

## HL First-Year Chemistry

### 1. Atomic Structure and Stoichiometry

Solutions to further problems:

- (a)  ${}_{26}^{55}\text{Fe}$  (b)  ${}_{37}^{86}\text{Rb}$  (c)  ${}_{81}^{123}\text{Tl}$  (d)  ${}_{24}^{123}\text{Cr}$
- Introduction of the source, vaporization of the sample, ionization, acceleration, velocity selection, deflection, and detection.
- $$A_{\text{R}}(\text{Ti}) = \left(\frac{7.95}{100}\right) \times 46 + \left(\frac{7.75}{100}\right) \times 47 + \left(\frac{73.45}{100}\right) \times 48 + \left(\frac{5.51}{100}\right) \times 49 + \left(\frac{5.34}{100}\right) \times 50$$
$$= 47.9254 = 47.9 \text{ or } 48.$$
- Relative atomic masses of some elements are so far off from whole numbers because it represents the average mass of multiple isotopes that make up the element.
- (a) protons = 55, neutrons = 77, electrons = 55,  
(b) protons = 48, neutrons = 67, electrons = 46,  
(c) protons = 81, neutrons = 113, electrons = 81,  
(d) protons = 47, neutrons = 58, electrons = 46,  
(e) protons = 34, neutrons = 44, electrons = 36
- The sample is first vaporized (into gaseous atoms) and then they are ionized, and then accelerated by a strong electric field into an area where they are further exposed to a varying magnetic or electric field ultimately reaching the detector where the charged ions are detected as electric current.
- (a) 2 atoms of Na, 1 atom of C, and 3 atoms of O (b) 3 atoms of N, 12 of H, 1 of P and 4 of O (c) 3 atoms of Na, 1 of Ag, 4 of S, 6 of O
- (a) 2 moles of Na atoms, 1 mole of C atoms, and 3 moles of O atoms (b) 3 moles of N atoms, 12 moles of H atoms, 1 mole of P atoms and 4 moles of O atoms (c) 3 moles of Na atoms, 1 mole of Ag atoms, 4 moles of S atoms, 6 moles of O atoms
- Fe and  $\text{CO}_2$  are the products. The equation is NOT balanced.
- (a) Products:  $\text{CaCl}_2$ ,  $\text{CO}_2$  and  $\text{H}_2\text{O}$ , reactants:  $\text{CaCO}_3$  and  $\text{HCl}$   
(b) equation is unbalanced.  
(c)  $\text{CaCO}_3$  = solid,  $\text{HCl}$  = aqueous,  $\text{CaCl}_2$  = aqueous,  $\text{CO}_2$  = gas, and  $\text{H}_2\text{O}$  = liquid
- $\text{HgS}_{(\text{s})} + \text{O}_{2(\text{g})} \rightarrow \text{Hg}_{(\text{l})} + \text{SO}_{2(\text{g})}$
- (a) there is no unit for relative molecular mass (b) g
- (a) 60.06 g (b) 174.26 g (c) 74.10 g
- a) 
$$\left(\frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}}\right) \left(\frac{1 \text{ mol Fe atoms}}{6.022 \times 10^{23} \text{ atoms of Fe}}\right) = 9.2743 \times 10^{-23} \text{ g Fe/atom of Fe}$$
$$= 9.274 \times 10^{-23} \text{ g Fe/atom of Fe (4 sig figs)}$$
  
b) 
$$(5 \times 10^{10} \text{ molecules N}_2\text{O}) \left(\frac{1 \text{ mol N}_2\text{O}}{6.022 \times 10^{23} \text{ molecules N}_2\text{O}}\right) \left(\frac{44.00 \text{ g N}_2\text{O}}{1 \text{ mol N}_2\text{O}}\right)$$
$$= 3.65 \times 10^{-12} \text{ g N}_2\text{O} = 4 \times 10^{-12} \text{ g N}_2\text{O (1 sig fig)}$$
- Percent composition of:  
a)  $\text{NH}_3$  (ammonia):  $A_{\text{R}}(\text{N}) = 14.01$ ,  $A_{\text{R}}(\text{H}) = 1.01$   
 $M_{\text{R}}(\text{NH}_3) = 14.01 + 3(1.01) = 14.01 + 3.03 = 17.04$  [2 decimal places]  
 $\% \text{ N} = \left(\frac{14.01}{17.04}\right) \times 100 = 82.218 = 82.22$  [4 sig. figs]

$$\% \text{ H} = \left( \frac{3.03}{17.04} \right) \times 100 = 17.781 = 17.8 \text{ [3 sig. figs.]}$$

[**NB**: The total of the percent compositions add up to 100.02, which when rounded off to one decimal place—to be consistent with the percent composition of hydrogen—it comes out to 100.0.]

