

Thus, there were $2(0.0840) \text{ mol} = 0.1680 \text{ mol}$ of H in the original compound.

That is, $m = nM = (0.1680)(1.01) \text{ g} = 0.1697 \text{ g}$ of H.

The mass of O is found using conservation of mass.

	Mass (g)	n (mol)	ratio to smallest n
compound	2.524		
carbon	1.0105	0.0841	1
hydrogen	0.1697	0.1680	2
oxygen	1.3438	$\frac{1.3438}{16.00} = 0.0839$	1

Thus, the empirical formula is CH_2O .

- (b) Since it contains 12 hydrogen atoms, the molecular formula must be six times the empirical formula or $\text{C}_6\text{H}_{12}\text{O}_6$.

Chapter 6 Review Answers

Student Textbook pages 229–231

Answers to Knowledge/Understanding Questions

1. It is for convenience. Using a one-mole sample allows use of the molar mass values from the periodic table.
2. It is exactly the same according to the law of definite proportions. Each molecule of water has two hydrogen atoms and one oxygen atom.
3. (a) The required measurements are: the initial mass of the compound, the water trap, and the carbon dioxide trap; and the final mass of the water trap and the carbon dioxide trap.
(b) Having the same empirical formula means they have the same percentage composition. Thus, the combustion analysis results would be the same.
4. No, there is insufficient information. The relative proportions of the constituent elements must be known—that is, either percentage composition or empirical formula.

Answers to Inquiry Questions

5. Find the molar mass of the hydrate.

$$\begin{aligned}M &= 2(22.99 \text{ g/mol}) + 4(10.81 \text{ g/mol}) + 7(16.00 \text{ g/mol}) \\ &\quad + 10[2(1.01 \text{ g/mol}) + 16.00 \text{ g/mol}] \\ &= 45.98 \text{ g/mol} + 43.24 \text{ g/mol} + 112.00 \text{ g/mol} + 180.20 \text{ g/mol} \\ &= 201.22 \text{ g/mol} + 180.20 \text{ g/mol} \\ &= 381.42 \text{ g/mol}\end{aligned}$$

$$\text{The percent of anhydrous compound is } \left(\frac{201.22 \text{ g/mol}}{381.42 \text{ g/mol}}\right)(100)\% = 52.8\%$$

The mass of anhydrous compound from 5.00 g of hydrate would be $\left(\frac{52.8}{100}\right)(5.00) \text{ g} = 2.64 \text{ g}$.

6. (a) Find the molar mass for CCl_2F_2 .

$$\begin{aligned}M &= 12.01 \text{ g/mol} + 2(35.45 \text{ g/mol}) + 2(19.00 \text{ g/mol}) \\ &= 12.01 \text{ g/mol} + 70.90 \text{ g/mol} + 38.00 \text{ g/mol} \\ &= 120.91 \text{ g/mol}\end{aligned}$$

$$\% \text{ C} = \left(\frac{12.01 \text{ g/mol}}{120.91 \text{ g/mol}}\right)(100)\% = 9.93\%$$

$$\% \text{ Cl} = \left(\frac{70.90 \text{ g/mol}}{120.91 \text{ g/mol}}\right)(100)\% = 58.6\%$$

$$\% \text{ F} = \left(\frac{38.00 \text{ g/mol}}{120.91 \text{ g/mol}}\right)(100)\% = 31.4\%$$

Freon-12 is 9.93% C; 58.6% Cl; 31.4% F

(b) Find the molar mass for white lead, $\text{Pb}_3(\text{OH})_2(\text{CO}_3)_2$.

$$M = 3(207.20 \text{ g/mol}) + 2(16.00 \text{ g/mol} + 1.01 \text{ g/mol}) + 2[12.01 \text{ g/mol} + 3(16.00 \text{ g/mol})]$$

$$= 775.64 \text{ g/mol}$$

List the masses of the constituent elements for 1 mol of the compound.

