<u>**HL First-Year Chemistry**</u>

1. Atomic Structure and Stoichiometry

Solutions to further problems:

1. (a) ${}^{55}_{26}Fe$ (b) ${}^{86}_{37}Rb$ (c) ${}^{123}_{81}Tl$ (d) ${}^{123}_{24}Cr$
2. Introduction of the source, vaporization of the sample, ionization, acceleration, velocity selection, deflection, and detection.
3. A_R (Ti) = $\left(\frac{7.95}{100}\right)$ X 46 + $\left(\frac{7.75}{100}\right)$X 47 + $\left(\frac{73.45}{100}\right)$X 48 + $\left(\frac{5.51}{100}\right)$ X 49 + $\left(\frac{5.34}{100}\right)$ X 50
   = 47.9254 = 47.9 or 48.
4. 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.
5. (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
6. 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.
7. (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
8. (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
9. Fe and CO_2 are the products. The equation is NOT balanced.
10. (a) Products: $CaCl_2$, CO_2 and H_2O, reactants: $CaCO_3$ and HCl
    (b) equation is unbalanced.
    (c) $CaCO_3$ = solid, HCl = aqueous, $CaCl_2$ = aqueous, CO_2 = gas, and H_2O = liquid
11. $HgS_{(s)} + O_{2\,(g)} \rightarrow Hg_{(l)} + SO_{2\,(g)}$
12. (a) there is no unit for relative molecular mass (b) g
13. (a) 60.06 g (b) 174.26 g (c) 74.10 g
14. a) $\left(\frac{55.85\text{ g Fe}}{1\text{ mol Fe}}\right)\left(\frac{1\text{ mol Fe atoms}}{6.022\text{ X }10^{23}\text{ atoms of Fe}}\right)$= 9.2743 X 10^{-23} g Fe/atom of Fe

    = 9.274 X 10^{-23} g Fe/atom of Fe (4 sig figs)

    b) (5 X 10^{10} molecules N_2O)$\left(\frac{1\text{ mol }N_2O}{6.022\text{ X }10^{23}\text{ molecules }N_2O}\right)\left(\frac{44.00\text{ g }N_2O}{1\text{ mol }N_2O}\right)$

    = 3.65 X 10^{-12} g N_2O = 4 X 10^{-12} g N_2O (1 sig fig)
15. Percent composition of:
    a) NH_3 (ammonia): A_R (N) = 14.01, A_R (H) = 1.01
       M_R (NH_3) = 14.01 + 3(1.01) = 14.01 + 3.03 = 17.04 [2 decimal places]
       % N = $\left(\frac{14.01}{17.04}\right)$ X 100 = 82.218 = 82.22 [4 sig. figs]

$\% H = \left(\frac{3.03}{17.04}\right)$ X 100 = 17.781 = 17.8 [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

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

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

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

b) $C_2H_4O_2$ (ethanoic acid): A_R (C) = 12.01, A_R (H) = 1.01, A_R (O) = 16.00

M_R ($C_2H_4O_2$) = 2(12.01) + 4(1.01) + 2(16.00) = 60.06 [2 decimal places]

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

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

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

c) K_2SO_4 (potassium sulfate): A_R (K) = 39.10, A_R (S) = 32.06, A_R (O) = 16.00

formula mass (K_2SO_4) = 2(39.10) + 32.06 + 4(16.00) = 174.26 [2 decimal places]

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

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

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

d) $Ca(OH)_2$: A_R (Ca) = 40.08, A_R (O) = 16.00, A_R (H) = 1.01

formula mass ($Ca(OH)_2$) = 74.10 [2 decimal places]

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

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

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

(e) $Zn(NO_3)_2 \bullet 6H_2O$:

formula mass = 297.51 [2 decimal places]

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

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

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

$\% H = \left(\frac{12.12}{297.51}\right) X 100 = 4.074$ [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) $C_{22}H_{20}O_{12}N_7$

M_R ($C_{22}H_{20}O_{12}N_7$) = 574.49

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

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

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

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

16. a) 1.00 lb NH_3 $\left(\frac{1\,kg\ NH_3}{2.2\,lb\ NH_3}\right)\left(\frac{1000\,g\ NH_3}{1\,kg\ NH_3}\right)\left(\frac{1\,mol\ NH_3}{17.04\,g\ NH_3}\right)$ = 26.7 mol NH_3 [3 sig figs]

b) 1.00 lb $C_2H_4O_2$ $\left(\frac{1\,kg\ C_2H_4O_2}{2.2\,lb\ C_2H_4O_2}\right)\left(\frac{1000\,g\ C_2H_4O_2}{1\,kg\ C_2H_4O_2}\right)\left(\frac{1\,mol\ C_2H_4O_2}{60.06\,g\ C_2H_4O_2}\right)$ = 7.50 mol $C_2H_4O_2$ [3 sig figs]