**OR** considering a mole of ammonia:

Mass of N in a mole of ammonia = 14.01 g

Mass of H in a mole of ammonia = 3 (1.01 g) = 3.03 g

Molar mass of ammonia = 17.04 g

$$\% \text{ N} = \left( \frac{14.01 \text{ g}}{17.04 \text{ g}} \right) \times 100 = 82.218 = 82.22 \text{ [4 sig. figs]}$$

$$\% \text{ H} = \left( \frac{3.03 \text{ g}}{17.04 \text{ g}} \right) \times 100 = 17.781 = 17.8 \text{ [3 sig. figs.]}$$

[**NB**: It makes no difference whether you determine percent compositions using relative mass or molar mass!]

b)  $\text{C}_2\text{H}_4\text{O}_2$  (ethanoic acid):  $A_{\text{R}}(\text{C}) = 12.01$ ,  $A_{\text{R}}(\text{H}) = 1.01$ ,  $A_{\text{R}}(\text{O}) = 16.00$

$M_{\text{R}}(\text{C}_2\text{H}_4\text{O}_2) = 2(12.01) + 4(1.01) + 2(16.00) = 60.06$  [2 decimal places]

$$\% \text{ C} = \left( \frac{24.02}{60.06} \right) \times 100 = 39.99 \text{ [4 sig. figs.]}$$

$$\% \text{ H} = \left( \frac{4.04}{60.06} \right) \times 100 = 6.73 \text{ [3 sig. figs.]}$$

$$\% \text{ O} = \left( \frac{32.00}{60.06} \right) \times 100 = 53.28 \text{ [4 sig. figs.]}$$

c)  $\text{K}_2\text{SO}_4$  (potassium sulfate):  $A_{\text{R}}(\text{K}) = 39.10$ ,  $A_{\text{R}}(\text{S}) = 32.06$ ,  $A_{\text{R}}(\text{O}) = 16.00$

formula mass ( $\text{K}_2\text{SO}_4$ ) =  $2(39.10) + 32.06 + 4(16.00) = 174.26$  [2 decimal places]

$$\% \text{ K} = \left( \frac{78.20}{174.26} \right) \times 100 = 44.88 \text{ [4 sig. figs.]}$$

$$\% \text{ S} = \left( \frac{32.06}{174.26} \right) \times 100 = 18.40 \text{ [4 sig. figs.]}$$

$$\% \text{ O} = \left( \frac{64.00}{174.26} \right) \times 100 = 36.73 \text{ [4 sig. figs.]}$$

d)  $\text{Ca}(\text{OH})_2$ :  $A_{\text{R}}(\text{Ca}) = 40.08$ ,  $A_{\text{R}}(\text{O}) = 16.00$ ,  $A_{\text{R}}(\text{H}) = 1.01$

formula mass ( $\text{Ca}(\text{OH})_2$ ) = 74.10 [2 decimal places]

$$\% \text{ Ca} = \left( \frac{40.08}{74.10} \right) \times 100 = 54.09 \text{ [4 sig. figs.]}$$

$$\% \text{ O} = \left( \frac{32.00}{74.10} \right) \times 100 = 43.18 \text{ [4 sig. figs.]}$$

$$\% \text{ H} = \left( \frac{2.02}{74.10} \right) \times 100 = 2.73 \text{ [3 sig. figs.]}$$