Pb: 621.60 g; O: 128.00 g; H: 2.02g; C: 24.02 g

$$\% \text{ Pb} = \left(\frac{621.60 \text{ g/mol}}{775.64 \text{ g/mol}} \right) (100)\% = 80.1\%$$

$$\% \text{ O} = \left(\frac{128.00 \text{ g/mol}}{775.64 \text{ g/mol}} \right) (100)\% = 16.5\%$$

$$\% \text{ H} = \left(\frac{2.02 \text{ g/mol}}{775.64 \text{ g/mol}} \right) (100)\% = 0.260\%$$

$$\% \text{ C} = \left(\frac{24.02 \text{ g/mol}}{775.64 \text{ g/mol}} \right) (100)\% = 3.10\%$$

White lead is 80.1% Pb; 16.5% O; 0.260% H; 3.10% C.

7. (a) Find the molar mass of $\text{MgCl}_2 \cdot 2\text{H}_2\text{O}$.

$$M = 24.31 \text{ g/mol} + 2(35.45 \text{ g/mol}) + 2[2(1.01 \text{ g/mol}) + 16.00 \text{ g/mol}]$$

$$= 95.21 \text{ g/mol} + 36.04 \text{ g/mol}$$

$$= 131.25 \text{ g/mol}$$

$$\% \text{ H}_2\text{O} = \left(\frac{36.04 \text{ g/mol}}{131.25 \text{ g/mol}} \right) (100)\% = 27.46\%$$

The mass of water in 25.0 g of hydrate is $\left(\frac{27.46}{100} \right) (25.0) \text{ g} = 6.86 \text{ g}$

- (b) Find the molar mass of KMnO_4 .

$$M = 39.10 \text{ g/mol} + 54.94 \text{ g/mol} + 4(16.00 \text{ g/mol})$$

$$= 158.04 \text{ g/mol}$$

$$\% \text{ Mn} = \left(\frac{54.94 \text{ g/mol}}{158.04 \text{ g/mol}} \right) (100)\% = 34.8\%$$

Mass of Mn in 5.00 g of KMnO_4 is $\left(\frac{34.8}{100} \right) (5.00) \text{ g} = 1.74 \text{ g}$

8. (a) Find the molar mass of AgNO_3 .

$$M = 107.87 \text{ g/mol} + 14.01 \text{ g/mol} + 3(16.00 \text{ g/mol}) = 169.88 \text{ g/mol}$$

$$\% \text{ Ag} = \left(\frac{107.87 \text{ g/mol}}{169.88 \text{ g/mol}} \right) (100)\% = 63.5\%$$

- (b) Mass of Ag in $(2.00)(10^2) \text{ kg}$ of AgNO_3 is $\left(\frac{63.5}{100} \right) (2.00)(10^2) \text{ kg}$ or $(1.27)(10^2) \text{ kg}$

9. Find the molar mass of BaSO_4 .

$$M = 137.33 \text{ g/mol} + 32.07 \text{ g/mol} + 4(16.00 \text{ g/mol}) = 233.40 \text{ g/mol}$$

$$\% \text{ Ba} = \left(\frac{137.33 \text{ g/mol}}{233.40 \text{ g/mol}} \right) (100)\% = 58.8\%$$

Mass of Ba in 45.8g of BaSO_4 is $\left(\frac{58.8}{100} \right) (45.8) \text{ g} = 26.9 \text{ g}$.

10. Find the molar mass of $\text{Bi}(\text{NO}_3)_2$.

$$M = 208.98 \text{ g/mol} + 2[14.01 \text{ g/mol} + 3(16.00 \text{ g/mol})] = 333.00 \text{ g}$$

$$\% \text{ Bi} = \left(\frac{208.98 \text{ g/mol}}{333.00 \text{ g/mol}} \right) (100)\% = 62.8\%$$

Mass of Bi in 268 g of $\text{Bi}(\text{NO}_3)_2$ is $\left(\frac{62.8}{100} \right) (268) \text{ g} = 168 \text{ g}$.

