g C, g H and get g O by subtraction.

$$0.0064281 \text{ mol C} \times 12.01 \text{ g C mol}^{-1} = 0.07720 \text{ g C}$$

$$0.012863 \text{ mol H} \times 1.008 \text{ g H mol}^{-1} = 0.01297 \text{ g H}$$

$$\text{mass O} = 0.1005 \text{ g sample} - (0.07720 \text{ g C} + 0.01297 \text{ g H}) = 0.01033 \text{ g O}$$

Calculate mol O and find integer ratio of mol C: mol H: mol O.

$$\frac{0.01033 \text{ g O}}{16.00 \text{ g mol}^{-1}} = 6.456 \times 10^{-4} \text{ mol O}$$

Divide moles by 6.456×10^{-4} .

$$\text{C: } \frac{0.006428}{6.456 \times 10^{-4}} \approx 10; \quad \text{H: } \frac{0.01286}{6.456 \times 10^{-4}} \approx 20; \quad \text{O: } \frac{6.456 \times 10^{-4}}{6.456 \times 10^{-4}} = 1$$

The empirical formula is $\text{C}_{10}\text{H}_{20}\text{O}$.

$$\text{FM} = 10(12 \text{ g mol}^{-1}) + 20(1 \text{ g mol}^{-1}) + 16 \text{ g mol}^{-1} = 156 \text{ g mol}^{-1}$$

$$\frac{\text{MM}}{\text{FM}} = \frac{156 \text{ g mol}^{-1}}{156 \text{ g mol}^{-1}} = 1$$

The molecular formula is the same as the empirical formula, $\text{C}_{10}\text{H}_{20}\text{O}$.

Check. The mass of O was not negative or greater than the sample mass; empirical and molecular formulae are reasonable.

- 3.39 (a) Calculate mol C and mol H, then g C and g H; get g O by subtraction.

$$\frac{6.32 \times 10^{-3} \text{ g CO}_2}{44.01 \text{ g mol}^{-1}} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 1.436 \times 10^{-4} \text{ mol C} = 1.44 \times 10^{-4} \text{ mol C}$$

$$\frac{2.58 \times 10^{-3} \text{ g H}_2\text{O}}{18.02 \text{ g mol}^{-1}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 2.863 \times 10^{-4} \text{ mol H} = 2.86 \times 10^{-4} \text{ mol H}$$

$$1.436 \times 10^{-4} \text{ mol C} \times 12.01 \text{ g mol}^{-1} = 1.725 \times 10^{-3} \text{ g C} = 1.73 \text{ mg C}$$

$$2.863 \times 10^{-4} \text{ mol H} \times 1.008 \text{ g mol}^{-1} = 2.886 \times 10^{-4} \text{ g H} = 0.289 \text{ mg H}$$

$$\begin{aligned} \text{mass of O} &= 2.78 \text{ mg sample} - (1.725 \text{ mg C} + 0.289 \text{ mg H}) \\ &= 0.77 \text{ mg O} \end{aligned}$$

$$\frac{0.77 \times 10^{-3} \text{ g O}}{16.00 \text{ g mol}^{-1}} = 4.81 \times 10^{-5} \text{ mol O}$$

Divide moles by 4.81×10^{-5} mol.

$$\text{C: } \frac{1.44 \times 10^{-4}}{4.81 \times 10^{-5}} \approx 3; \quad \text{H: } \frac{2.86 \times 10^{-4}}{4.81 \times 10^{-5}} \approx 6; \quad \text{O: } \frac{4.81 \times 10^{-5}}{4.81 \times 10^{-5}} = 1$$

The empirical formula is $\text{C}_3\text{H}_6\text{O}$.

- (b) Calculate mol C and mol H, then g C and g H. In this case, get N by subtraction.

$$\frac{14.242 \times 10^{-3} \text{ g CO}_2}{44.01 \text{ g mol}^{-1}} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 3.2361 \times 10^{-4} \text{ mol C}$$

$$\frac{4.083 \times 10^{-3} \text{ g H}_2\text{O}}{18.02 \text{ g mol}^{-1}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 4.5136 \times 10^{-4} \text{ mol H} = 4.532 \times 10^{-4} \text{ mol H}$$

$$3.2361 \times 10^{-4} \text{ mol C} \times 12.01 \text{ g mol}^{-1} = 3.8866 \times 10^{-3} \text{ g C} = 3.8866 \text{ mg C}$$

$$4.532 \times 10^{-4} \text{ mol H} \times 1.008 \text{ g mol}^{-1} = 0.45683 \times 10^{-3} \text{ g H} = 0.4568 \text{ mg H}$$

$$\begin{aligned} \text{mass of N} &= 5.250 \text{ mg sample} - (3.8866 \text{ mg C} + 0.4568 \text{ mg H}) \\ &= 0.9066 \text{ mg N} = 0.907 \text{ mg N} \end{aligned}$$

$$\frac{0.9066 \times 10^{-3} \text{ g N}}{14.01 \text{ g mol}^{-1}} = 6.47 \times 10^{-5} \text{ mol N}$$

Divide moles by $6.47 \times 10^{-5} \text{ mol}$.

$$\text{C: } \frac{3.24 \times 10^{-4}}{6.47 \times 10^{-5}} \approx 5; \quad \text{H: } \frac{4.53 \times 10^{-4}}{6.47 \times 10^{-5}} \approx 7; \quad \text{N: } \frac{6.47 \times 10^{-5}}{6.47 \times 10^{-5}} = 1$$

The empirical formula is $\text{C}_5\text{H}_7\text{N}$, FM = 81 g mol^{-1} . A molar mass of $160 \pm 5 \text{ g mol}^{-1}$ indicates a factor of 2 and a molecular formula of $\text{C}_{10}\text{H}_{14}\text{N}_2$.

- 3.40 The reaction involved is $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}(\text{s}) \rightarrow \text{Na}_2\text{CO}_3(\text{s}) + x\text{H}_2\text{O}(\text{g})$.