(e)  $\text{Zn}(\text{NO}_3)_2 \cdot 6\text{H}_2\text{O}$ :

formula mass = 297.51 [2 decimal places]

$$\% \text{ Zn} = \left( \frac{65.37}{297.51} \right) \times 100 = 21.97 \text{ [4 sig. figs.]}$$

$$\% \text{ N} = \left( \frac{28.02}{297.51} \right) \times 100 = 9.418 \text{ [4 sig. figs.]}$$

$$\% \text{ O} = \left( \frac{192.00}{297.51} \right) \times 100 = 64.536 \text{ [5 sig. figs.]}$$

$$\% \text{ H} = \left( \frac{12.12}{297.51} \right) \times 100 = 4.074 \text{ [4 sig. figs.]}$$

[NB: The sum of the percentages adds up to 99.998 which when rounded off to 2 decimal places, which is what you have to do since the percentage of zinc is correct only to 2 decimal place, the sum does result in 100.]

f)  $\text{C}_{22}\text{H}_{20}\text{O}_{12}\text{N}_7$

$$M_{\text{R}} (\text{C}_{22}\text{H}_{20}\text{O}_{12}\text{N}_7) = 574.49$$

$$\% \text{ C} = \left( \frac{264.22}{574.49} \right) \times 100 = 45.992 \text{ [5 sig. figs.]}$$

$$\% \text{ H} = \left( \frac{20.20}{574.49} \right) \times 100 = 4.516 \text{ [4 sig. figs.]}$$

$$\% \text{ O} = \left( \frac{192.00}{574.49} \right) \times 100 = 33.421 \text{ [5 sig. figs.]}$$

$$\% \text{ N} = \left( \frac{98.07}{574.49} \right) \times 100 = 17.07 \text{ [4 sig. figs.]}$$

$$16. \text{ a) } 1.00 \text{ lb NH}_3 \left( \frac{1 \text{ kg NH}_3}{2.2 \text{ lb NH}_3} \right) \left( \frac{1000 \text{ g NH}_3}{1 \text{ kg NH}_3} \right) \left( \frac{1 \text{ mol NH}_3}{17.04 \text{ g NH}_3} \right) = 26.7 \text{ mol NH}_3 \text{ [3 sig figs]}$$

$$\text{b) } 1.00 \text{ lb C}_2\text{H}_4\text{O}_2 \left( \frac{1 \text{ kg C}_2\text{H}_4\text{O}_2}{2.2 \text{ lb C}_2\text{H}_4\text{O}_2} \right) \left( \frac{1000 \text{ g C}_2\text{H}_4\text{O}_2}{1 \text{ kg C}_2\text{H}_4\text{O}_2} \right) \left( \frac{1 \text{ mol C}_2\text{H}_4\text{O}_2}{60.06 \text{ g C}_2\text{H}_4\text{O}_2} \right) = 7.50 \text{ mol C}_2\text{H}_4\text{O}_2 \text{ [3}$$

sig  
figs]

$$\text{c) } 1.00 \text{ lb K}_2\text{SO}_4 \left( \frac{1 \text{ kg K}_2\text{SO}_4}{2.2 \text{ lb K}_2\text{SO}_4} \right) \left( \frac{1000 \text{ g K}_2\text{SO}_4}{1 \text{ kg K}_2\text{SO}_4} \right) \left( \frac{1 \text{ mol K}_2\text{SO}_4}{174.26 \text{ g K}_2\text{SO}_4} \right) = 2.61 \text{ mol K}_2\text{SO}_4 \text{ [3 sig$$

figs]

$$\text{d) } 1.00 \text{ lb Ca(OH)}_2 \left( \frac{1 \text{ kg Ca(OH)}_2}{2.2 \text{ lb Ca(OH)}_2} \right) \left( \frac{1000 \text{ g Ca(OH)}_2}{1 \text{ kg Ca(OH)}_2} \right) \left( \frac{1 \text{ mol Ca(OH)}_2}{74.10 \text{ g Ca(OH)}_2} \right) = 6.13 \text{ mol}$$