11. Let M_e and M_m be the molar masses of the empirical and molecular formulas of the compound. Given $M_m \cong 121 \text{ g/mol}$:

from the empirical formula CH_2O ,

$$M_e = 12.01 \text{ g/mol} + 2(1.01 \text{ g/mol}) + 16.00 \text{ g/mol} = 30.03 \text{ g/mol}$$

$$\frac{M_m}{M_e} = \frac{121}{30.03} = \frac{4}{1}$$

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

12. Given $M_m = 322 \text{ g/mol}$, empirical formula $\text{C}_6\text{H}_2\text{OCl}_2$:

$$M_e = 6(12.01 \text{ g/mol}) + 2(1.01 \text{ g/mol}) + 16.00 \text{ g/mol} + 2(35.45 \text{ g/mol})$$

$$= 160.98 \text{ g/mol}$$

$$\text{Ratio } M_m : M_e = 160.98 : 322 = 2 : 1$$

The molecular formula is $\text{C}_{12}\text{H}_4\text{O}_2\text{Cl}_4$.

13. (a) The subscripts represent numbers of atoms, so they must be whole numbers.
 (b) Since 2.67 is the decimal equivalent of 8:3, multiply the subscripts by 3 to obtain C_3H_8 .

14. Consider a 100 g sample.

Element	$n = \frac{m}{M}$ (mol)	Ratio to Smallest n	Revised Ratio
C	$\frac{80.2}{12.01} = 6.677$	10.498	21
O	$\frac{10.18}{16.00} = 0.636$	1.00	2
H	$\frac{9.62}{1.01} = 9.524$	14.975	30

The empirical formula is $\text{C}_{21}\text{O}_2\text{H}_{30}$.

15. Consider a 100 g sample.

Element	$n = \frac{m}{M}$ (mol)	Ratio to Smallest n	Revised Ratio
Na	$\frac{17.6}{22.99} = 0.766$	1.00	2
Cr	$\frac{39.7}{52.00} = 0.763$	1.00	2
O	$\frac{42.8}{16.00} = 2.675$	3.50	7

The empirical formula is $\text{Na}_2\text{Cr}_2\text{O}_7$.

16. Consider a 100 g sample.

Element	$n = \frac{m}{M}$ (mol)	Ratio to Smallest n
Hg	$\frac{67.6}{200.59} = 0.337$	1.00
S	$\frac{10.8}{32.07} = 0.337$	1.00
O	$\frac{21.6}{16.00} = 1.35$	4.00

The empirical formula is HgSO_4 .

17. (a) Consider a 100 g sample.

Element	$n = \frac{m}{M}$ (mol)	Ratio to Smallest n	Revised Ratio
Ca	$\frac{38.8}{40.08} = 0.968$	1.50	3
P	$\frac{20.0}{30.97} = 0.646$	1.00	2
O	$\frac{41.2}{16.00} = 2.575$	3.99	8

The empirical formula is $\text{Ca}_3\text{P}_2\text{O}_8$.

- (b) Since each formula unit contains 2 PO_4^{3-} ions, the molecular formula is $\text{Ca}_3(\text{PO}_4)_2$.

18. (a) Consider a 100 g sample.

Element	$n = \frac{m}{M}$ (mol)	Ratio to Smallest n
C	$\frac{71.0}{12.01} = 5.91$	18.02
H	$\frac{8.60}{1.01} = 8.51$	25.95
O	$\frac{15.8}{16.00} = 0.988$	3.01
N	$\frac{4.60}{14.01} = 0.328$	1.00

The empirical formula is $C_{18}H_{26}O_3N$.

- (b) Since each molecule contains just one nitrogen atom, the molecular formula is the same as the empirical formula.

19. Let the molar mass of x be x g.

Then, for the compound, the molar mass is

$$m = 2x + 5(16.00 \text{ g/mol}) = 2x + 80.00 \text{ g/mol.}$$

$$\text{So, } \% \text{ O} = \left(\frac{80.00}{2x} + 80.00 \right) (100)\% = 44.0\% \text{ (given)}$$

$$\frac{8000}{(2x + 80)} = 44 \text{ and then } 8000 = 44(2x + 80) \text{ or}$$

$$8000 = 88x + 3520 \text{ which gives } 88x = 4480 \text{ or } x = 50.9$$

The molar mass of x is 50.9 g/mol. From the periodic table, x is vanadium.