Calculate the mass of H_2O lost and then the mole ratio of Na_2CO_3 and H_2O .

$$\text{g H}_2\text{O lost} = 2.558 \text{ g sample} - 0.948 \text{ g Na}_2\text{CO}_3 = 1.610 \text{ g}$$

$$\frac{0.948 \text{ g Na}_2\text{CO}_3}{106.0 \text{ g mol}^{-1}} = 0.00894 \text{ mol Na}_2\text{CO}_3$$

$$\frac{1.610 \text{ g H}_2\text{O}}{18.02 \text{ g mol}^{-1}} = 0.08935 \text{ mol H}_2\text{O}$$

The formula is $\text{Na}_2\text{CO}_3 \cdot \underline{10}\text{H}_2\text{O}$.

Check. x is an integer.

- 3.41 The reaction involved is $\text{MgSO}_4 \cdot x\text{H}_2\text{O}(\text{s}) \rightarrow \text{MgSO}_4(\text{s}) + x\text{H}_2\text{O}(\text{g})$. First, calculate the number of moles of product MgSO_4 ; this is the same as the number of moles of starting hydrate.

$$\frac{2.472 \text{ g MgSO}_4}{120.4 \text{ g mol}^{-1}} \times \frac{1 \text{ mol MgSO}_4 \cdot x\text{H}_2\text{O}}{1 \text{ mol MgSO}_4} = 0.02053 \text{ mol MgSO}_4 \cdot x\text{H}_2\text{O}$$

$$\text{Thus, } \frac{5.061 \text{ g MgSO}_4 \cdot x\text{H}_2\text{O}}{0.02053 \text{ mol}} = 246.5 \text{ g mol}^{-1} = \text{FM of MgSO}_4 \cdot x\text{H}_2\text{O}$$

$$\text{FM of MgSO}_4 \cdot x\text{H}_2\text{O} = \text{FM of MgSO}_4 + x(\text{FM of H}_2\text{O})$$

$$246.5 \text{ g mol}^{-1} = 120.4 \text{ g mol}^{-1} + x(18.02 \text{ g mol}^{-1})$$

$$x = 6.998$$

The hydrate formula is $\text{MgSO}_4 \cdot \underline{7}\text{H}_2\text{O}$.

Alternatively, we could calculate the number of moles of water represented by mass loss: $(5.061 \text{ g} - 2.472 \text{ g}) = 2.589 \text{ g H}_2\text{O}$ lost.

$$\frac{2.589 \text{ g H}_2\text{O}}{18.02 \text{ g mol}^{-1} \text{ H}_2\text{O}} = 0.1437 \text{ mol H}_2\text{O}$$

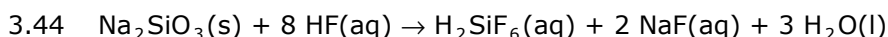
$$\frac{\text{mol H}_2\text{O}}{\text{mol MgSO}_4} = \frac{0.1437}{0.02053} = 7.000$$

Again the correct formula is $\text{MgSO}_4 \cdot \underline{7}\text{H}_2\text{O}$.

Calculations Based on Chemical Equations

3.42 The mole ratios, implicit in the coefficients of a balanced chemical equation, express the fundamental relationship between amounts of reactants and products. If the equation is not balanced, the mole ratios will be incorrect and lead to erroneous calculated amounts of products.

3.43 The **integer coefficients** immediately preceding each molecular formula in a chemical equation give information about relative numbers of moles of reactants and products involved in a reaction.



(a) Use the mole ratio 8 HF : 1 Na_2SiO_3 from the balanced equation to relate moles of the two reactants.

$$0.300 \text{ mol Na}_2\text{SiO}_3 \times \frac{8 \text{ mol HF}}{1 \text{ mol Na}_2\text{SiO}_3} = 2.40 \text{ mol HF}$$

Check. Mol HF should be greater than mol Na_2SiO_3 .

(b) Use the mole ratio 2 NaF : 8 HF to change mol HF to mol NaF, then molar mass to get g NaF.

$$0.500 \text{ mol HF} \times \frac{2 \text{ mol NaF}}{8 \text{ mol HF}} \times 41.99 \text{ g mol}^{-1} = 5.25 \text{ g NaF}$$

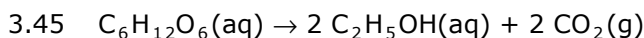
Check. $(0.5 \text{ mol}/4) = 0.125 \text{ mol}$; $0.13 \text{ mol} \times 42 \text{ g mol}^{-1} > 4 \text{ g NaF}$

(c) $\text{g HF} \rightarrow \text{mol HF} \left(\frac{\text{mol}}{\text{ratio}} \right) \rightarrow \text{mol Na}_2\text{SiO}_3 \rightarrow \text{g Na}_2\text{SiO}_3$

The mole ratio is at the heart of every stoichiometry problem. Molar mass is used to change to and from grams.

$$\frac{0.800 \text{ g HF}}{20.01 \text{ g mol}^{-1}} \times \frac{1 \text{ mol Na}_2\text{SiO}_3}{8 \text{ mol HF}} \times 122.1 \text{ g mol}^{-1} = 0.610 \text{ g Na}_2\text{SiO}_3$$

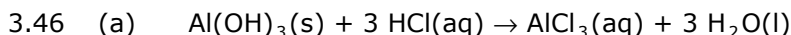
Check. $0.8 \text{ g} (120/160) < 0.75 \text{ g}$



(a) $0.400 \text{ mol C}_6\text{H}_{12}\text{O}_6 \times \frac{2 \text{ mol CO}_2}{1 \text{ mol C}_6\text{H}_{12}\text{O}_6} = 0.800 \text{ mol CO}_2$

(b) $\frac{7.50 \text{ g C}_2\text{H}_5\text{OH}}{46.07 \text{ g mol}^{-1}} \times \frac{1 \text{ mol C}_6\text{H}_{12}\text{O}_6}{2 \text{ mol C}_2\text{H}_5\text{OH}} \times 180.2 \text{ g mol}^{-1} = 14.7 \text{ g C}_6\text{H}_{12}\text{O}_6$

(c) $\frac{7.50 \text{ g C}_2\text{H}_5\text{OH}}{46.07 \text{ g mol}^{-1}} \times \frac{2 \text{ mol CO}_2}{2 \text{ mol C}_2\text{H}_5\text{OH}} \times 44.01 \text{ g mol}^{-1} = 7.16 \text{ g CO}_2$



(b) Follow the logic in SAMPLE EXERCISE 3.16. Calculate mol $\text{Al}(\text{OH})_3$ in 0.500 g $\text{Al}(\text{OH})_3$ separately, since it will be used several times.

