CHAPTER 3
CHEMICAL COMPOUNDS

PRACTICE EXAMPLES

1A **(E)** First we convert the number of chloride ions to the mass of $MgCl_2$.

$$\text{mass}_{MgCl_2} = 5.0\times10^{23}\,Cl^- \times \frac{1\ MgCl_2}{2\ Cl^-}\times\frac{1\text{ mol }MgCl_2}{6.022\times10^{23}\,MgCl_2}\times\frac{95.205\text{ g }MgCl_2}{1\text{ mol }MgCl_2} = 4.0\times10^{1}\text{ g }MgCl_2$$

1B **(M)** First we convert mass $Mg(NO_3)_2$ to moles $Mg(NO_3)_2$ and formula units $Mg(NO_3)_2$ then finally to NO_3^- ions.

$$1.00\ \mu\text{g } Mg(NO_3)_2\times\frac{1\text{ g } Mg(NO_3)_2}{1{,}000{,}000\ \mu\text{g } Mg(NO_3)_2}\times\frac{1\text{ mol } Mg(NO_3)_2}{148.313\text{ g } Mg(NO_3)_2}\times\frac{6.022\times10^{22}\text{ formula units } Mg(NO_3)_2}{1\text{ mol } Mg(NO_3)_2}$$

$$= 4.06\times10^{15}\text{ formula units } Mg(NO_3)_2\times\frac{2\ NO_3^-\text{ ions}}{1\text{ formula unit } Mg(NO_3)_2} = 8.12\times10^{15}\ NO_3^-\text{ ions}$$

Next, determine the number of oxygen atoms by multiplying by the appropriate ratio.

$$\#\text{ O atoms} = 4.06\times10^{15}\text{ formula units } Mg(NO_3)_2\times\frac{6\text{ atoms O}}{1\text{ formula unit } Mg(NO_3)_2} = 2.44\times10^{16}\text{ O}$$

2A **(M)** The volume of gold is converted to its mass and then to the amount in moles.

$$\#\text{Au atoms} = (2.50\text{ cm})^2\times\left(0.100\text{ mm}\times\frac{1\text{ cm}}{10\text{ mm}}\right)\times\frac{19.32\text{ g}}{1\text{ cm}^3}\times\frac{1\text{ mol Au}}{196.97\text{ g Au}}\times\frac{6.022\times10^{23}\text{ atoms}}{1\text{ mol Au}}$$

$$= 3.69\times10^{21}\text{ Au atoms}$$

2B **(M)** We need the molar mass of ethyl mercaptan for one conversion factor.

$$M = (2\times12.011\text{ g C}) + (6\times1.008\text{ g H}) + (1\times32.06\text{ g S}) = 62.13\text{ g/mol } C_2H_6S$$

Volume of room: $62\text{ ft}\times35\text{ ft}\times14\text{ ft} = 3.0\underline{4}\times10^4\text{ ft}^3$. We also need to convert ft^3 to m^3.

$$3.0\underline{4}\times10^4\text{ ft}^3\times\left(\frac{12\text{ in}}{1\text{ ft}}\right)^3\times\left(\frac{2.54\text{ cm}}{1\text{ in}}\right)^3\times\left(\frac{1\text{ m}}{100\text{ cm}}\right)^3 = 8.6\times10^2\text{ m}^3$$

$$[C_2H_6S] = \frac{1.0\ \mu\text{L } C_2H_6S}{8.6\times10^2\text{ m}^3}\times\frac{1\text{ L}}{1\times10^6\ \mu\text{L}}\times\frac{1000\text{ mL}}{1\text{ L}}\times\frac{0.84\text{ g}}{1\text{ mL}}\times\frac{1\text{ mol } C_2H_6S}{62.13\text{ g}}\times\frac{10^6\ \mu\text{mol}}{1\text{ mol}}$$

$$= 0.016\ \mu\text{mol/m}^3 > 9.0\times10^{-4}\ \mu\text{mol/m}^3 = \text{the detectable limit}$$

Thus, the vapor will be detectable.

3A **(M)** The molar mass of halothane is given in Example 3-3 *in the textbook* as 197.4 g/mol. The rest of the solution uses conversion factors to change units.

$$\text{mass Br} = 25.0\ \text{mL}\ C_2HBrClF_3 \times \frac{1.871\ \text{g}\ C_2HBrClF_3}{1\ \text{mL}\ C_2HBrClF_3} \times \frac{1\ \text{mol}\ C_2HBrClF_3}{197.4\ \text{g}\ C_2HBrClF_3} \times \frac{1\ \text{mol Br}}{1\ \text{mol}\ C_2HBrClF_3}$$
$$\times \frac{79.904\ \text{g Br}}{1\ \text{mol Br}} = 18.9\ \text{g Br}$$

3B **(M)** Again, the molar mass of halothane is given in Example 3-3 *in the textbook* as 197.4 g/mol.

$$V_{\text{halothane}} = 1.00\times10^{24}\ \text{Br} \times \frac{1\ \text{mol Br}}{6.022\times10^{23}\ \text{Br}} \times \frac{1\ \text{mol}\ C_2HBrClF_3}{1\ \text{mol Br}} \times \frac{197.4\ \text{g}\ C_2HBrClF_3}{1\ \text{mol}\ C_2HBrClF_3} \times \frac{1\ \text{mL}}{1.871\ \text{g}}$$
$$= 175\ \text{mL}\ C_2HBrClF_3$$

4A **(M)** We use the same technique as before: determine the mass of each element in a mole of the compound. Their sum is the molar mass of the compound. The percent composition is determined by comparing the mass of each element with the molar mass of the compound.

$$M = (10\times12.011\ \text{g C}) + (16\times1.008\ \text{g H}) + (5\times14.01\ \text{g N}) + (3\times30.97\ \text{g P}) + (13\times15.999\ \text{g O})$$
$$= 120.11\ \text{g C} + 16.13\ \text{g H} + 70.05\ \text{g N} + 92.91\ \text{g P} + 207.99\ \text{g O} = 507.19\ \text{g ATP/mol}$$

$$\%\text{C} = \frac{120.11\ \text{g C}}{507.19\ \text{g ATP}}\times100\% = 23.681\%\ \text{C} \quad \%\text{H} = \frac{16.13\ \text{g H}}{507.19\ \text{g ATP}}\times100\% = 3.180\%\ \text{H}$$

$$\%\text{N} = \frac{70.05\ \text{g N}}{507.19\ \text{g ATP}}\times100\% = 13.81\%\ \text{N} \quad \%\text{P} = \frac{92.91\ \text{g P}}{507.19\ \text{g ATP}}\times100\% = 18.32\%\ \text{P}$$

$$\%\text{O} = \frac{207.99\ \text{g O}}{507.19\ \text{g ATP}}\times100\% = 41.008\%\ \text{O}\ \ \text{(NOTE: the mass percents sum to 99.999\%)}$$

4B **(E)** Both (b) and (e) have the same empirical formula, that is, CH_2O. These two molecules have the same percent oxygen by mass.