20. First, find the mass of H in the HCl produced.

	HCl	H
Mass in sample	4.730 g	x g
Mass in one mole	36.46 g	1.01 g

$$\frac{x}{1.01} = \frac{4.730}{36.46} \text{ or } x = 0.1310 \text{ g}$$

Now, find mass of C in the CCl_4 produced.

	CCl_4	C
Mass in sample	9.977 g	y g
Mass in one mole	153.81 g	12.01 g

$$\frac{y}{12.01} = \frac{9.977}{153.81} \text{ or } y = 0.7790 \text{ g}$$

The original compound contains $1.254 \text{ g} - (0.1310 \text{ g} + 0.7790 \text{ g}) = 0.3440 \text{ g}$ of O (using conservation of mass). Thus, the masses are:

Compound	Carbon	Hydrogen	Oxygen
1.254 g	0.7790 g	0.1310 g	0.3440 g

Element	$n = \frac{m}{M}$ (mol)	Ratio to Smallest n
C	$\frac{0.7790}{12.01} = 0.06486$	3
H	$\frac{0.1310}{1.01} = 0.1297$	6
O	$\frac{0.3440}{16.00} = 0.0215$	1

The empirical formula is C_3H_6O .

21. Mass of $FeSO_4 \cdot xH_2O = 2.78 \text{ g}$

$$\text{Mass of } FeSO_4 = 1.52 \text{ g}$$

$$\text{Mass of } H_2O \text{ in sample} = 2.78 \text{ g} - 1.52 \text{ g} = 1.26 \text{ g}$$

$$\begin{aligned} \text{Moles of } FeSO_4 &= \frac{1.52 \text{ g}}{151.92 \text{ g/mol}} \\ &= 0.01 \text{ mol } FeSO_4 \end{aligned}$$

$$\text{Moles of H}_2\text{O} = \frac{1.26 \text{ g}}{18.02 \text{ g/mol}}$$

$$= 0.07 \text{ mol H}_2\text{O}$$

Divide moles of iron sulfate and water by 0.01 mol, this gives the ratio 1 mol FeSO₄ : 7 mol H₂O

Therefore, $x = 7$ and the formula for the compound is FeSO₄ • 7H₂O.

22. (a) First, find the mass of C in the CO₂ produced.

	CO ₂	C
Mass in sample	0.6871 g	x g
Mass in one mole	44.01 g	12.01 g

$$\frac{x}{12.01 \text{ g}} = \frac{0.6871 \text{ g}}{44.01 \text{ g}} \text{ or } x = 0.1875 \text{ g}$$

Now find the mass of H in the H₂O produced.

	H ₂ O	H
Mass in sample	0.1875 g	y g
Mass in one mole	18.02 g	2.02 g

$$\frac{y}{2.02} = \frac{0.1875}{18.02} \text{ or } y = 0.0210 \text{ g}$$

In the original compound using conservation of mass:

	Mass (g)	Mass (%)
compound	0.5000	100
C	0.1875	$\left(\frac{0.1875}{0.500}\right)(100) = 37.50$
H	0.0210	$\left(\frac{0.0210}{0.500}\right)(100) = 4.200$
O	0.2915	$\left(\frac{0.2915}{0.500}\right)(100) = 58.30$

Citric acid is 37.50% C; 4.20% H; 58.30% O.

- (b) Consider 100 g of citric acid.

Element	$n = \frac{m}{M}$ (mol)	Ratio to Smallest n	Revised Ratio
C	$\frac{37.50}{12.01} = 3.122$	1.00	6
H	$\frac{4.20}{1.01} = 4.158$	1.33	8
O	$\frac{58.30}{16.00} = 3.643$	1.17, which is $\frac{7}{6}$	7

The empirical formula is C₆H₈O₇.