$$\frac{0.500 \text{ g Al(OH)}_3}{78.00 \text{ g mol}^{-1}} = 6.410 \times 10^{-3} \text{ mol Al(OH)}_3 = 6.41 \times 10^{-3} \text{ mol Al(OH)}_3$$

$$6.410 \times 10^{-3} \text{ mol Al(OH)}_3 \times \frac{3 \text{ mol HCl}}{1 \text{ mol Al(OH)}_3} \times 36.46 \text{ g mol}^{-1} = 0.7012 \text{ g HCl}$$

$$= 0.701 \text{ g HCl}$$

$$(c) \quad 6.410 \times 10^{-3} \text{ mol Al(OH)}_3 \times \frac{1 \text{ mol AlCl}_3}{1 \text{ mol Al(OH)}_3} \times 133.34 \text{ g mol}^{-1} = 0.8547 \text{ g AlCl}_3$$

$$= 0.855 \text{ g AlCl}_3$$

$$6.410 \times 10^{-3} \text{ mol Al(OH)}_3 \times \frac{3 \text{ mol H}_2\text{O}}{1 \text{ mol Al(OH)}_3} \times 18.02 \text{ g mol}^{-1} = 0.3465 \text{ g H}_2\text{O}$$

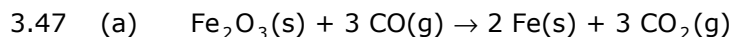
$$= 0.347 \text{ g H}_2\text{O}$$

(d) Conservation of mass: mass of products = mass of reactants

reactants: $\text{Al(OH)}_3 + \text{HCl}$, $0.500 \text{ g} + 0.701 \text{ g} = 1.201 \text{ g}$

products: $\text{AlCl}_3 + \text{H}_2\text{O}$, $0.855 \text{ g} + 0.347 \text{ g} = 1.202 \text{ g}$

The 0.001 g difference is due to rounding ($0.8547 + 0.3465 = 1.2012$). This is an excellent *check* of results.



$$(b) \quad \frac{(0.150 \times 1 \times 1000) \text{ g Fe}_2\text{O}_3}{159.688 \text{ g mol}^{-1}} = 0.9393 \text{ mol Fe}_2\text{O}_3$$

$$0.939 \text{ mol Fe}_2\text{O}_3$$

$$0.9393 \text{ mol Fe}_2\text{O}_3 \times \frac{3 \text{ mol CO}}{1 \text{ mol Fe}_2\text{O}_3} \times 28.01 \text{ g mol}^{-1}$$

$$= 78.929 \text{ g CO} = 78.9 \text{ g CO}$$

$$(c) \quad 0.9393 \text{ mol Fe}_2\text{O}_3 \times \frac{2 \text{ mol Fe}}{1 \text{ mol Fe}_2\text{O}_3} \times 55.845 \text{ g mol}^{-1} = 104.914 \text{ g Fe}$$

$$= 105 \text{ g Fe}$$

$$0.9393 \text{ mol Fe}_2\text{O}_3 \times \frac{3 \text{ mol CO}_2}{1 \text{ mol Fe}_2\text{O}_3} \times 44.01 \text{ g mol}^{-1} = 124.015 \text{ g CO}_2$$

$$= 124 \text{ g CO}_2$$

(d) reactants: $150 \text{ g Fe}_2\text{O}_3 + 78.9 \text{ g CO} = 228.9 \text{ g} = 229 \text{ g}$

products: $104.9 \text{ g Fe} + 124.0 \text{ g CO}_2 = 228.9 \text{ g} = 229 \text{ g}$

Mass is conserved.

3.48 (a) Use mole ratio from balanced equation.

$$1.50 \text{ mol NaN}_3 \times \frac{3 \text{ mol N}_2}{2 \text{ mol NaN}_3} = 2.25 \text{ mol N}_2$$

Check. The resulting mol N_2 should be greater than mol NaN_3 (the N_2 : NaN_3 ratio is > 1); and it is.

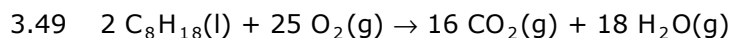
(b) Given: g N₂. Find: g NaN₃.

Use molar masses to get from and to grams, mol ratio to relate moles of the two substances.

$$\frac{10.0 \text{ g N}_2}{28.01 \text{ g mol}^{-1}} \times \frac{2 \text{ mol NaN}_3}{3 \text{ mol N}_2} \times 65.01 \text{ g mol}^{-1} = 15.5 \text{ g NaN}_3$$

Check. Mass relations are less intuitive than mole relations. Estimating the ratio of molar masses is sometimes useful. In this case 65 g NaN₃/28 g N₂ ≈ 2.25. Then, (10 g × 2/3 × 2.25) ≈ 14 g NaN₃. The calculated result looks reasonable.

$$\begin{aligned} \text{(c)} \quad \text{mol N}_2 &= \frac{(1.25 \text{ g dm}^{-3} \times 280 \text{ dm}^3)}{28.01 \text{ g mol}^{-1}} = 12.5 \text{ mol} \\ \text{mol NaN}_3 &= (12.5 \times 2/3) \text{ mol} = 8.33 \text{ mol} \\ \text{g NaN}_3 &= 8.33 \text{ mol} \times 65.01 \text{ g mol}^{-1} = 542 \text{ g} \end{aligned}$$



$$\text{(a)} \quad 1.25 \text{ mol C}_8\text{H}_{18} \times \frac{25 \text{ mol O}_2}{2 \text{ mol C}_8\text{H}_{18}} = 15.625 \text{ mol O}_2 = 15.6 \text{ mol O}_2$$

$$\text{(b)} \quad \frac{10.0 \text{ g C}_8\text{H}_{18}}{114.2 \text{ g mol}^{-1}} \times \frac{25 \text{ mol O}_2}{2 \text{ mol C}_8\text{H}_{18}} \times 32.00 \text{ g mol}^{-1} = 35.0 \text{ g O}_2$$

$$\begin{aligned} \text{(c)} \quad \text{mol C}_8\text{H}_{18} &= \frac{(4.54 \times 1 \times 1000) \text{ cm}^3 \times 0.692 \text{ g cm}^{-3}}{114.2 \text{ g mol}^{-1}} = 27.5 \text{ mol} \\ \text{g O}_2 &= (27.5 \times 25/2) \text{ mol} \times 32.00 \text{ g mol}^{-1} = 1.10 \times 10^4 \text{ g} \end{aligned}$$

Limiting Reactants; Theoretical Yields

3.50 (a) The *limiting reactant* determines the maximum number of moles of product resulting from a chemical reaction; any other reactant is an *excess reactant*.