5A **(M)** Once again, we begin with a 100.00 g sample of the compound. In this way, each elemental mass in grams is numerically equal to its percent. We convert each mass to an amount in moles, and then determine the simplest integer set of molar amounts. This determination begins by dividing all three molar amounts by the smallest.

$$39.56\ \text{g C} \times \frac{1\ \text{mol C}}{12.011\ \text{g C}} = 3.294\ \text{mol C} \div 3.294 \rightarrow 1.000\ \text{mol C} \times 3.000 = 3.000\ \text{mol C}$$

$$7.74\ \text{g H} \times \frac{1\ \text{mol H}}{1.008\ \text{g H}} = 7.68\ \text{mol H} \div 3.294 \rightarrow 2.33\ \text{mol H} \times 3.000 = 6.99\ \text{mol H}$$

$$52.70\ \text{g O} \times \frac{1\ \text{mol O}}{15.999\ \text{g O}} = 3.294\ \text{mol O} \div 3.294 \rightarrow 1.000\ \text{mol O} \times 3.000 = 3.000\ \text{mol O}$$

Thus, the empirical formula of the compound is $C_3H_7O_3$. The empirical molar mass of this compound is:

$$(3\times 12.011\,\text{g C})+(7\times 1.008\,\text{g H})+(3\times 15.999\,\text{g O}) = 36.033\text{ g}+7.056\text{ g}+47.997\text{ g} = 91.09\,\text{g/mol}$$

The empirical mass is almost precisely one half the reported molar mass, leading to the conclusion that the molecular formula must be twice the empirical formula in order to double the molar mass. Thus, the molecular formula is $C_6H_{14}O_6$.

5B **(M)** To answer this question, we start with a 100.00 g sample of the compound. In this way, each elemental mass in grams is numerically equal to its percent. We convert each mass to an amount in moles, and then determine the simplest integer set of molar amounts. This determination begins by dividing all molar amounts by the smallest number of moles in the group of four, which is 1.1025 moles. Multiplication of the resulting quotients by eight produces the smallest possible set of whole numbers.

$$21.51\,\text{g C}\times\frac{1\,\text{mol C}}{12.011\,\text{g C}} = 1.791\text{ mol C}\div 1.102\underline{5}\rightarrow 1.624\text{ mol C}\times 8 = 12.99\,\text{mol C}$$

$$2.22\text{ g H}\times\frac{1\,\text{mol H}}{1.008\,\text{g H}} = 2.20\text{ mol H}\div 1.102\underline{5}\rightarrow 2.00\text{ mol H}\times 8 = 16.0\text{ mol H}$$

$$17.64\text{ g O}\times\frac{1\,\text{mol O}}{15.999\,\text{g O}} = 1.102\underline{5}\,\text{mol O}\div 1.102\underline{5}\rightarrow 1.000\,\text{mol O}\times 8 = 8.000\,\text{mol O}$$

$$58.63\,\text{g Cl}\times\frac{1\,\text{mol Cl}}{35.45\,\text{g Cl}} = 1.654\text{ mol Cl}\div 1.102\underline{5}\rightarrow 1.500\text{ mol C}\times 8 = 12.00\,\text{mol Cl}$$

Thus, the empirical formula of the compound is $C_{13}H_{16}O_8Cl_{12}$. The empirical molar mass of this compound is 725.7 g/mol.

The empirical mass is almost precisely the same as the reported molar mass, leading to the conclusion that the molecular formula must be the same as the empirical formula. Thus, the molecular formula is $C_{13}H_{16}O_8Cl_{12}$.

6A **(M)** We calculate the amount in moles of each element in the sample (determining the mass of oxygen by difference) and transform these molar amounts to the simplest integral amounts, by first dividing all three by the smallest.

$$2.726\,\text{g CO}_2\times\frac{1\,\text{mol CO}_2}{44.009\,\text{g CO}_2}\times\frac{1\,\text{mol C}}{1\,\text{mol CO}_2} = 0.06194\,\text{mol C}\times\frac{12.011\,\text{g C}}{1\,\text{mol C}} = 0.7440\,\text{g C}$$

$$1.116\,\text{g H}_2\text{O}\times\frac{1\,\text{mol H}_2\text{O}}{18.015\,\text{g H}_2\text{O}}\times\frac{2\,\text{mol H}}{1\,\text{mol H}_2\text{O}} = 0.1239\,\text{mol H}\times\frac{1.008\,\text{g H}}{1\,\text{mol H}} = 0.1249\,\text{g H}$$

$$(1.152\,\text{g cmpd} - 0.7440\,\text{g C} - 0.1249\,\text{g H}) = 0.283\,\text{g O}\times\frac{1\,\text{mol O}}{15.999\,\text{g O}} = 0.0177\,\text{mol O}$$

$$\left.\begin{array}{lll}0.06194\,\text{mol C} \div 0.0177 & \rightarrow 3.50 \\ 0.1239\,\text{mol H} \div 0.0177 & \rightarrow 7.00 \\ 0.0177\,\text{mol O} \div 0.0177 & \rightarrow 1.00\end{array}\right\}$$

All of these amounts in moles are multiplied by 2 to make them integral. Thus, the empirical formula of isobutyl propionate is $C_7H_{14}O_2$.

6B **(M)** Notice that we do not have to obtain the mass of any element in this compound by difference; there is no oxygen present in the compound. We calculate the amount in one mole of each element in the sample and transform these molar amounts to the simplest integral amounts, by first dividing all three by the smallest.

$$3.149\,\text{g CO}_2 \times \frac{1\,\text{mol CO}_2}{44.009\,\text{g CO}_2} \times \frac{1\,\text{mol C}}{1\,\text{mol CO}_2} = 0.07155\,\text{mol C} \div 0.01789 = 3.999\,\text{mol C}$$

$$0.645\,\text{g H}_2\text{O} \times \frac{1\,\text{mol H}_2\text{O}}{18.015\,\text{g H}_2\text{O}} \times \frac{2\,\text{mol H}}{1\,\text{mol H}_2\text{O}} = 0.0716\,\text{mol H} \div 0.01789 = 4.00\,\text{mol H}$$

$$1.146\,\text{g SO}_2 \times \frac{1\,\text{mol SO}_2}{64.058\,\text{g SO}_2} \times \frac{1\,\text{mol S}}{1\,\text{mol SO}_2} = 0.01789\,\text{mol S} \div 0.01789 = 1.000\,\text{mol S}$$

Thus, the empirical formula of thiophene is C_4H_4S.

7A **(E)**

$\underline{S}_8$ For an atom of a free element, the oxidation state is 0 (rule 1).