$\text{Ca(OH)}_2$

[3 sig figs]

$$\text{e) } 1.00 \text{ lb Zn(NO}_3)_2 \cdot 6\text{H}_2\text{O} \left( \frac{1 \text{ kg Zn(NO}_3)_2 \cdot 2\text{H}_2\text{O}}{2.2 \text{ lb Zn(NO}_3)_2 \cdot 2\text{H}_2\text{O}} \right) \left( \frac{1000 \text{ g Zn(NO}_3)_2 \cdot 2\text{H}_2\text{O}}{1 \text{ kg Zn(NO}_3)_2 \cdot 2\text{H}_2\text{O}} \right)$$

$$\left( \frac{1 \text{ mol Zn(NO}_3)_2 \cdot 2\text{H}_2\text{O}}{297.51 \text{ g Zn(NO}_3)_2 \cdot 2\text{H}_2\text{O}} \right) = 1.53 \text{ mol Zn(NO}_3)_2 \cdot 2\text{H}_2\text{O} \text{ [3 sig figs]}$$

$$\text{f) } 1.00 \text{ lb C}_{22}\text{H}_{20}\text{O}_{12}\text{N}_7 \left( \frac{1 \text{ kg C}_{22}\text{H}_{20}\text{O}_{12}\text{N}_7}{2.2 \text{ lb C}_{22}\text{H}_{20}\text{O}_{12}\text{N}_7} \right) \left( \frac{1000 \text{ g C}_{22}\text{H}_{20}\text{O}_{12}\text{N}_7}{1 \text{ kg C}_{22}\text{H}_{20}\text{O}_{12}\text{N}_7} \right) \left( \frac{1 \text{ mol C}_{22}\text{H}_{20}\text{O}_{12}\text{N}_7}{574.49 \text{ g C}_{22}\text{H}_{20}\text{O}_{12}\text{N}_7} \right)$$

$$= 0.791 \text{ mol C}_2\text{H}_4\text{O}_2 \text{ [3 sig figs]}$$

$$17. \text{ (a) } (15.4 \text{ g Al}) \left( \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \right) \left( \frac{6.022 \times 10^{23} \text{ atoms Al}}{1 \text{ mol Al}} \right) = 3.43 \times 10^{23} \text{ atoms Al (3 sig figs)}$$

$$(b) (14.8 \text{ g N}_2\text{O}_5) \left( \frac{1 \text{ mol N}_2\text{O}_5}{108.00 \text{ g N}_2\text{O}_5} \right) \left( \frac{6.022 \times 10^{23} \text{ molecules N}_2\text{O}_5}{1 \text{ mol N}_2\text{O}_5} \right)$$

$$= 8.25 \times 10^{22} \text{ molecules of N}_2\text{O}_5$$

$$(c) (83.2 \text{ g Na}_2\text{CO}_3) \left( \frac{1 \text{ mol Na}_2\text{CO}_3}{105.99 \text{ g Na}_2\text{CO}_3} \right) \left( \frac{2 \text{ mol Na ions}}{1 \text{ mol Na}_2\text{CO}_3} \right) \left( \frac{6.022 \times 10^{23} \text{ Na ions}}{1 \text{ mol Na ions}} \right)$$

$$18. (7 \text{ mol C}) \left( \frac{1 \text{ mol C}_3\text{H}_8\text{O}_3}{3 \text{ mol C}} \right) \left( \frac{3 \text{ mol O}}{1 \text{ mol C}_3\text{H}_8\text{O}_3} \right) = 7 \text{ mol O, which you should have been able to}$$

determine without even setting it up as shown above. Since in a mole of glycerol the number of moles of C and O are equal (3:3) any sample of glycerol would have equal number of moles of C and O.

$$19. \text{Equation is: } 2\text{Cu}_{(s)} + \text{S}_{(aq)} \rightarrow \text{Cu}_2\text{S}_{(s)}$$

$$(100.0 \text{ g Cu}) \left( \frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}} \right) \left( \frac{1 \text{ mol S}}{2 \text{ mol Cu}} \right) \left( \frac{32.06 \text{ g S}}{1 \text{ mol S}} \right) = 25.224 = 25.22 \text{ g S}$$

Limiting reagent is copper.

$$(100.0 \text{ g Cu}) \left( \frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}} \right) \left( \frac{1 \text{ mol Cu}_2\text{S}}{2 \text{ mol Cu}} \right) \left( \frac{159.16 \text{ g Cu}_2\text{S}}{1 \text{ mol Cu}_2\text{S}} \right) = 125.22 = 125.2 \text{ g. (4 sig figs)}$$

$$20. \text{Equation is: } \text{C}_{(s)} + 2\text{Cl}_2_{(g)} \rightarrow \text{CCl}_4_{(l)}$$

$$(10.0 \text{ g C}) \left( \frac{1 \text{ mol C}}{12.01 \text{ g C}} \right) \left( \frac{2 \text{ mol Cl}_2}{1 \text{ mol C}} \right) \left( \frac{70.90 \text{ g Cl}_2}{1 \text{ mol Cl}_2} \right) = 59.0 \text{ Cl}_2$$