(b) The limiting reactant regulates the amount of products, because it is completely used up during the reaction; no more product can be made when one of the reactants is unavailable.

3.51 (a) *Theoretical yield* is the maximum amount of product possible, as predicted by stoichiometry, assuming that the limiting reactant is converted entirely to product.

Actual yield is the amount of product actually obtained, less than or equal to the theoretical yield. *Percent yield* is the ratio of (actual yield to theoretical yield) × 100.

(b) No reaction is perfect. Not all reactant molecules come together effectively to form products; alternative reaction pathways may produce secondary products and reduce the amount of desired product actually

obtained, or it might not be possible to completely isolate the desired product from the reaction mixture. All these factors reduce the actual yield of a reaction.

- 3.52 Amounts of more than one reactant are given, so we must determine which reactant regulates (limits) the product. Then apply the appropriate mole ratio from the balanced equation.

The mole ratio is 2 NaOH:1 CO₂, so 1.00 mol CO₂ requires 2.00 mol NaOH for complete reaction. Less than 2.00 mol NaOH are present, so NaOH is the limiting reactant.

$$1.85 \text{ mol NaOH} \times \frac{1 \text{ mol Na}_2\text{CO}_3}{2 \text{ mol NaOH}} = 0.925 \text{ mol Na}_2\text{CO}_3 \text{ can be produced}$$

The Na₂CO₃:CO₂ ratio is 1:1, so 0.925 mol Na₂CO₃ produced requires 0.925 mol CO₂ consumed. (Alternately, 1.85 mol NaOH \times 1 mol CO₂/2 mol NaOH = 0.925 mol CO₂ reacted).

$$1.00 \text{ mol CO}_2 \text{ initial} - 0.925 \text{ mol CO}_2 \text{ reacted} = 0.075 \text{ mol CO}_2 \text{ remain}$$

<i>Check.</i>	2 NaOH(s)	+	CO ₂ (g)	\rightarrow	Na ₂ CO ₃ (s)	+	H ₂ O(l)
initial	1.85 mol		1.00 mol		0 mol		
change (reaction)	-1.85 mol		-0.925 mol		+0.925 mol		
final	0 mol		0.075 mol		0.925 mol		

Note that the 'change' line (but not necessarily the 'final' line) is proportional to the mole ratios from the balanced equation.

3.53 $0.500 \text{ mol Al(OH)}_3 \times \frac{3 \text{ mol H}_2\text{SO}_4}{2 \text{ mol Al(OH)}_3}$
 $= 0.750 \text{ mol H}_2\text{SO}_4$ needed for complete reaction

Only 0.500 mol H₂SO₄ available, so H₂SO₄ limits.

$$0.500 \text{ mol H}_2\text{SO}_4 \times \frac{1 \text{ mol Al}_2(\text{SO}_4)_3}{3 \text{ mol H}_2\text{SO}_4} = 0.1667 \text{ mol Al}_2(\text{SO}_4)_3$$

0.167 mol Al₂(SO₄)₃ can form

$$0.500 \text{ mol H}_2\text{SO}_4 \times \frac{2 \text{ mol Al(OH)}_3}{3 \text{ mol H}_2\text{SO}_4} = 0.3333 \text{ mol Al(OH)}_3$$

$$= 0.333 \text{ mol Al(OH)}_3 \text{ react}$$

$$0.500 \text{ mol Al(OH)}_3 \text{ initial} - 0.333 \text{ mol react} = 0.167 \text{ mol Al(OH)}_3 \text{ remain}$$

- 3.54 $3 \text{ NaHCO}_3(\text{aq}) + \text{H}_3\text{C}_6\text{H}_5\text{O}_7(\text{aq}) \rightarrow 3 \text{ CO}_2(\text{g}) + 3 \text{ H}_2\text{O}(\text{l}) + \text{Na}_3\text{C}_6\text{H}_5\text{O}_7(\text{aq})$

(a) Follow the approach in SAMPLE EXERCISE 3.18. In the working citric acid is abbreviated as H₃Cit.

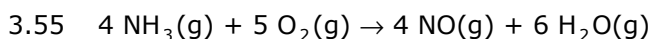
$$\frac{1.00 \text{ g NaHCO}_3}{84.01 \text{ g mol}^{-1}} = 1.190 \times 10^{-2} \text{ mol NaHCO}_3 = 1.19 \times 10^{-2} \text{ mol NaHCO}_3$$

$$\frac{1.00 \text{ g H}_3\text{C}_6\text{H}_5\text{O}_7}{192.1 \text{ g mol}^{-1}} = 5.206 \times 10^{-3} \text{ mol H}_3\text{Cit} = 5.21 \times 10^{-3} \text{ mol H}_3\text{Cit}$$

But NaHCO_3 and H_3Cit react in a 3:1 ratio, so $5.21 \times 10^{-3} \text{ mol H}_3\text{Cit}$ require $3(5.21 \times 10^{-3} \text{ mol}) = 1.56 \times 10^{-2} \text{ mol NaHCO}_3$. We have only $1.19 \times 10^{-2} \text{ mol NaHCO}_3$, so NaHCO_3 is the limiting reactant.

$$\begin{aligned} \text{(b)} \quad & 1.190 \times 10^{-2} \text{ mol NaHCO}_3 \times \frac{3 \text{ mol CO}_2}{3 \text{ mol NaHCO}_3} \times 44.01 \text{ g mol}^{-1} \\ & = 0.524 \text{ g CO}_2 \end{aligned}$$

$$\begin{aligned} \text{(c)} \quad & 1.190 \times 10^{-2} \text{ mol NaHCO}_3 \times \frac{1 \text{ mol H}_3\text{Cit}}{3 \text{ mol NaHCO}_3} = 3.968 \times 10^{-3} \text{ mol H}_3\text{Cit} \\ & = 3.97 \times 10^{-3} \text{ mol H}_3\text{Cit react} \\ & 5.206 \times 10^{-3} \text{ mol H}_3\text{Cit initial} - 3.968 \times 10^{-3} \text{ mol H}_3\text{Cit react} \\ & = 1.238 \times 10^{-3} \text{ mol H}_3\text{Cit} = 1.24 \times 10^{-3} \text{ mol H}_3\text{Cit remain} \\ & 1.238 \times 10^{-3} \text{ mol H}_3\text{Cit} \times 192.1 \text{ g mol}^{-1} = 0.238 \text{ g H}_3\text{Cit remain} \end{aligned}$$