$\underline{Cr}_2O_7^{2-}$ The sum of all the oxidation numbers in the ion is -2 (rule 2). The O.S. of each oxygen is -2 (rule 6). Thus, the total for all seven oxygens is -14. The total for both chromiums must be $+12$. Thus, each Cr has an $\text{O.S.} = +6$.

$\underline{Cl}_2O$ The sum of all oxidation numbers in the compound is 0 (rule 2). The O.S. of oxygen is -2 (rule 6). The total for the two chlorines must be $+2$. Thus, each chlorine must have $\text{O.S.} = +1$.

$K\underline{O}_2$ The sum for all the oxidation numbers in the compound is 0 (rule 2). The O.S. of potassium is $+1$ (rule 3). The sum of the oxidation numbers of the two oxygens must be -1. Thus, each oxygen must have $\text{O.S.} = -1/2$.

7B **(E)**

$\underline{S}_2O_3^{2-}$ The sum of all the oxidation numbers in the ion is -2 (rule 2). The O.S. of oxygen is -2 (rule 6). Thus, the total for three oxygens must be -6. The total for both sulfurs must be $+4$. Thus, each S has an $\text{O.S.} = +2$.

$\underline{Hg_2}Cl_2$ The O.S. of each Cl is -1 (rule 7). The sum of all O.S. is 0 (rule 2). Thus, the total for two Hg is $+2$ and each Hg has $\text{O.S.} = +1$.

$K\underline{Mn}O_4$ The O.S. of each O is -2 (rule 6). Thus, the total for 4 oxygens must be -8. The K has $\text{O.S.} = +1$ (rule 3). The total of all O.S. is 0 (rule 2). Thus, the O.S. of Mn is +7.

$H_2\underline{C}O$ The O.S. of each H is $+1$ (rule 5), producing a total for both hydrogens of $+2$. The O.S. of O is -2 (rule 6). Thus, the O.S. of C is 0, because the total of all O.S. values is 0 (rule 2).

8A **(E)** In each case, we determine the formula *with its accompanying charge* of each ion in the compound. We then produce a formula for the compound in which the total positive charge equals the total negative charge.

lithium oxide	Li^+ and O^{2-}	*two* Li^+ and *one* O^{2-}	Li_2O
tin(II) fluoride	Sn^{2+} and F^-	*one* Sn^{2+} and *two* F^-	SnF_2
lithium nitride	Li^+ and N^{3-}	*three* Li^+ and *one* N^{3-}	Li_3N

8B **(E)** Using a similar procedure as that provided in **8A**

aluminum sulfide	Al^{3+} and S^{2-}	*two* Al^{3+} and *three* S^{2-}	Al_2S_3
magnesium nitride	Mg^{2+} and N^{3-}	*three* Mg^{2+} and *two* N^{3-}	Mg_3N_2
vanadium(III) oxide	V^{3+} and O^{2-}	*two* V^{3+} and *three* O^{2-}	V_2O_3

9A **(E)** The name of each of these ionic compounds is the name of the cation followed by that of the anion. Each anion name is a modified (with the ending -*ide*) version of the name of the element. Each cation name is the name of the metal, with the oxidation state appended in Roman numerals in parentheses if there is more than one type of cation for that metal.

CsI cesium iodide

CaF_2 calcium fluoride

FeO The O.S. of $O = -2$ (rule 6). Thus, the O.S. of $Fe = +2$ (rule 2).
The cation is iron(II). The name of the compound is iron(II) oxide.

$CrCl_3$ The O.S. of $Cl = -1$ (rule 7). Thus, the O.S. of $Cr = +3$ (rule 2).
The cation is chromium(III). The compound is chromium(III) chloride.

9B **(E)** The name of each of these ionic compounds is the name of the cation followed by that of the anion. Each anion name is a modified (with the ending -*ide*) version of the name of the element. Each cation name is the name of the metal, with the oxidation state appended in Roman numerals in parentheses if there is more than one type of cation for that metal.

The oxidation state of Ca is +2 (rule 3). Hydrogen would therefore have an oxidation number of –1 (which is an exception to rule 5), based on rule 2.
CaH_2 calcium hydride

The oxidation number of sulfur is –2 (rule 7), and therefore silver would be +1 for each silver atom based on rule 2.
Ag_2S silver(I) sulfide

In the next two compounds, the oxidation state of chlorine is -1 (rule 7) and thus the oxidation state of the metal in each cation must be $+1$ (rule 2).
$CuCl$ copper(I) chloride Hg_2Cl_2 mercury(I) chloride

10A (E)

SF_6	Both S and F are nonmetals. This is a binary molecular compound: sulfur hexafluoride.
HNO_2	The NO_2^- ion is the nitrite ion. Its acid is nitrous acid.
$Ca(HCO_3)_2$	HCO_3^- is the bicarbonate ion or the hydrogen carbonate ion. This compound is calcium bicarbonate or calcium hydrogen carbonate.
$FeSO_4$	The SO_4^{2-} ion is the sulfate ion. The cation is Fe^{2+}, iron(II). This compound is iron(II) sulfate.

10B (E)

NH_4NO_3	The cation is NH_4^+, ammonium ion. The anion is NO_3^-, nitrate ion. This compound is ammonium nitrate.
PCl_3	Both P and Cl are nonmetals. This is a binary molecular compound: phosphorus trichloride.
HBrO	BrO^- is hypobromite, this is hypobromous acid.
$AgClO_4$	The anion is perchlorate ion, ClO_4^-. The compound is silver(I) perchlorate.
$Fe_2(SO_4)_3$	The SO_4^{2-} ion is the sulfate ion. The cation is Fe^{3+}, iron(III). This compound is iron(III) sulfate.

11A (E)

boron trifluoride	Both elements are nonmetals. This is a binary molecular compound: BF_3.
potassium dichromate	Potassium ion is K^+, and dichromate ion is $Cr_2O_7^{2-}$. This is $K_2Cr_2O_7$.
sulfuric acid	The anion is sulfate, SO_4^{2-}. There must be two H^+ s. This is H_2SO_4.
calcium chloride	The ions are Ca^{2+} and Cl^-. There must be one Ca^{2+} and two Cl^- s: $CaCl_2$.

11B (E)

aluminum nitrate	Aluminum is Al^{3+};; the nitrate ion is NO_3^-. This is $Al(NO_3)_3$.
tetraphosphorus decoxide	Both elements are nonmetals. This is a binary molecular compound, P_4O_{10}.
chromium(III) hydroxide	Chromium(III) ion is Cr^{3+}; the hydroxide ion is OH^-. This is $Cr(OH)_3$.
iodic acid	The halogen “ic” acid has the halogen in a $+5$ oxidation state. This is HIO_3.

12A (E)

(a) Not isomers: molecular formulas are different (C_8H_{18} vs C_9H_{20}).

(b) Molecules are isomers (same formula C_7H_{16}).

12B (E)

(a) Molecules are isomers (same formula C_7H_{14}).