Limiting reagent is carbon.

$$(10.0 \text{ g C}) \left( \frac{1 \text{ mol C}}{12.01 \text{ g C}} \right) \left( \frac{1 \text{ mol CCl}_4}{1 \text{ mol C}} \right) \left( \frac{141.80 \text{ g CCl}_4}{1 \text{ mol CCl}_4} \right) = 118.0 = 118 \text{ g CCl}_4 \text{ (3 sig figs)}$$

$$\text{Mass of excess reagent left unreacted} = 100.0 \text{ g} - 59.0 \text{ g} = 41.0 \text{ g}$$

$$21. \text{The equation for the reaction is: } 4\text{Fe}_{(s)} + 3\text{O}_2_{(g)} \rightarrow 2\text{Fe}_2\text{O}_3_{(s)}$$

$$(4.80 \text{ g O}_2) \left( \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} \right) \left( \frac{4 \text{ mol Fe}}{3 \text{ mol O}_2} \right) \left( \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} \right) = 11.17 \text{ g Fe} = 11.2 \text{ g Fe (required)}$$

$$(0.150 \text{ mol Fe}) \left( \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} \right) = 8.3775 \text{ g Fe} = 8.38 \text{ g Fe (available)}$$

Therefore, limiting reagent is Fe.

$$(0.150 \text{ mol Fe}) \left( \frac{2 \text{ mol Fe}_2\text{O}_3}{4 \text{ mol Fe}} \right) \left( \frac{159.70 \text{ g Fe}_2\text{O}_3}{1 \text{ mol Fe}_2\text{O}_3} \right) = 11.977 \text{ g Fe}_2\text{O}_3 = 12.0 \text{ g Fe}_2\text{O}_3 \text{ (produced)}$$

Mass of Fe remaining = 0.

$$(0.150 \text{ mol Fe}) \left( \frac{3 \text{ mol O}_2}{4 \text{ mol Fe}} \right) \left( \frac{16.00 \text{ g O}_2}{1 \text{ mol O}_2} \right) = 1.8 \text{ g O}_2 = 1.80 \text{ g O}_2$$

$$\text{Mass of oxygen remaining} = 4.80 - 1.80 = 3.00 \text{ g (2 dp)}$$

$$22. (a) \text{NH}_2\text{SO}_4, (b) \text{CH}_4, (c) \text{C}_3\text{H}_8\text{O}_3, (d) \text{Hg}_2\text{SO}_4.$$

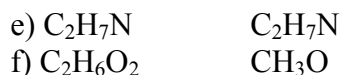
$$23. \quad \begin{array}{ll} \text{Molecular} & \text{Empirical formula} \end{array}$$

$$a) \text{C}_3\text{H}_7\text{O}_2 \quad \text{C}_3\text{H}_7\text{O}_2$$

$$b) \text{CH}_4\text{O} \quad \text{CH}_4\text{O}$$

$$c) \text{C}_4\text{H}_8\text{O}_2 \quad \text{C}_2\text{H}_4\text{O}$$

$$d) \text{C}_2\text{H}_6\text{O} \quad \text{C}_2\text{H}_6\text{O}$$



24.	Tungsten	Oxygen
Mass ratio	4.23 g	5.34 – 4.23 = 1.11 g
Molar ratio	$4.23 \text{ g} \left( \frac{1 \text{ mol}}{183.85 \text{ g}} \right) = 0.02300 \text{ mol}$	$1.11 \text{ g} \left( \frac{1 \text{ mol}}{16.00 \text{ g}} \right) = 0.06937 \text{ mol}$
Dividing by the smaller of the two numbers (to get whole number ratios)	$\left( \frac{0.02300}{0.02300} \right) = 1$	$\left( \frac{0.06937}{0.02300} \right) = 3.016 \approx 3$

Therefore the empirical formula is WO<sub>3</sub>.