(a) Follow the approach in SAMPLE EXERCISE 3.18.

$$1.50 \text{ g NH}_3 \times \frac{1 \text{ mol NH}_3}{17.03 \text{ g NH}_3} = 0.08808 \text{ mol NH}_3 = 0.0881 \text{ mol NH}_3$$

$$2.75 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} = 0.08594 \text{ mol O}_2 = 0.0859 \text{ mol O}_2$$

$$0.08594 \text{ mol O}_2 \times \frac{4 \text{ mol NH}_3}{5 \text{ mol O}_2} = 0.06875 \text{ mol NH}_3 = 0.0688 \text{ mol NH}_3 \text{ required}$$

More than 0.0688 mol NH_3 is available, so O_2 is the limiting reactant.

$$\begin{aligned} \text{(b)} \quad & 0.08594 \text{ mol O}_2 \times \frac{4 \text{ mol NO}}{5 \text{ mol O}_2} \times 30.01 \text{ g mol}^{-1} = 2.063 \text{ g NO} \\ & = 2.06 \text{ g NO produced} \end{aligned}$$

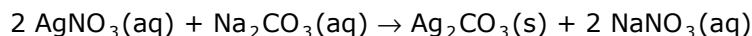
$$\begin{aligned} & 0.08594 \text{ mol O}_2 \times \frac{6 \text{ mol H}_2\text{O}}{5 \text{ mol O}_2} \times 18.02 \text{ g mol}^{-1} = 1.8583 \text{ g H}_2\text{O} \\ & = 1.86 \text{ g H}_2\text{O} \end{aligned}$$

$$\begin{aligned} \text{(c)} \quad & 0.08808 \text{ mol NH}_3 - 0.06875 \text{ mol NH}_3 \text{ reacted} = 0.01933 \text{ mol NH}_3 \\ & = 0.0193 \text{ mol NH}_3 \text{ remain} \\ & 0.01933 \text{ mol NH}_3 \times 17.03 \text{ g mol}^{-1} = 0.32919 \text{ g NH}_3 \\ & = 0.329 \text{ g NH}_3 \text{ remain} \end{aligned}$$

$$\begin{aligned} \text{(d)} \quad & \text{mass products} = 2.06 \text{ g NO} + 1.86 \text{ g H}_2\text{O} + 0.329 \text{ g NH}_3 \text{ remaining} \\ & = 4.25 \text{ g products} \\ & \text{mass reactants} = 1.50 \text{ g NH}_3 + 2.75 \text{ g O}_2 = 4.25 \text{ g reactants} \end{aligned}$$

(For comparison purposes, the mass of excess reactant can be either added to the products, as above, or subtracted from reactants.)

- 3.56 Write balanced equation; determine limiting reactant; calculate amounts of excess reactant remaining and products, based on limiting reactant.



$$\frac{3.50 \text{ g Na}_2\text{CO}_3}{106.0 \text{ g mol}^{-1}} = 0.03302 \text{ mol Na}_2\text{CO}_3 = 0.0330 \text{ mol Na}_2\text{CO}_3$$

$$\frac{5.00 \text{ g AgNO}_3}{169.9 \text{ g mol}^{-1}} = 0.02943 \text{ mol AgNO}_3 = 0.0294 \text{ mol AgNO}_3$$

$$0.02943 \text{ mol AgNO}_3 \times \frac{1 \text{ mol Na}_2\text{CO}_3}{2 \text{ mol AgNO}_3} = 0.01471 \text{ mol Na}_2\text{CO}_3$$

$$= 0.0147 \text{ mol Na}_2\text{CO}_3 \text{ required}$$

AgNO_3 is the limiting reactant and Na_2CO_3 is present in excess.

	$2 \text{AgNO}_3(\text{aq})$	$+ \text{Na}_2\text{CO}_3(\text{aq})$	$\rightarrow \text{Ag}_2\text{CO}_3(\text{s})$	$+ 2 \text{NaNO}_3(\text{aq})$
initial	0.0294 mol	0.0330 mol	0 mol	0 mol
reaction	-0.0294 mol	-0.0147 mol	+0.0147 mol	+0.0294 mol
final	0 mol	0.0183 mol	0.0147 mol	0.0294 mol

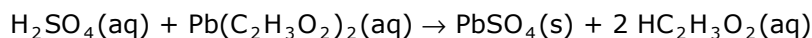
$$0.01830 \text{ mol Na}_2\text{CO}_3 \times 106.0 \text{ g mol}^{-1} = 1.940 \text{ g Na}_2\text{CO}_3 = 1.94 \text{ g Na}_2\text{CO}_3$$

$$0.01471 \text{ mol Ag}_2\text{CO}_3 \times 275.8 \text{ g mol}^{-1} = 4.057 \text{ g Ag}_2\text{CO}_3 = 4.06 \text{ g Ag}_2\text{CO}_3$$

$$0.02943 \text{ mol NaNO}_3 \times 85.00 \text{ g mol}^{-1} = 2.502 \text{ g NaNO}_3 = 2.50 \text{ g NaNO}_3$$

Check. The initial mass of reactants was 8.50 g, and the final mass of excess reactant and products is 8.50 g; mass is conserved.