(b) Not isomers: molecular formulas are different (C_4H_8 vs C_5H_{10}).

13A (E)

(a) The carbon to carbon bonds are all single bonds in this hydrocarbon. This compound is an alkane.

(b) In this compound, there are only single bonds, and a Cl atom has replaced one H atom. This compound is a chloroalkane.

(c) The presence of the carboxyl group (—CO_2H) in this molecule means that the compound is a carboxylic acid.

(d) There is a carbon to carbon double bond in this hydrocarbon. This is an alkene.

13B (E)

(a) The presence of the hydroxyl group (—OH) in this molecule means that this compound is an alcohol.

(b) The presence of the carboxyl group (—CO_2H) in this molecule means that the compound is a carboxylic acid. This molecule also contains the hydroxyl group (—OH).

(c) The presence of the carboxyl group (—CO_2H) in this molecule means that the compound is a carboxylic acid. As well, a Cl atom has replaced one H atom. This compound is a chloroalkane. The compound is a chloro carboxylic acid.

(d) There is a carbon to carbon double bond in this compound; hence, it is an alkene. There is also one H atom that has been replaced by a Br atom. This compound is also a bromoalkene.

14A (E)

(a) The structure is that of an alcohol with the hydroxyl group on the second carbon atom of a three carbon chain. The compound is propan-2-ol (commonly isopropyl alcohol).

(b) The structure is that of an iodoalkane molecule with the I atom on the first carbon of a three-carbon chain. The compound is called 1-iodopropane.

(c) The carbon chain in this structure is four carbon atoms long with the end C atom in a carboxyl group. There is also a methyl group on the third carbon in the chain. The compound is 3-methylbutanoic acid.

(d) The structure is that of a three carbon chain that contains a carbon to carbon double bond. This compound is propene.

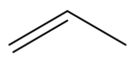
14B (E)

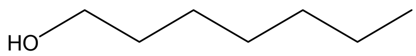
(a) 2-chloropropane **(b)** 1,4-dichlorobutane **(c)** 2-methylpropanoic acid

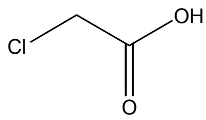
15A **(E)**

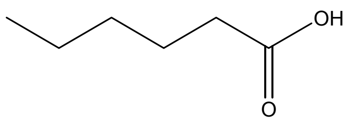
(a) pentane: $CH_3(CH_2)_3CH_3$
(b) ethanoic acid: CH_3CO_2H
(c) 1-iodooctane: $ICH_2(CH_2)_6CH_3$
(d) pentan-1-ol: $CH_2(OH)(CH_2)_3CH_3$

15B **(E)**

(a) propene 

(b) heptan-1-ol 

(c) chloroacetic acid 

(d) hexanoic acid 

INTEGRATIVE EXAMPLES

A. **(M)**
First, determine the mole ratios of the dehydrated compound:
27.74 g Mg × (1 mol Mg / 24.305 g Mg) = 1.141 mol Mg
23.57 g P × (1 mol P / 30.974 g P) = 0.7610 mol P
48.69 g O × (1 mol O / 15.999 g O) = 3.043 mol O

Mole ratios are determined by dividing by the smallest number:
1.141 mol Mg / 0.7610 mol P ≈ 1.5
0.7610 mol P / 0.7610 mol P ≈ 1.0
3.043 mol O / 0.7610 mol P ≈ 4.0

Multiplying by 2 to get whole numbers, the empirical formula becomes $Mg_3P_2O_8$. The compound is magnesium phosphate, $Mg_3(PO_4)_2$.

To determine the number of waters of hydration, determine the mass of water driven off.

mass of H_2O = 2.4917 g – 1.8558 g = 0.6359 g
mol H_2O = 0.6359 g × (1 mol H_2O/18.015 g H_2O) = 0.0353 mol.

Then, calculate the number of moles of dehydrated $Mg_3(PO_4)_2$ in the same manner above. The number of moles (using 262.86 g/mol for molecular weight) is 0.00706. Dividing the number of moles of H_2O by $Mg_3(PO_4)_2$ gives a ratio of 5. Therefore, the compound is $Mg_3(PO_4)_2 \cdot 5\ H_2O$

B. **(M)**
First, determine the mole ratio of the elements in this compound:

17.15 g Cu × (1 mol Cu / 63.546 g Cu) = 0.2699 mol Cu
19.14 g Cl × (1 mol Cl / 35.45 g Cl) = 0.5399 mol Cl
60.45 g O × (1 mol O / 15.999 g O) = 3.778 mol O

Mass of H: 100 – (17.15 + 19.14 + 60.45) = 3.26 g H
3.26 g H × (1 mol H / 1.008 g H) = 3.23 mol H
Mole ratios are determined by dividing by the smallest number:

0.2699 mol Cu / 0.2699 mol Cu = 1.000
0.5399 mol Cl / 0.2699 mol Cu = 2.000 (mol Cl per mol Cu)
3.778 mol O / 0.2699 mol Cu = 14.00 (mol O per mol Cu)
3.23 mol H / 0.2699 mol Cu = 12.0 (mol H per mol Cu)

Now we know that since all the hydrogen atoms are taken up as water, half as many moles of O are also taken up as water. Therefore, if there are 12 moles of H, 6 moles of O are needed, 6 moles of H_2O are generated, and 8 moles of O are left behind.

To determine the oxidation state of Cu and Cl, we note that there are 4 times as many moles of O as there is Cl. If the Cl and O are associated, we have the perchlorate ion (ClO_4^-) and the formula of the compound is $Cu(ClO_4)_2 \cdot 6\ H_2O$. The oxidation state of Cu is +2 and Cl is +7.

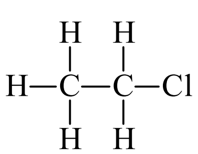
EXERCISES

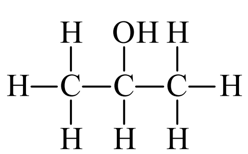
Representing Molecules

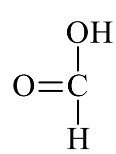
1. **(E)**
(a) H_2O_2 **(b)** CH_3CH_2Cl **(c)** P_4O_{10}
(d) $CH_3CH(OH)CH_3$ **(e)** HCO_2H

2. **(E)**
(a) N_2H_4 **(b)** CH_3CH_2OH **(c)** P_4O_6
(d) $CH_3(CH_2)_3CO_2H$ **(e)** $CH_2(OH)CH_2Cl$

3. **(E)**
(b) CH_3CH_2Cl **(d)** $CH_3CH(OH)CH_3$ **(e)** HCO_2H

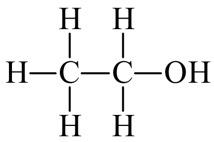




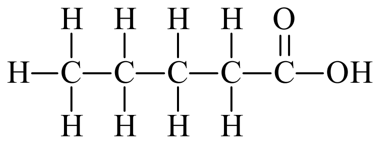


4. **(E)**

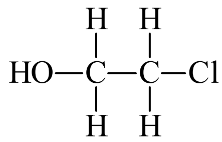
(b) CH_3CH_2OH



(d) $CH_3(CH_2)_3CO_2H$



(e) $CH_2(OH)CH_2Cl$



The Avogadro Constant and the Mole

<u>5.</u> **(M)**

(a) A trinitrotoluene molecule, $CH_3C_6H_2(NO_2)_3$, contains 7 C atoms, 5 H atoms, 3 N atoms, and 3×2 O atoms = 6 O atoms, for a total of 7 + 5 + 3 + 6 = 21 atoms.