25. a)	C	O
Percentage	27.3%	72.7%
Then in 100 g	42.9 g	57.1 g
Molar ratio	$(42.9 \text{ g}) \left( \frac{1 \text{ mol}}{12.01 \text{ g}} \right) = 3.57 \text{ mol}$	$(57.1 \text{ g}) \left( \frac{1 \text{ mol}}{16.00 \text{ g}} \right) = 3.57 \text{ mol}$
Number of moles of O that combine with 1 mol of C	$\left( \frac{3.57}{3.57} \right) = 1$	$\left( \frac{3.57}{3.57} \right) = 1$

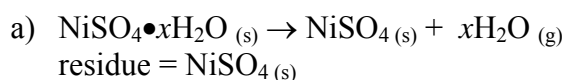
Therefore the empirical formula is CO.

b) CO<sub>2</sub> c) NaSO<sub>4</sub> d) Na<sub>2</sub>S<sub>2</sub>O<sub>3</sub>.

26. There are a couple of different ways of tackling these kinds of questions.

**Method A.** The following is the long version with explanations. And it uses the stoichiometric ratio between the anhydrous salt and the hydrated form.

What the solution boils down to is determination of the number of moles of water produced by a mole of the hydrate knowing each of the moles of water and that of the hydrate in the sample of hydrate analyzed.



$$0.306 \text{ g NiSO}_4 \left( \frac{1 \text{ mol NiSO}_4}{154.77 \text{ g NiSO}_4} \right) = 0.001977 \text{ mol NiSO}_4$$

Since 1 mole of NiSO<sub>4</sub>•xH<sub>2</sub>O produces 1 mole of NiSO<sub>4</sub>, number of moles of NiSO<sub>4</sub>•xH<sub>2</sub>O in the

original sample = number of moles of NiSO<sub>4</sub> produced = 0.001977.

Since the original mass was 0.520 g, then the mass of 1 mole of NiSO<sub>4</sub>•xH<sub>2</sub>O is given by:

$$\left( \frac{0.520 \text{ g NiSO}_4 \cdot x\text{H}_2\text{O}}{0.001977 \text{ mol NiSO}_4 \cdot x\text{H}_2\text{O}} \right) = 263 \text{ g/mol NiSO}_4 \cdot x\text{H}_2\text{O}$$

Mass of water in 1 mole of NiSO<sub>4</sub>•xH<sub>2</sub>O = molar mass of NiSO<sub>4</sub>•xH<sub>2</sub>O – molar mass of NiSO<sub>4</sub> (since molar ratio of NiSO<sub>4</sub>•xH<sub>2</sub>O: NiSO<sub>4</sub> = 1:1)

$$= 263 - 154.77 = 108 \text{ g}$$

$$\text{Number of moles of water in 108 g of water} = 108 \text{ g H}_2\text{O} \left( \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) = 5.99 = 6$$

Therefore, x = 6. That is, in every mole of NiSO<sub>4</sub>•xH<sub>2</sub>O, there are 6 moles of water molecules.

**Method B.** This method uses the relationship between the products:

What the solution boils down to is determination of the number of moles of water produced for every mole of the anhydrous salt produced knowing each of the moles of water and that of the anhydrous salt in the sample of hydrate analyzed.

From the equation, the ratio of the products  $\text{NiSO}_4$  to  $\text{H}_2\text{O} = 1$  to  $x$

$$\text{That is, } \left( \frac{\text{moles of NiSO}_4}{\text{moles of H}_2\text{O}} \right) = \left( \frac{1 \text{ mole of NiSO}_4}{x} \right)$$

The ratio of  $\left( \frac{\text{moles of NiSO}_4}{\text{moles of H}_2\text{O}} \right)$  in the experiment must also be the same.