c) 1.00 lb K_2SO_4 $\left(\frac{1\,kg\ K_2SO_4}{2.2\,lb\ K_2SO_4}\right)\left(\frac{1000\,g\ K_2SO_4}{1\,kg\ K_2SO_4}\right)\left(\frac{1\,mol\ K_2SO_4}{174.26\,g\ K_2SO_4}\right)$ = 2.61 mol K_2SO_4 [3 sig figs]

d) 1.00 lb $Ca(OH)_2$ $\left(\frac{1\,kg\ Ca(OH)_2}{2.2\,lb\ Ca(OH)_2}\right)\left(\frac{1000\,g\ Ca(OH)_2}{1\,kg\ Ca(OH)_2}\right)\left(\frac{1\,mol\ Ca(OH)_2}{74.10\,g\ Ca(OH)_2}\right)$ = 6.13 mol $Ca(OH)_2$
[3 sig figs]

e) 1.00 lb $Zn(NO_3)_2{\bullet}6H_2O$ $\left(\frac{1\,kg\ Zn(NO_3)_2 \bullet 2H_2O}{2.2\,lb\ Zn(NO_3)_2 \bullet 2H_2O}\right)\left(\frac{1000\,g\ Zn(NO_3)_2 \bullet 2H_2O}{1\,kg\ Zn(NO_3)_2 \bullet 2H_2O}\right)$

$\left(\frac{1\,mol\ Zn(NO_3)_2 \bullet 2H_2O}{297.51\,g\ Zn(NO_3)_2 \bullet 2H_2O}\right)$ = 1.53 mol $Zn(NO_3)_2 \bullet 2H_2O$ [3 sig figs]

f) 1.00 lb $C_{22}H_{20}O_{12}N_7$ $\left(\frac{1\,kg\ C_{22}H_{04}O_{12}N_7}{2.2\,lb\ C_{22}H_{04}O_{12}N_7}\right)\left(\frac{1000\,g\ C_{22}H_{04}O_{12}N_7}{1\,kg\ C_{22}H_{04}O_{12}N_7}\right)\left(\frac{1\,mol\ C_{22}H_{04}O_{12}N_7}{574.49\,g\ C_{22}H_{04}O_{12}N_7}\right)$

= 0.791 mol $C_2H_4O_2$ [3 sig figs]

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

(b) $(14.8\text{ g }N_2O_5)\left(\frac{1\text{ mol }N_2O_5}{108.00\text{ g }N_2O_5}\right)\left(\frac{6.022\text{ X }10^{23}\text{ molecules }N_2O_5}{1\text{ mol }N_2O_5}\right)$

$= 8.25\text{ X }10^{22}$ molecules of N_2O_5

(c) $(83.2\text{ g }Na_2CO_3)\left(\frac{1\text{ mol }Na_2CO_3}{105.99\text{ g }Na_2CO_3}\right)\left(\frac{2\text{ mol Na ions}}{1\text{ mol }Na_2CO_3}\right)\left(\frac{6.022\text{ X }10^{23}\text{ Na ions}}{1\text{ mol Na ions}}\right)$

18. $(7\text{ mol C})\left(\frac{1\text{ mol }C_3H_8O_3}{3\text{ mol C}}\right)\left(\frac{3\text{ mol O}}{1\text{ mol }C_3H_8O_3}\right) = 7$ 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. Equation is: $2Cu_{(s)} + S_{(aq)} \rightarrow Cu_2S_{(s)}$

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

Limiting reagent is copper.

$(100.0\text{ g Cu})\left(\frac{1\text{ mol Cu}}{63.55\text{ g Cu}}\right)\left(\frac{1\text{ mol }Cu_2S}{2\text{ mol Cu}}\right)\left(\frac{159.16\text{ g }Cu_2S}{1\text{ mol }Cu_2S}\right) = 125.22 = 125.2\text{ g.}$ (4 sig figs)

20. Equation is: $C_{(s)} + 2Cl_{2\,(g)} \rightarrow CCl_{4\,(l)}$

$(10.0\text{ g C})\left(\frac{1\text{ mol C}}{12.01\text{ g C}}\right)\left(\frac{2\text{ mol }Cl_2}{1\text{ mol C}}\right)\left(\frac{70.90\text{ g }Cl_2}{1\text{ mol }Cl_2}\right) = 59.0\ Cl_2$

Limiting reagent is carbon.