- 3.57 Write balanced equation; determine limiting reactant; calculate amounts of excess reactant remaining and products, based on limiting reactant.



$$\frac{7.50 \text{ g H}_2\text{SO}_4}{98.09 \text{ g mol}^{-1}} = 0.07646 \text{ mol H}_2\text{SO}_4 = 0.0765 \text{ mol H}_2\text{SO}_4$$

$$\frac{7.50 \text{ g Pb}(\text{C}_2\text{H}_3\text{O}_2)_2}{325.3 \text{ g mol}^{-1}} = 0.023056 \text{ mol Pb}(\text{C}_2\text{H}_3\text{O}_2)_2 = 0.0231 \text{ mol Pb}(\text{C}_2\text{H}_3\text{O}_2)_2$$

1 mol H_2SO_4 :1 mol $\text{Pb}(\text{C}_2\text{H}_3\text{O}_2)_2$, so $\text{Pb}(\text{C}_2\text{H}_3\text{O}_2)_2$ is the limiting reactant.

when 0 mol $\text{Pb}(\text{C}_2\text{H}_3\text{O}_2)_2$ left, then $(0.07646 - 0.023056) \text{ mol} = 0.0534 \text{ mol H}_2\text{SO}_4$ and 0.0231 mol PbSO_4 remain

$(0.023056 \text{ mol} \times 2) = 0.0461 \text{ mol HC}_2\text{H}_3\text{O}_2$ are present after reaction

$$0.053405 \text{ mol H}_2\text{SO}_4 \times 98.09 \text{ g mol}^{-1} = 5.2385 \text{ g H}_2\text{SO}_4 = 5.24 \text{ g H}_2\text{SO}_4$$

$$0.023056 \text{ mol PbSO}_4 \times 303.3 \text{ g mol}^{-1} = 6.9928 \text{ g PbSO}_4 = 6.99 \text{ g PbSO}_4$$

$$0.046111 \text{ mol HC}_2\text{H}_3\text{O}_2 \times 60.05 \text{ g mol}^{-1} = 2.7690 \text{ g HC}_2\text{H}_3\text{O}_2$$

$$= 2.77 \text{ g HC}_2\text{H}_3\text{O}_2$$

Check. The initial mass of reactants was 15.00 g; and the final mass of excess reactant and products is 15.00 g; mass is conserved.

- 3.58 Determine the limiting reactant and the maximum amount of product it could produce. Then calculate yield in %.

$$(a) \quad \frac{30.0 \text{ g C}_6\text{H}_6}{78.11 \text{ g mol}^{-1}} = 0.3841 \text{ mol C}_6\text{H}_6 = 0.384 \text{ mol C}_6\text{H}_6$$

$$\frac{65.0 \text{ g Br}_2}{159.8 \text{ g mol}^{-1}} = 0.4068 \text{ mol Br}_2 = 0.407 \text{ mol Br}_2$$

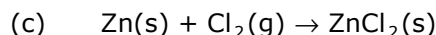
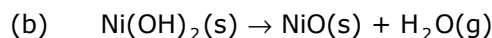
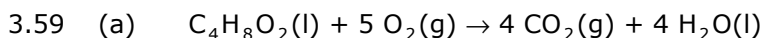
Since C_6H_6 and Br_2 react in a 1:1 mole ratio, C_6H_6 is the limiting reactant and determines the theoretical yield.

$$0.3841 \text{ mol C}_6\text{H}_6 \times \frac{1 \text{ mol C}_6\text{H}_5\text{Br}}{1 \text{ mol C}_6\text{H}_6} \times 157.0 \text{ g mol}^{-1} = 60.30 \text{ g C}_6\text{H}_5\text{Br} = 60.3 \text{ g C}_6\text{H}_5\text{Br}$$

Check. $30/78 \sim 3/8 \text{ mol C}_6\text{H}_6$. $65/160 \sim 3/8 \text{ mol Br}_2$. Since moles of the two reactants are similar, a precise calculation is needed to determine the limiting reactant. $3/8 \times 160 \approx 60 \text{ g product}$.

$$(b) \quad \% \text{ yield} = \frac{56.7 \text{ g C}_6\text{H}_5\text{Br actual}}{60.3 \text{ g C}_6\text{H}_5\text{Br theoretical}} \times 100 = 94.0\%$$

Additional Exercise



3.60 (a) $1.25 \text{ carat} \times 0.200 \text{ g carat}^{-1} \times \frac{1}{12.01 \text{ g mol}^{-1}} = 0.020816 \text{ mol C}$
 $= 0.0208 \text{ mol C}$

$$0.020816 \text{ mol C} \times 6.022 \times 10^{23} \text{ atoms mol}^{-1} = 1.25 \times 10^{22} \text{ C atoms}$$

(b) $\frac{0.500 \text{ g C}_9\text{H}_8\text{O}_4}{180.2 \text{ g mol}^{-1}} = 2.7747 \times 10^{-3} \text{ mol HC}_9\text{H}_7\text{O}_4 = 2.77 \times 10^{-3} \text{ mol HC}_9\text{H}_7\text{O}_4$

$$0.0027747 \text{ mol C}_9\text{H}_8\text{O}_4 \times 6.022 \times 10^{23} \text{ molecules mol}^{-1}$$

$$= 1.67 \times 10^{21} \text{ HC}_9\text{H}_7\text{O}_4 \text{ molecules}$$

3.61 (a) $5.342 \times 10^{-21} \text{ g molecule}^{-1} \times 6.0221 \times 10^{23} \text{ molecules mol}^{-1}$
 $= 3217 \text{ g mol}^{-1}$

(b) 1.00 g haemoglobin (hem) contains $3.40 \times 10^{-3} \text{ g Fe}$.

$$\frac{1.00 \text{ g hem}}{3.40 \times 10^{-3} \text{ g Fe}} \times 55.85 \text{ g mol}^{-1} \times \frac{4 \text{ mol Fe}}{1 \text{ mol hem}}$$

$$= 6.57 \times 10^4 \text{ g mol}^{-1} \text{ haemoglobin}$$

$$3.62 \quad (a) \quad \frac{1.000 \times 10^4 \text{ Si atoms}}{6.022 \times 10^{23} \text{ atoms mol}^{-1}} \times 28.0855 \text{ g mol}^{-1} = 4.6638 \times 10^{-19} \text{ g Si}$$

$$(b) \quad 4.6638 \times 10^{-19} \text{ g Si} \times \frac{1}{2.3 \text{ g cm}^{-3}} = 2.03 \times 10^{-19} \text{ cm}^3 = 2.0 \times 10^{-19} \text{ cm}^3$$

$$(c) \quad V = l^3; l = (V)^{1/3} = (2.03 \times 10^{-19} \text{ cm}^3)^{1/3} = 6.0 \times 10^{-7} \text{ cm} \\ (= 6.0 \text{ nm})$$