Therefore,

$$\left( \frac{\text{experimental moles of NiSO}_4}{\text{experimental moles of H}_2\text{O}} \right) = \left( \frac{1 \text{ mole of NiSO}_4}{x} \right) \quad (\text{Equation I})$$

Moles of  $\text{NiSO}_4$  produced in the experiment

$$0.306 \text{ g NiSO}_4 \left( \frac{1 \text{ mol NiSO}_4}{154.77 \text{ g NiSO}_4} \right) = 0.001978 \text{ mol NiSO}_4$$

Mass of water in the experiment =  $(0.520 - 0.306) \text{ g} = 0.214 \text{ g}$

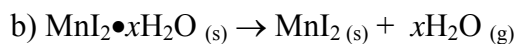
Moles of  $\text{H}_2\text{O}$  produced in the experiment:

$$0.214 \text{ g H}_2\text{O} \left( \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) = 0.0118 \text{ mol H}_2\text{O}$$

Substituting in equation I, and solving for  $x$ ,

$$\left( \frac{0.00198 \text{ mole NiSO}_4}{0.0118 \text{ mole H}_2\text{O}} \right) = \left( \frac{1 \text{ mole of NiSO}_4}{x} \right)$$

$$\Rightarrow x = (1 \text{ mole of NiSO}_4) \left( \frac{0.0118 \text{ mole H}_2\text{O}}{0.00198 \text{ mole of NiSO}_4} \right) = 5.95 \text{ mol H}_2\text{O} = 6 \text{ mole H}_2\text{O}$$



$$x = (1 \text{ mole of MnI}_2) \left( \frac{\text{moles of H}_2\text{O produced}}{\text{moles of MnI}_2 \text{ produced}} \right)$$

Mass of water produced =  $(0.895 - 0.726) \text{ g} = 0.169 \text{ g}$

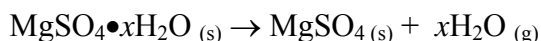
$$0.169 \text{ g H}_2\text{O} \left( \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) = 0.00937 \text{ mol H}_2\text{O}$$

Moles of  $\text{MnI}_2$  produced

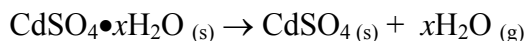
$$0.726 \text{ g MnI}_2 \left( \frac{1 \text{ mol MnI}_2}{308.74 \text{ g MnI}_2} \right) = 0.00235 \text{ mol MnI}_2$$

$$x = (1 \text{ mole of MnI}_2) \left( \frac{0.00937 \text{ moles of H}_2\text{O}}{0.00235 \text{ moles of MnI}_2} \right) = 3.98 \text{ mol H}_2\text{O} = 4 \text{ mol H}_2\text{O}$$

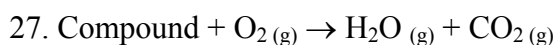
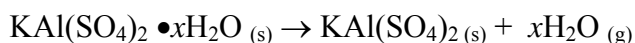
c) similar to the ones above.



d) similar to the ones above.



e) Similar to the ones above.



Mass of compound = 12.13 mg

Mass of H<sub>2</sub>O produced = 5.36 mg

Mass of hydrogen in the sample of water produced

$$5.36 \text{ mg H}_2\text{O} \left( \frac{1 \text{ g H}_2\text{O}}{1000 \text{ mg H}_2\text{O}} \right) \left( \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) \left( \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \right) \left( \frac{1.01 \text{ g H}}{1 \text{ mol H}} \right) = 0.000601 \text{ g H}$$

Mass of CO<sub>2</sub> produced = 30.6 mg

Mass of carbon in the carbon dioxide produced

$$30.6 \text{ mg CO}_2 \left( \frac{1 \text{ g CO}_2}{1000 \text{ mg CO}_2} \right) \left( \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} \right) \left( \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \right) \left( \frac{12.01 \text{ g C}}{1 \text{ mol C}} \right) = 0.00835 \text{ g C}$$

Since all the hydrogen in water and carbon in carbon dioxide came from the original organic compound, the mass of hydrogen in the original compound = 0.000601 g, and mass of carbon in the

original compound = 0.00835 g.