(b) $CH_3(CH_2)_4CH_2OH$ contains 6 C atoms, 14 H atoms, and 1 O atom, for a total of 21 atoms.

Conversion pathway approach:

$$\#\text{ of atoms} = 0.00102\,\text{mol CH}_3(\text{CH}_2)_4\text{CH}_2\text{OH}\times\frac{6.022\times10^{23}\,\text{molecules CH}_3(\text{CH}_2)_4\text{CH}_2\text{OH}}{1\,\text{mol CH}_3(\text{CH}_2)_4\text{CH}_2\text{OH}}$$

$$\times\frac{21\text{ atoms}}{1\,\text{molecule CH}_3(\text{CH}_2)_4\text{CH}_2\text{OH}} = 1.29\times10^{22}\text{ atoms}$$

Stepwise approach:

$$0.00102\,\text{mol CH}_3(\text{CH}_2)_4\text{CH}_2\text{OH}\times\frac{6.022\times10^{23}\,\text{molecules CH}_3(\text{CH}_2)_4\text{CH}_2\text{OH}}{1\,\text{mol CH}_3(\text{CH}_2)_4\text{CH}_2\text{OH}} = 6.14\times10^{20}\text{molecules}$$

$$6.14\times10^{20}\text{molecules}\times\frac{21\text{ atoms}}{1\text{ CH}_3(\text{CH}_2)_4\text{CH}_2\text{OH molecule}} = 1.29\times10^{22}\text{ atoms}$$

(c) *Conversion pathway approach:*

$$\#\text{ of F atoms} = 12.15\,\text{mol C}_2\text{HBrClF}_3\times\frac{3\,\text{mol F}}{1\,\text{mol C}_2\text{HBrClF}_3}\times\frac{6.022\times10^{23}\,\text{F atoms}}{1\,\text{mol F atoms}}$$

$$= 2.195\times10^{25}\text{ F atoms}$$

Stepwise approach:

$$12.15\,\text{mol C}_2\text{HBrClF}_3\times\frac{3\,\text{mol F}}{1\,\text{mol C}_2\text{HBrClF}_3} = 36.45\text{ mol F}$$

$$36.45\text{ mol F}\times\frac{6.022\times10^{23}\,\text{F atoms}}{1\,\text{mol F atoms}} = 2.195\times10^{25}\text{ F atoms}$$

6. (E)

(a) To convert the amount in moles to mass, we need the molar mass of NO_2 (46.005 g mol^{-1}).

$$\text{mass } NO_2 = 14.68 \text{ mol } NO_2 \times \frac{46.005 \text{ g } NO_2}{1 \text{ mol } NO_2} = 675 \text{ g } NO_2$$

(b) $$\text{mass of } O_2 = 4.220 \times 10^{25} O_2 \text{ molecules} \times \frac{1 \text{ mol } O_2}{6.022 \times 10^{23} \text{ molecules } O_2} \times \frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2}$$

$$= 2242 \text{ g } O_2$$

(c) $$\text{mass of } CuSO_4 \cdot 5\,H_2O = 15.5 \text{ mol} \times \frac{249.7 \text{ g } CuSO_4 \cdot 5\,H_2O}{1 \text{ mol } CuSO_4 \cdot 5\,H_2O} = 3.87 \times 10^3 \text{ g } CuSO_4 \cdot 5\,H_2O$$

(d) $$\text{mass of } C_2H_4(OH)_2 = 2.25 \times 10^{24} \text{ molecules of } C_2H_4(OH)_2 \times \frac{1 \text{ mol } C_2H_4(OH)_2}{6.022 \times 10^{23} \text{ molecules of } C_2H_4(OH)_2}$$

$$\times \frac{62.07 \text{ g } C_2H_4(OH)_2}{1 \text{ mol } C_2H_4(OH)_2} = 232 \text{ g } C_2H_4(OH)_2$$

7. (M)

(a) molecular mass (mass of one molecule) of $C_5H_{11}NO_2S$ is:

$$(5 \times 12.011 \text{ u C}) + (11 \times 1.008 \text{ u H}) + 14.007 \text{ u N} + (2 \times 15.999 \text{ u O}) + 32.06 \text{ u S}$$

$$= 149.208 \text{ u}/C_5H_{11}NO_2S \text{ molecule}$$

(b) Since there are 11 H atoms in each $C_5H_{11}NO_2S$ molecule, there are 11 moles of H atoms in each mole of $C_5H_{11}NO_2S$ molecules.

(c) $$\text{mass C} = 1 \text{ mol } C_5H_{11}NO_2S \times \frac{5 \text{ mol C}}{1 \text{ mol } C_5H_{11}NO_2S} \times \frac{12.011 \text{ g C}}{1 \text{ mol C}} = 60.055 \text{ g C}$$

(d) $$\#\text{ C atoms} = 9.07 \text{ mol } C_5H_{11}NO_2S \times \frac{5 \text{ mol C}}{1 \text{ mol } C_5H_{11}NO_2S} \times \frac{6.022 \times 10^{23} \text{ atoms C}}{1 \text{ mol C}}$$

$$= 2.73 \times 10^{25} \text{ C atoms}$$

8. (M)

(a) $$\text{amount of } Br_2 = 4.04 \times 10^{22}\, Br_2 \text{ molecules} \times \frac{1 \text{ mole } Br_2}{6.022 \times 10^{23}\, Br_2 \text{ molecules}}$$

$$= 0.0671 \text{ mol } Br_2$$

(b) $$\text{amount of } Br_2 = 5.78 \times 10^{24} \text{ Br atoms} \times \frac{1\, Br_2 \text{ molecule}}{2 \text{ Br atoms}} \times \frac{1 \text{ mole } Br_2}{6.022 \times 10^{23}\, Br_2 \text{ molecules}}$$

$$= 4.80 \text{ mol } Br_2$$

(c) $$\text{amount of } Br_2 = 7.82 \text{ kg } Br_2 \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol } Br_2}{159.8 \text{ g } Br_2} = 48.9 \text{ mol } Br_2$$

(d) $$\text{amount of } Br_2 = 3.56 \text{ L } Br_2 \times \frac{1000 \text{ mL}}{1 \text{ L}} \times \frac{3.10 \text{ g } Br_2}{1 \text{ mL } Br_2} \times \frac{1 \text{ mol } Br_2}{159.8 \text{ g } Br_2} = 70.8 \text{ mol } Br_2$$

9. **(E)** The greatest number of N atoms is found in the compound with the greatest number of moles of N.