3.63 Assume 100 g, calculate mole ratios, empirical formula, then molecular formula from molar mass.

$$\frac{68.2 \text{ g C}}{12.01 \text{ g mol}^{-1}} = 5.68 \text{ mol C}; \quad 5.68 / 0.568 \approx 10$$

$$\frac{6.86 \text{ g H}}{1.008 \text{ g mol}^{-1}} = 6.81 \text{ mol H}; \quad 6.81 / 0.568 \approx 12$$

$$\frac{15.9 \text{ g N}}{14.01 \text{ g mol}^{-1}} = 1.13 \text{ mol N}; \quad 1.13 / 0.568 \approx 2$$

$$\frac{9.08 \text{ g O}}{16.00 \text{ g mol}^{-1}} = 0.568 \text{ mol O}; \quad 0.568 / 0.568 = 1$$

The empirical formula is $\text{C}_{10}\text{H}_{12}\text{N}_2\text{O}$, FM = 176 u (or g). Since the molar mass is 176, the empirical and molecular formula are the same, $\text{C}_{10}\text{H}_{12}\text{N}_2\text{O}$.

3.64 Because different sample sizes were used to analyse the different elements, calculate mass% of each element in the sample.

- i. Calculate mass% C from g CO_2 .
- ii. Calculate mass% Cl from AgCl.
- iii. Get mass% H by subtraction.
- iv. Calculate mole ratios and the empirical formula.

$$i. \quad \frac{3.52 \text{ g CO}_2}{44.01 \text{ g mol}^{-1}} \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \times 12.01 \text{ g mol}^{-1} = 0.9606 \text{ g C} = 0.961 \text{ g C} \\ \frac{0.9606 \text{ g C}}{1.50 \text{ g sample}} \times 100 = 64.04\% \text{ C} = 64.0\% \text{ C}$$

$$ii. \quad \frac{1.27 \text{ g AgCl}}{143.3 \text{ g mol}^{-1}} \times \frac{1 \text{ mol Cl}}{1 \text{ mol AgCl}} \times 35.45 \text{ g mol}^{-1} = 0.3142 \text{ g Cl} = 0.314 \text{ g Cl} \\ \frac{0.3142 \text{ g Cl}}{1.00 \text{ g sample}} \times 100 = 31.42\% \text{ Cl} = 31.4\% \text{ Cl}$$

$$iii. \quad \% \text{ H} = 100.0\% - (64.04\% \text{ C} + 31.42\% \text{ Cl}) = 4.54\% \text{ H} = 4.5\% \text{ H}$$

iv. Assume 100 g sample.

$$\frac{64.04 \text{ g C}}{12.01 \text{ g mol}^{-1}} = 5.33 \text{ mol C}; \quad 5.33 / 0.886 = 6.02$$

$$\frac{31.42 \text{ g Cl}}{35.45 \text{ g mol}^{-1}} = 0.886 \text{ mol Cl}; \quad 0.886 / 0.886 = 1.00$$

$$\frac{4.54 \text{ g H}}{1.008 \text{ g mol}^{-1}} = 4.50 \text{ mol H}; \quad 4.50 / 0.886 = 5.08$$

The empirical formula is probably C_6H_5Cl .

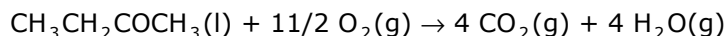
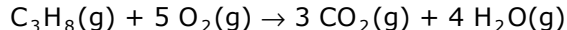
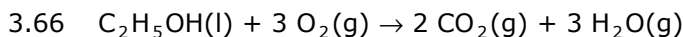
The subscript for H, 5.08, is relatively far from 5.00 as the relatively small weight of hydrogen has been determined by subtracting two larger values each with associated errors. It would be better in practice to determine H from the amount of H_2O generated in the combustion reaction. However, C_6H_5Cl makes chemical sense. More significant figures in the mass data are required for a more accurate mole ratio.

- 3.65 (a) Let AM = the atomic mass of X.

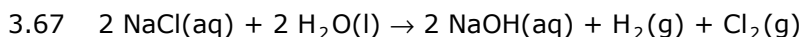
According to the chemical reaction, moles XI_3 reacted = moles XCl_3 produced

$$\begin{aligned}0.5000 \text{ g } XI_3 \times 1 \text{ mol } XI_3 / (AM + 380.71) \text{ g } XI_3 \\&= 0.2360 \text{ g } XCl_3 \times \frac{1 \text{ mol } XCl_3}{(AM + 106.36) \text{ g } XCl_3} \\0.5000 (AM + 106.36) &= 0.2360 (AM + 380.71) \\0.5000 AM + 53.180 &= 0.2360 AM + 89.848 \\0.2640 AM &= 36.67 \\AM &= 138.9 \text{ g}\end{aligned}$$

- (b) X is lanthanum, La, atomic number 57.



In a combustion reaction, all H in the fuel is transformed to H_2O in the products. The reactant with most mol H/mol fuel will produce the most H_2O . C_3H_8 and $CH_3CH_2COCH_3$ (C_4H_8O) both have 8 mol H/mol fuel, so 1.5 mol of either fuel will produce the same amount of H_2O . 1.5 mol C_2H_5OH will produce less H_2O .



Calculate mol Cl_2 and relate to mol H_2 , mol NaOH.

$$\begin{aligned}\frac{1.5 \times 10^6 \times 1 \times 10^3 \text{ g}}{70.91 \text{ g mol}^{-1}} &= 2.115 \times 10^7 \text{ mol } Cl_2 = 2.1 \times 10^7 \text{ mol } Cl_2 \\2.115 \times 10^7 \text{ mol } Cl_2 \times \frac{1 \text{ mol } H_2}{1 \text{ mol } Cl_2} \times 2.016 \text{ g mol}^{-1} &= 4.26 \times 10^7 \text{ g } H_2 = 4.3 \times 10^4 \text{ kg } H_2 \\&= 4.3 \text{ tonnes } H_2 \\2.115 \times 10^7 \text{ mol } Cl_2 \times \frac{2 \text{ mol } NaOH}{1 \text{ mol } Cl_2} \times 40.0 \text{ g mol}^{-1} &= 1.69 \times 10^9 \text{ g } NaOH \\&= 1.7 \times 10^9 \text{ g } NaOH \\1.7 \times 10^9 \text{ g } NaOH &= 1.7 \times 10^6 \text{ kg } NaOH = 1700 \text{ tonnes } NaOH\end{aligned}$$



(b) Following the approach in SAMPLE EXERCISE 3.18

$$\frac{10.0 \text{ g C}_2\text{H}_2}{26.04 \text{ g mol}^{-1}} \times \frac{5 \text{ mol O}_2}{2 \text{ mol C}_2\text{H}_2} \times 32.00 \text{ g mol}^{-1} = 30.7 \text{ g O}_2 \text{ required}$$

Only 10.0 g O₂ are available, so O₂ limits.