Therefore,

$$\begin{aligned} \text{mass of oxygen in the original compound} &= 0.01213 \text{ g} - (0.000601 + 0.00835) \text{ g} \\ &= 0.01213 \text{ g} - 0.00895 \text{ g} \\ &= 0.00318 \text{ g} \end{aligned}$$

	C	H	O
Mass ratio	0.00835	0.000601 g	0.00318 g
Molar ratio	$(0.00835 \text{ g}) \left( \frac{1 \text{ mol}}{12.01 \text{ g}} \right)$	$(0.000601 \text{ g}) \left( \frac{1 \text{ mol}}{1.01 \text{ g}} \right)$	$(0.00318 \text{ g}) \left( \frac{1 \text{ mol}}{16.00 \text{ g}} \right)$
	= 0.0006952 mol	= 0.0005950 mol	= 0.0001987 mol
Moles of C and H that combines with 1 mol of O	$\left( \frac{0.0006952}{0.0001987} \right) = 3.498$	$\left( \frac{0.0005950}{0.0001987} \right) = 2.994$	$\left( \frac{0.0001987}{0.0001987} \right) = 1$
Multiplying by 2	3.498 X 2 = 6.996 ≈ 7	2.994 X 2 = 5.988 ≈ 6	1 X 2 = 2

Therefore, the empirical formula is C<sub>7</sub>H<sub>6</sub>O<sub>2</sub>.

28. Out of 100.0 g compound:  $30.4 \text{ g N} \left( \frac{1 \text{ mol N}}{14.01 \text{ g N}} \right) = 2.17 \text{ mol N}$

$$\% \text{ O} = 100.0 - 30.4 = 69.6\%$$

$$69.6 \text{ g O} \left( \frac{1 \text{ mol O}}{16.00 \text{ g O}} \right) = 4.35 \text{ mol O}$$

$$\text{Number of moles of O per mol of N} = \frac{4.35}{2.17} = 2.00$$

Empirical formula = NO<sub>2</sub>

$$\text{Empirical formula mass} = 14.01 + 2(16.00) = 46.01 \text{ g/mol}$$

$$\frac{92 \text{ g}}{46.01 \text{ g}} = 2.0;$$

Therefore, the molecular formula is  $\text{N}_2\text{O}_4$ .

29. Assuming 100.00 g of compound

$$\text{Mass of hydrogen} = 100.00 - 49.31 - 43.79 = 6.90 \text{ g}$$

$$49.31 \text{ g C} \left( \frac{1 \text{ mol C}}{12.01 \text{ g C}} \right) = 4.106 \text{ mol C}; 43.79 \text{ g O} \left( \frac{1 \text{ mol O}}{16.00 \text{ g O}} \right) = 2.737 \text{ mol O};$$

$$6.90 \text{ g H} \left( \frac{1 \text{ mol H}}{1.01 \text{ g H}} \right) = 6.83 \text{ mol H}$$

Dividing all mole values by 2.737 to get the the number of moles of C and H per mole of O

$$\frac{4.106}{2.737} = 1.500 \text{ C}; \frac{6.83}{2.737} = 2.495 \text{ H}$$

Multiplying the ratio by 2 yields whole number ratios giving an empirical formula  $\text{C}_3\text{H}_5\text{O}_2$ .

$$\text{Empirical formula mass} = 3(12.01) + 5(1.01) + 2(16.00) = 73.08 \text{ g/mol}$$

$$\frac{\text{molar mass}}{\text{empirical formula mass}} = \frac{146.1}{73.08} = 1.999 = 2.00$$

Therefore, the molecular formula is  $\text{C}_6\text{H}_{10}\text{O}_4$ .

30. Assuming 100.00 g of compound

$$\text{Mass of oxygen} = 100.00 - 41.39 - 3.47 = 55.14 \text{ g}$$

$$41.39 \text{ g C} \left( \frac{1 \text{ mol C}}{12.01 \text{ g C}} \right) = 3.446 \text{ mol C}; 55.14 \text{ g O} \left( \frac{1 \text{ mol O}}{16.00 \text{ g O}} \right) = 3.446 \text{ mol O};$$

$$3.47 \text{ g H} \left( \frac{1 \text{ mol H}}{1.01 \text{ g H}} \right) = 3.435 \text{ mol H}$$

All three are the same mole values, therefore the empirical formula is CHO.

$$\text{Empirical formula mass} = 12.01 + 1.01 + 16.00 = 29.02 \text{ g/mol}$$

$$\text{Molar mass of compound} = \frac{15.0 \text{ g}}{0.129 \text{ mol}} = 116 \text{ g/mol}$$

$$\frac{\text{molar mass}}{\text{empirical formula mass}} = \frac{116}{29.02} = 4.00$$

Therefore, the molecular formula is  $\text{C}_4\text{H}_4\text{O}_4$ .