The molar mass of N_2O = $(2\,\text{mol N} \times 14.0\,\text{g N}) + (1\,\text{mol O} \times 16.0\,\text{g O}) = 44.0\,\text{g/mol N}_2\text{O}$. Thus, 50.0 g N_2O is slightly more than 1 mole of N_2O, and contains slightly more than 2 moles of N. Each mole of N_2 contains 2 moles of N. The molar mass of NH_3 is 17.0 g. Thus, there is 1 mole of NH_3 present, which contains 1 mole of N.

The molar mass of pyridine is $(5\,\text{mol C} \times 12.0\,\text{g C}) + (5\,\text{mol H} \times 1.01\,\text{g H}) + 14.0\,\text{g N} = 79.1\,\text{g/mol}$. Because each mole of pyridine contains 1 mole of N, we need slightly more than 2 moles of pyridine to have more N than is present in the N_2O. But that would be a mass of about 158 g pyridine, and 150 mL has a mass of less than 150 g. Thus, the greatest number of N atoms is present in 50.0 g N_2O.

10. **(E)** The greatest number of S atoms is contained in the compound with the greatest number of moles of S. The solid sulfur contains $8 \times 0.12\,\text{mol} = 0.96$ mol S atoms. There are 0.50×2 mol S atoms in 0.50 mol S_2O. There is slightly greater than 1 mole (64.1 g) of SO_2 in 65 g, and thus a bit more than 1 mole of S atoms. The molar mass of thiophene is 84.1 g and thus contains less than 1 mole of S. This means that 65 g SO_2 has the greatest number of S atoms.

11. **(M)**

(a) $$\text{moles N}_2\text{O}_4 = 115\,\text{g N}_2\text{O}_4 \times \frac{1\,\text{mol N}_2\text{O}_4}{92.01\,\text{g N}_2\text{O}_4} = 1.25\,\text{mol N}_2\text{O}_4$$

(b) $$\text{moles N} = 43.5\text{ g Mg}(\text{NO}_3)_2 \times \frac{1\,\text{mol Mg}(\text{NO}_3)_2}{148.31\,\text{g}} \times \frac{2\,\text{mol N}}{1\,\text{mol Mg}(\text{NO}_3)_2} = 0.587\,\text{mol N atoms}$$

(c) $$\text{moles N} = 12.4\text{ g C}_6\text{H}_{12}\text{O}_6 \times \frac{1\,\text{mol C}_6\text{H}_{12}\text{O}_6}{180.15\,\text{g}} \times \frac{6\,\text{mol O}}{1\,\text{mol C}_6\text{H}_{12}\text{O}_6} \times \frac{1\,\text{mol C}_7\text{H}_5(\text{NO}_2)_3}{6\,\text{mol O}}$$
$$\times \frac{3\,\text{mol N}}{1\,\text{mol C}_7\text{H}_5(\text{NO}_2)_3} = 0.206\,\text{mol N}$$

12. **(M)**

(a) $$\text{mass} = 2.10 \times 10^{-2}\text{ mol S}_8 \times \frac{8\text{ mol S}}{1\text{ mol S}_8} \times \frac{32.06\text{ g S}}{1\text{ mol S}} = 5.39\text{ g S}$$

(b) $$\text{mass} = 5.02 \times 10^{22}\text{ molecules} \times \frac{1\text{ mole}}{6.022 \times 10^{23}\text{ molecules}} \times \frac{256.43\text{ g C}_{16}\text{H}_{32}\text{O}_2}{1\text{ mole C}_{16}\text{H}_{32}\text{O}_2}$$
$$= 21.4\text{ g C}_{16}\text{H}_{32}\text{O}_2\text{ (palmitic acid).}$$

(c) $$\text{mass} = 2.95\text{ mol N} \times \frac{1\text{ mol C}_6\text{H}_9\text{N}_3\text{O}_2}{3\text{ mol N}} \times \frac{155.157\text{ g C}_6\text{H}_9\text{N}_3\text{O}_2}{1\text{ mol C}_6\text{H}_9\text{N}_3\text{O}_2} = 152.6\text{ g C}_6\text{H}_9\text{N}_3\text{O}_2$$

13. **(M)** The number of Fe atoms in 6 L of blood can be found using the following conversion pathway.

$$= 6\,\text{L blood}\times\frac{1000\,\text{mL}}{1\,\text{L}}\times\frac{15.5\,\text{g Hb}}{100\,\text{mL blood}}\times\frac{1\,\text{mol Hb}}{64,500\,\text{g Hb}}\times\frac{4\,\text{mol Fe}}{1\,\text{mol Hb}}\times\frac{6.022\times10^{23}\,\text{atoms Fe}}{1\,\text{mol Fe}}$$

$= 3\times10^{22}$ Fe atoms

14. **(M)**

(a) volume $= \pi r^2 \times h = 3.1416\times\left(\frac{1.22\text{ cm}}{2}\right)^2\times 6.50\text{ cm} = 7.60\text{ cm}^3$

$$\text{mol } P_4 = 7.60\text{ cm}^3\times\frac{1.823\text{ g } P_4}{1\text{ cm}^3}\times\frac{1\text{ mol } P_4}{123.896\text{ g } P_4} = 0.112\text{ mol } P_4$$

(b) $\#\text{ P atoms} = 0.112\text{ mol } P_4\times\frac{4\text{ mol P}}{1\text{ mol } P_4}\times\frac{6.022\times10^{23}\text{ P}}{1\text{ mol P}} = 2.70\times10^{23}\text{ P atoms}$

Chemical Formulas

15. **(E)** For glucose (blood sugar), $C_6H_{12}O_6$,

(a) FALSE The percentages by mass of C and O are *different* than in CO. For one thing, CO contains no hydrogen.

(b) TRUE In dihydroxyacetone, $(CH_2OH)_2CO$ or $C_3H_6O_3$, the ratio of C:H:O = 3:6:3 or 1:2:1. In glucose, this ratio is C:H:O = 6:12:6 = 1:2:1. Thus, the ratios are the same.

(c) FALSE The proportions, by number of atoms, of C and O are the same in glucose. Since, however, C and O have different molar masses, their proportions by mass must be *different*.

(d) FALSE Each mole of glucose contains $(12\times1.008=)12.1$ g H. But each mole also contains 72.1 g C and 96.0 g O. Thus, the highest percentage, by mass, is that of O. The highest percentage, by number of atoms, is that of H.