- (c) From the empirical formula, $M_e = 6(12.01) + 8(1.01) + 7(16.00) = 192.14 \text{ g/mol}$.
Given $M_m = 192 \text{ g/mol}$: Since $\frac{M_m}{M_e} = \frac{1}{1}$, the molecular formula is the same as the empirical formula.

23. Find the molar mass of CH₃OH.

$$m = 12.01 + 3(1.01) + 16.00 + 1.01 = 32.05 \text{ g/mol}$$

$$n = \frac{m}{M} = \frac{1.00}{32.05} = 0.03120 \text{ mol}$$

1 mole of CH₃OH contains 1 mole of C and 4 moles of H. So the 0.03120 mole sample contains 0.03120 moles of C and 4(0.03120) or 0.1248 moles of H.

When the products are formed, 0.03120 moles of C would yield 0.03120 moles of CO₂ and the 0.1248 moles of H would yield $\frac{0.1248}{2}$ or 0.0624 moles of H₂O.

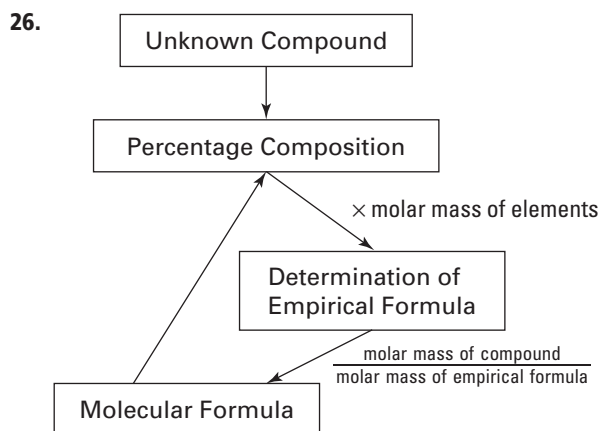
Now use $m = nM$.

Mass of CO₂ produced is (0.03120 mol)(44.01 g/mol) = 1.373 g.

Mass of H₂O produced is (0.0624 mol)(18.02 g/mol) = 1.124 g.

24. (a) Use small quantities to minimize the amount of heat released. Have sand on hand to extinguish any fire that gets out of control. Use safety goggles and gloves.
- (b) Materials: C (powdered charcoal) and Cu_xO . Apparatus: Evaporating dish, retort stand, ring clamp, wire gauze, Bunsen burner, stirring rod, and balance.
- (c) Choose a small sample of Cu_xO (1.00 g perhaps) and calculate how much C would be needed to react with this much Cu_2O (which has a higher percent Cu than the other possibility). Then, choose an amount of C that is in excess to ensure complete reaction of the copper oxide, but not a huge excess that would generate large amounts of heat. Find the mass of an empty evaporating dish. Add the sample of Cu_xO and find the total mass. Add the charcoal and stir. Heat strongly for several minutes until no further reaction occurs. Let dish and contents cool, then find total mass again.
- (d) The data needed for the calculations are:
mass of dish, mass of dish plus Cu_xO , and mass of dish plus Cu.
- (e) This procedure assumes that no Cu was lost from the dish, all the C was removed as CO_2 , and all the Cu_xO reacted.
25. (a) Find the mass of a sample of the substance. Heat the sample. Find the mass again.
- (b) If the sample is a hydrate, the mass should decrease.
- (c) For anhydrous crystals, the mass would be unchanged.

Answers to Communication Questions



27. If a one-mol sample of dimethyl ether ($\text{C}_2\text{H}_6\text{O}$) were passed through a combustion analyzer, the following would result. Once complete combustion of the dimethyl ether has occurred, the resulting carbon dioxide and water vapour will pass into the water-absorbing chamber. At this point, three moles of water would be absorbed by the desiccant, and only the carbon dioxide would pass on to the carbon dioxide absorber. Here, two moles of CO_2 would be absorbed, and any impurities would be passed on to the final filter. From the mass data collected at each absorber, the percentage composition of dimethyl ether could be calculated.