$(10.0\text{ g C})\left(\frac{1\text{ mol C}}{12.01\text{ g C}}\right)\left(\frac{1\text{ mol }CCl_4}{1\text{ mol C}}\right)\left(\frac{141.80\text{ g }CCl_4}{1\text{ mol }CCl_4}\right) = 118.0 = 118\text{ g }CCl_4$ (3 sig figs)

Mass of excess reagent left unreacted = 100.0 g – 59.0 g = 41.0 g

21. The equation for the reaction is: $4Fe_{(s)} + 3O_{2\,(g)} \rightarrow 2Fe_2O_{3\,(s)}$

$(4.80\text{ g }O_2)\left(\frac{1\text{ mol }O_2}{32.00\text{ g }O_2}\right)\left(\frac{4\text{ mol Fe}}{3\text{ mol }O_2}\right)\left(\frac{55.85\text{ g Fe}}{1\text{ mol Fe}}\right) = 11.17\text{ g Fe} = 11.2\text{ g Fe}$ (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.

$(0.150\text{ mol Fe})\left(\frac{2\text{ mol }Fe_2O_3}{4\text{ mol Fe}}\right)\left(\frac{159.70\text{ g }Fe_2O_3}{1\text{ mol }Fe_2O_3}\right) = 11.977\text{ g }Fe_2O_3 = 12.0\text{ g }Fe_2O_3$ (produced)

Mass of Fe remaining = 0.

$(0.150\text{ mol Fe})\left(\frac{3\text{ mol }O_2}{4\text{ mol Fe}}\right)\left(\frac{16.00\text{ g }O_2}{1\text{ mol }O_2}\right) = 1.8\text{ g }O_2 = 1.80\text{ g }O_2$

Mass of oxygen remaining = 4.80 – 1.80 = 3.00 g (2 dp)

22. (a) NH_2SO_4, (b) CH, (c) $C_3H_8O_3$, (d) Hg_2SO_4.

23.

Molecular	Empirical formula
a) $C_3H_7O_2$	$C_3H_7O_2$
b) CH_4O	CH_4O
c) $C_4H_8O_2$	C_2H_4O
d) C_2H_6O	C_2H_6O

e) C_2H_7N C_2H_7N
f) $C_2H_6O_2$ CH_3O

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_3.

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_2 c) $NaSO_4$ d) $Na_2S_2O_3$.

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.

a) $NiSO_4 \bullet xH_2O_{(s)} \rightarrow NiSO_{4\,(s)} + xH_2O_{(g)}$
residue = $NiSO_{4\,(s)}$

$0.306\text{ g }NiSO_4\left(\frac{1\text{ mol }NiSO_4}{154.77\text{ g }NiSO_4}\right) = 0.001977\text{ mol }NiSO_4$

Since 1 mole of $NiSO_4 \bullet xH_2O$ produces 1 mole of $NiSO_4$, number of moles of $NiSO_4 \bullet xH_2O$ in the
original sample = number of moles of $NiSO_4$ produced = 0.001977.

Since the original mass was 0.520 g, then the mass of 1 mole of $NiSO_4 \bullet xH_2O$ is given by:

$\left(\frac{0.520\text{ g }NiSO_4 \bullet xH_2O}{0.001977\text{ mol }NiSO_4 \bullet xH_2O}\right) = 263\text{ g/mol }NiSO_4 \bullet xH_2O$

Mass of water in 1 mole of $NiSO_4 \bullet xH_2O$ = molar mass of $NiSO_4 \bullet xH_2O$ – molar mass of $NiSO_4$ (since molar ratio of $NiSO_4 \bullet xH_2O$: $NiSO_4$= 1:1)

= 263 – 154.77 = 108 g

Number of moles of water in 108 g of water = $108\text{ g }H_2O\left(\frac{1\text{ mol }H_2O}{18.02\text{ g }H_2O}\right) = 5.99 = 6$

Therefore, $x = 6$. That is, in every mole of $NiSO_4 \bullet xH_2O$, 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 $NiSO_4$ to H_2O = 1 to x

That is, $\left(\frac{\text{moles of } NiSO_4}{\text{moles of } H_2O}\right) = \left(\frac{1 \text{ mole of } NiSO_4}{x}\right)$

The ratio of $\left(\frac{\text{moles of } NiSO_4}{\text{moles of } H_2O}\right)$ in the experiment must also be the same.