16. **(E)** For sorbic acid, $C_6H_8O_2$,

(a) FALSE The C:H:O mole ratio is 3:4:1, but the mass ratio differs because moles of different elements have different molar masses.

(b) TRUE Since the two compounds have the same empirical formula, they have the same mass percent composition.

(c) TRUE Aspidinol, $C_{12}H_{16}O_4$, and sorbic acid have the same empirical formula, C_3H_4O.

(d) TRUE The ratio of H atoms to O atoms is $8:2 = 4:1$. Thus, the mass ratio is $(4\,\text{mol H}\times1\,\text{g H}):(1\,\text{mol O}\times16.0\,\text{g O}) = 4\,\text{g H}:16\,\text{g O} = 1\,\text{g H}:4\,\text{g O}$.

17. **(M)**

(a) $Cu(UO_2)_2(PO_4)_2\cdot 8\,H_2O$ has 1 Cu, 2 U, 2 P, 20 O, and 16 H, or a total of 41 atoms.

(b) By number, $Cu(UO_2)_2(PO_4)_2\cdot 8\,H_2O$ has a H to O ratio of 16:20 or 4:5 or 0.800 H atoms/O atom.

(c) By number, $Cu(UO_2)_2(PO_4)_2 \cdot 8\ H_2O$ has a Cu to P ratio of 1:2.

The mass ratio of Cu:P is $\dfrac{1\,\text{mol Cu}\times\dfrac{63.546\text{ g Cu}}{1\,\text{mol Cu}}}{2\text{ mol P}\times\dfrac{30.974\text{ g P}}{1\,\text{mol P}}}=1.026.$

(d) With a mass percent slightly greater than 50%, U has the largest mass percent, with oxygen coming in at ~34%.

$$\text{mass \% U}=\frac{\text{mass U in Cu(UO}_2)_2(\text{PO}_4)_2\cdot 8\text{ H}_2\text{O}}{\text{total mass of Cu(UO}_2)_2(\text{PO}_4)_2\cdot 8\text{ H}_2\text{O}}\times 100\%$$
$$=\frac{2\times 238.03\text{ g/mol}}{937.660\text{ g/mol}}\times 100\%$$
$$=50.77\%$$
$$\text{mass \% O}=\frac{\text{mass O in Cu(UO}_2)_2(\text{PO}_4)_2\cdot 8\text{H}_2\text{O}}{\text{total mass of Cu(UO}_2)_2(\text{PO}_4)_2\cdot 8\text{H}_2\text{O}}\times 100\%$$
$$=\frac{20\times 15.999\text{ g/mol}}{937.660\text{ g/mol}}\times 100\%$$
$$=34.13\%$$

(e) $$1.00\text{ g P}\times\frac{1\text{ mol P}}{30.974\text{ g P}}\times\frac{1\text{ mol Cu(UO}_2)_2(\text{PO}_4)_2\cdot 8\text{ H}_2\text{O}}{2\text{ mol P}}\times$$
$$\frac{937.666\text{ g Cu(UO}_2)_2(\text{PO}_4)_2\cdot 8\text{ H}_2\text{O}}{1\text{ mol Cu(UO}_2)_2(\text{PO}_4)_2\cdot 8\text{ H}_2\text{O}}$$
$= 15.1$ g of $Cu(UO_2)_2(PO_4)_2 \cdot 8\ H_2O$

18. (M)

(a) A formula unit of $\text{Ge}\left[\text{S}\left(\text{CH}_2\right)_4\text{CH}_3\right]_4$ contains:

1 Ge atom 4 S atoms $4(4+1)=20$ C atoms $4\left[4(2)+3\right]=44$ H atoms

For a total of $1+4+20+44=69$ atoms per formula unit

(b) $$\frac{\#\text{ of C atoms}}{\#\text{ of H atoms}}=\frac{20\,\text{C atoms}}{44\,\text{H atoms}}=\frac{5\,\text{C atoms}}{11\,\text{H atoms}}=0.455\,\text{C atom/H atom}$$

(c) $$\frac{\text{mass Ge}}{\text{mass S}}=\frac{1\,\text{mol Ge}\times\dfrac{72.630\text{ g Ge}}{1\text{ mol Ge}}}{4\,\text{mol S}\times\dfrac{32.06\text{ g S}}{1\,\text{mol S}}}=\frac{72.630\,\text{g Ge}}{128.24\,\text{g S}}=0.566\,\text{g Ge/g S}$$

(d) $$\text{mass of S}=1\,\text{mol Ge}\left[\text{S}\left(\text{CH}_2\right)_4\text{CH}_3\right]_4\times\frac{4\,\text{mol S}}{1\,\text{mol Ge}\left[\text{S}\left(\text{CH}_2\right)_4\text{CH}_3\right]_4}\times\frac{32.06\,\text{g S}}{1\,\text{mol S}}=128.264\text{ g S}$$

(e) $$\#\text{ of C atoms}=33.10\,\text{g cmpd}\times\frac{1\,\text{mol cmpd}}{485.44\,\text{g cmpd}}\times\frac{20\,\text{mol C}}{1\,\text{mol cmpd}}\times\frac{6.0221\times 10^{23}\,\text{C atoms}}{1\,\text{mol C}}$$
$$=8.212\times 10^{23}\,\text{C atoms}$$

Percent Composition of Compounds

19. **(E)** The information obtained in the course of calculating the molar mass is used to determine the mass percent of H in decane.

$$\text{molar mass } C_{10}H_{22} = \left(\frac{10\,\text{mol C}}{1\,\text{mol } C_{10}H_{22}} \times \frac{12.011\,\text{g C}}{1\,\text{mol C}}\right) + \left(\frac{22\,\text{mol H}}{1\,\text{mol } C_{10}H_{22}} \times \frac{1.008\,\text{g H}}{1\,\text{mol H}}\right)$$

$$= \frac{120.11\,\text{g C}}{1\,\text{mol } C_{10}H_{22}} + \frac{22.176\,\text{g H}}{1\,\text{mol } C_{10}H_{22}} = \frac{142.29\,\text{g}}{1\,\text{mol } C_{10}H_{22}}$$

$$\%\text{H} = \frac{22.176\,\text{g H/mol decane}}{142.29\,\text{g } C_{10}H_{22}/\text{mol decane}} \times 100\% = 15.369\%\,\text{H}$$

20. **(E)** Determine the mass of O in one mole of $Cu_2(OH)_2CO_3$ and the molar mass of $Cu_2(OH)_2CO_3$.

$$\text{mass O/mol } Cu_2(OH)_2CO_3 = \frac{5\text{ mol O}}{1\text{ mol } Cu_2(OH)_2CO_3} \times \frac{15.999\text{ g O}}{1\text{ mol O}} = 80.00\,\text{g O/mol } Cu_2(OH)_2CO_3$$

$$\text{molar mass } Cu_2(OH)_2CO_3 = (2\times 63.546\text{ g Cu}) + (5\times 15.999\text{ g O}) + (2\times 1.008\text{ g H}) + 12.011\text{ g C}$$

$$= 221.114\,\text{g/mol } Cu_2(OH)_2CO_3$$

$$\text{percent oxygen in sample} = \frac{79.995\text{ g}}{221.114\text{ g}} \times 100\% = 36.18\%\text{ O}$$