(c) Since O₂ limits, 0.0 g O₂ remain.

Next, calculate the g C₂H₂ consumed and the amounts of CO₂ and H₂O produced by reaction of 10.0 g O₂.

$$\frac{10.0 \text{ g O}_2}{32.00 \text{ g mol}^{-1}} \times \frac{2 \text{ mol C}_2\text{H}_2}{5 \text{ mol O}_2} \times 26.04 \text{ g mol}^{-1} = 3.26 \text{ g C}_2\text{H}_2 \text{ consumed}$$

$$10.0 \text{ g C}_2\text{H}_2 \text{ initial} - 3.26 \text{ g C}_2\text{H}_2 \text{ consumed} = 6.74 \text{ g C}_2\text{H}_2$$

$$= 6.7 \text{ g C}_2\text{H}_2 \text{ remain}$$

$$\frac{10.0 \text{ g O}_2}{32.00 \text{ g mol}^{-1}} \times \frac{4 \text{ mol CO}_2}{5 \text{ mol O}_2} \times 44.01 \text{ g mol}^{-1} = 11.0 \text{ g CO}_2 \text{ produced}$$

$$\frac{10.0 \text{ g O}_2}{32.00 \text{ g mol}^{-1}} \times \frac{2 \text{ mol H}_2\text{O}}{5 \text{ mol O}_2} \times 18.02 \text{ g mol}^{-1} = 2.25 \text{ g H}_2\text{O produced}$$

Integrative Exercise

3.69 Volume cube $\xrightarrow{\text{density}}$ mass CaCO₃ \rightarrow moles CaCO₃ \rightarrow moles O \rightarrow O atoms

$$(5.093)^3 \text{ cm}^3 \times 2.71 \text{ g cm}^{-3} \text{ CaCO}_3 \times \frac{1}{100.1 \text{ g mol}^{-1}} \times \frac{3 \text{ mol O}}{1 \text{ mol CaCO}_3} \times 6.022 \times 10^{23} \text{ O atoms mol}^{-1} \\ = 6.46 \times 10^{24} \text{ O atoms}$$

3.70 (a) Volume of Ag cube $\xrightarrow{\text{density}}$ mass of Ag \rightarrow mol Ag \rightarrow Ag atoms

$$(1.000)^3 \text{ cm}^3 \text{ Ag} \times 10.49 \text{ g cm}^{-3} \times \frac{1}{107.87 \text{ g mol}^{-1}} \times 6.022 \times 10^{23} \text{ atoms mol}^{-1} \\ = 5.8562 \times 10^{22} \text{ Ag atoms} = 5.856 \times 10^{22} \text{ Ag atoms}$$

(b) 1.000 cm³ cube volume, 74% is occupied by Ag atoms

$$0.7400 \text{ cm}^3 = \text{volume of } 5.856 \times 10^{22} \text{ Ag atoms}$$

$$\frac{0.7400 \text{ cm}^3}{5.8562 \times 10^{22} \text{ Ag atoms}} = 1.2636 \times 10^{-23} \text{ cm}^3 (\text{Ag atom})^{-1} \\ = 1.264 \times 10^{-23} \text{ cm}^3 (\text{Ag atom})^{-1}$$

Since atomic dimensions are usually given in Å, we will show this conversion.

$$1.264 \times 10^{-23} \text{ cm}^3 (\text{Ag atom})^{-1} = 1.264 \times 10^{-23} \text{ cm}^3 \times (1 \times 10^8)^3 \text{ Å}^3 \text{ cm}^{-3} \\ (\text{Ag atom})^{-1} = 12.64 \text{ Å}^3 (\text{Ag atom})^{-1}$$

$$12.64 \text{ Å}^3 = 12640000 \text{ pm}^3$$

$$(c) \quad V = \frac{4}{3} \pi r^3; r^3 = \frac{3V}{4\pi}; r = \left(\frac{3V}{4\pi}\right)^{1/3}$$

$$r_A = (3 \times 12.636 \text{ \AA}^3 / 4\pi)^{1/3} = 1.4449 \text{ \AA} = 1.445 \text{ \AA}$$

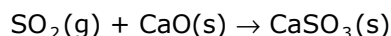
$$1.445 \text{ \AA} = 144.5 \text{ pm}$$

- 3.71 We can proceed by writing the ratio of masses of Ag to AgNO_3 , where y is the atomic mass of nitrogen.

$$\frac{m \text{ Ag}}{m \text{ AgNO}_3} = 0.634985 = \frac{107.8682}{107.8682 + 3(15.9994) + y}$$

Solve for y to obtain $y = 14.0088$. This is to be compared with the currently accepted value of 14.0067.

- 3.72 (a) $\text{S(s)} + \text{O}_2(\text{g}) \rightarrow \text{SO}_2(\text{g})$



Overall equation is; $\text{S(s)} + \text{O}_2(\text{g}) + \text{CaO(s)} \rightarrow \text{CaSO}_3(\text{s})$

$$(b) \quad \text{mol S} = \frac{2000 \times 1 \times 10^6 \text{ g}}{32.1 \text{ g mol}^{-1}} \times \frac{2.5}{100} = 1.56 \times 10^6 \text{ mol}$$

$$\text{g CaSO}_3 = 1.56 \times 10^6 \text{ mol S} \times \frac{1 \text{ mol CaSO}_3}{1 \text{ mol S}} \times 120 \text{ g mol}^{-1} = 1.87 \times 10^8 \text{ g}$$

$$= 1.87 \times 10^5 \text{ kg}$$

This corresponds to about 190 tonnes of CaSO_3 per day as a waste product.