Therefore,

$$\left(\frac{\text{experimental moles of } NiSO_4}{\text{experimental moles of } H_2O}\right) = \left(\frac{1 \text{ mole of } NiSO_4}{x}\right) \qquad \text{(Equation I)}$$

Moles of $NiSO_4$ 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) g = 0.214 g

Moles of H_2O produced in the experiment:

$0.214 \text{ g } H_2O \left(\frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O}\right) = 0.0118 \text{ mol } H_2O$

Substituting in equation I, and solving for x,

$\left(\frac{0.00198 \text{ mole } NiSO_4}{0.0118 \text{ mole } H_2O}\right) = \left(\frac{1 \text{ mole of } NiSO_4}{x}\right)$

$\Rightarrow x = (1 \text{ mole of } NiSO_4) \left(\frac{0.0118 \text{ mole } H_2O}{0.00198 \text{ mole of } NiSO_4}\right) = 5.95 \text{ mol } H_2O = 6 \text{ mole } H_2O$

b) $MnI_2 \bullet xH_2O_{(s)} \rightarrow MnI_{2(s)} + xH_2O_{(g)}$

$x = (1 \text{ mole of } MnI_2)\left(\frac{\text{moles of } H_2O \text{ produced}}{\text{moles of } MnI_2 \text{ produced}}\right)$

Mass of water produced = (0.895 – 0.726) g = 0.169 g

$0.169 \text{ g } H_2O \left(\frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O}\right) = 0.00937 \text{ mol } H_2O$

Moles of MnI_2 produced

$0.726 \text{ g } MnI_2 \left(\frac{1 \text{ mol } MnI_2}{308.74 \text{ g } MnI_2}\right) = 0.00235 \text{ mol } MnI_2$

$x = (1 \text{ mole of } MnI_2)\left(\frac{0.00937 \text{ moles of } H_2O}{0.00235 \text{ moles of } MnI_2}\right) = 3.98 \text{ mol } H_2O = 4 \text{ mol } H_2O$

c) similar to the ones above.

$MgSO_4 \bullet xH_2O_{(s)} \rightarrow MgSO_{4(s)} + xH_2O_{(g)}$

d) similar to the ones above.

$CdSO_4 \bullet xH_2O_{(s)} \rightarrow CdSO_{4(s)} + xH_2O_{(g)}$

e) Similar to the ones above.

$KAl(SO_4)_2 \bullet xH_2O_{(s)} \rightarrow KAl(SO_4)_{2(s)} + xH_2O_{(g)}$

27. Compound + $O_{2(g)} \rightarrow H_2O_{(g)} + CO_{2(g)}$

Mass of compound = 12.13 mg

Mass of H_2O produced = 5.36 mg

Mass of hydrogen in the sample of water produced

$$5.36 \text{ mg } H_2O \left(\frac{1\,g\,H_2O}{1000\,mg\,H_2O}\right)\left(\frac{1\,mol\,H_2O}{18.02\,g\,H_2O}\right)\left(\frac{2\,mol\,H}{1\,mol\,H_2O}\right)\left(\frac{1.01\,g\,H}{1\,mol\,H}\right) = 0.000601 \text{ g H}$$

Mass of CO_2 produced = 30.6 mg

Mass of carbon in the carbon dioxide produced

$$30.6 \text{ mg } CO_2 \left(\frac{1\,g\,CO_2}{1000\,mg\,CO_2}\right)\left(\frac{1\,mol\,CO_2}{44.01\,g\,CO_2}\right)\left(\frac{1\,mol\,C}{1\,mol\,CO_2}\right)\left(\frac{12.01\,g\,C}{1\,mol\,C}\right) = 0.00835 \text{ g C}$$

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,

mass of oxygen in the original compound = 0.01213 g – (0.000601 + 0.00835) g

= 0.01213 g – 0.00895 g

= 0.00318 g

	C	H	O
Mass ratio	0.00835	0.000601 g	0.00318 g
Molar ratio	$(0.00835\text{ g})\left(\frac{1\,mol}{12.01\,g}\right)$	$(0.000601\text{ g})\left(\frac{1\,mol}{1.01\,g}\right)$	$(0.00318\text{ g})\left(\frac{1\,mol}{16.00\,g}\right)$
	= 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 $C_7H_6O_2$.

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

% O = 100.0 – 30.4 = 69.6%

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

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

Empirical formula = NO_2

Empirical formula mass = 14.01 + 2(16.00) = 46.01 g/mol

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

Therefore, the molecular formula is N_2O_4.

29. Assuming 100.00 g of compound

Mass of hydrogen = 100.00 – 49.31 – 43.79 = 6.90 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 givening an empirical formula $C_3H_5O_2$.

Empirical formula mass = 3(12.01) + 5(1.01) + 2(16.00) = 73.08 g/mol

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

Therefore, the molecular formula is $C_6H_{10}O_4$.

30. Assuming 100.00 g of compound

Mass of oxygen = 100.00 – 41.39 – 3.47 = 55.14 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.

Empirical formula mass = 12.01 + 1.01 + 16.00 = 29.02 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 $C_4H_4O_4$.