21. **(E)** $C(CH_3)_3CH_2CH(CH_3)_2$ has a molar mass of 114.231 g/mol and one mole contains 18.143 g of H.

$$\text{percent hydrogen in sample} = \frac{18.143\text{ g}}{114.231\text{g}} \times 100\% = 15.88\%\text{ H}$$

22. **(E)** Determine the mass of a mole of $Cr(NO_3)_3 \cdot 9\,H_2O$, and then the mass of water in a mole.

$$\text{molar mass } Cr(NO_3)_3 \cdot 9\,H_2O = 51.996\text{ g Cr} + (3\times 14.007\text{ g N}) + (18\times 15.999\text{ g O})$$

$$+ (18\times 1.008\text{ g H}) = 400.148\text{ g/mol } Cr(NO_3)_3 \cdot 9\,H_2O$$

$$\text{mass } H_2O = \frac{9\text{ mol } H_2O}{1\text{ mol } Cr(NO_3)_3 \cdot 9\,H_2O} \times \frac{18.015\,\text{g } H_2O}{1\text{ mol } H_2O} = 162.14\,\text{g } H_2O/\text{mol } Cr(NO_3)_3 \cdot 9\,H_2O$$

$$\frac{162.14\text{ g } H_2O/\text{mol } Cr(NO_3)_3 \cdot 9\,H_2O}{400.148\text{ g/mol } Cr(NO_3)_3 \cdot 9\,H_2O} \times 100\% = 40.5\%\ H_2O$$

23. **(E)** molar mass $= (20 \text{ mol C} \times 12.011 \text{ g C}) + (24 \text{ mol H} \times 1.008 \text{ g H}) + (2 \text{ mol N} \times 14.007 \text{ g N})$
$+ (2 \text{ mol O} \times 15.999 \text{ g O}) = 324.42 \text{ g/mol}$

$$\%\text{C} = \frac{240.22}{324.42} \times 100\% = 74.046\,\%\text{C} \qquad \%\text{H} = \frac{24.192}{324.42} \times 100\% = 7.4570\,\%\text{H}$$

$$\%\text{N} = \frac{28.014}{324.42} \times 100\% = 8.6351\,\%\text{N} \qquad \%\text{O} = \frac{31.998}{324.42} \times 100\% = 9.8631\,\%\text{O}$$

24. **(E)** The molar mass of $Cu(C_{18}H_{33}O_2)_2$ is 626.466 g/mol. One mole contains 66 H (66.528 g H), 36 C (432.396 g C), 4 O (63.996 g O) and 1 Cu (63.546 g Cu).

$$\%\text{C} = \frac{432.396 \text{ g}}{626.466 \text{ g}} \times 100\% = 69.0215\,\%\text{C} \qquad \%\text{H} = \frac{66.528 \text{ g}}{626.466 \text{ g}} \times 100\% = 10.619\,\%\text{H}$$

$$\%\text{Cu} = \frac{63.546 \text{ g}}{626.466 \text{ g}} \times 100\% = 10.144\,\%\text{Cu} \qquad \%\text{O} = \frac{63.996 \text{ g}}{626.466 \text{ g}} \times 100\% = 10.216\,\%\text{O}$$

25. **(E)** In each case, we first determine the molar mass of the compound, and then the mass of the indicated element in one mole of the compound. Finally, we determine the percent by mass of the indicated element to four significant figures.

(a) molar mass $\text{Pb}(\text{C}_2\text{H}_5)_4 = 207.2 \text{ g Pb} + (8 \times 12.011 \text{ g C}) + (20 \times 1.008 \text{ g H})$
$= 323.448 \text{ g/mol Pb}(\text{C}_2\text{H}_5)_4$

$$\text{mass Pb/mol Pb}(\text{C}_2\text{H}_5)_4 = \frac{1 \text{ mol Pb}}{1 \text{ mol Pb}(\text{C}_2\text{H}_5)_4} \times \frac{207.2 \text{ g Pb}}{1 \text{ mol Pb}} = 207.2 \text{ g Pb/mol Pb}(\text{C}_2\text{H}_5)_4$$

$$\%\,\text{Pb} = \frac{207.2 \text{ g Pb}}{323.448 \text{ g Pb}(\text{C}_2\text{H}_5)_4} \times 100\% = 64.06\%\ \text{Pb}$$

(b) molar mass $\text{Fe}_4\left[\text{Fe}(\text{CN})_6\right]_3 = (7 \times 55.845 \text{ g Fe}) + (18 \times 12.011 \text{ g C}) + (18 \times 14.007 \text{ g N})$
$= 859.253 \text{ g/mol Fe}_4\left[\text{Fe}(\text{CN})_6\right]_3$

$$\frac{\text{mass Fe}}{\text{mol Fe}_4\left[\text{Fe}(\text{CN})_6\right]_3} = \frac{7 \text{ mol Fe}}{1 \text{ mol Fe}_4\left[\text{Fe}(\text{CN})_6\right]_3} \times \frac{55.845 \text{ g Fe}}{1 \text{ mol Fe}} = 390.915 \text{ g Fe/mol Fe}_4\left[\text{Fe}(\text{CN})_6\right]_3$$

$$\%\text{Fe} = \frac{390.915 \text{ g Fe}}{859.253 \text{ g Fe}_4[\text{Fe}(\text{CN})_6]_3} \times 100\% = 45.495\%\ \text{Fe}$$

(c) molar mass $C_{55}H_{72}MgN_4O_5$

$= (55 \times 12.011\text{ g C}) + (72 \times 1.008\text{ g H}) + (1 \times 24.305\text{ g Mg}) + (4 \times 14.007\text{ g N}) + (5 \times 15.999\text{ g O})$

$= 893.509\text{ g/mol } C_{55}H_{72}MgN_4O_5$

$$\frac{\text{mass Mg}}{\text{mol } C_{55}H_{72}MgN_4O_5} = \frac{1\text{ mol Mg}}{1\text{ mol } C_{55}H_{72}MgN_4O_5} \times \frac{24.305\text{ g Mg}}{1\text{ mol Mg}} = \frac{24.305\text{ g Mg}}{\text{mol } C_{55}H_{72}MgN_4O_5}$$

$$\%\text{Mg} = \frac{24.305\text{ g Mg}}{893.509\text{g } C_{55}H_{72}MgN_4O_5} \times 100\% = 2.7202\%\text{ Mg}$$

26. **(E)**

(a) $$\%\text{Zr} = \frac{1\text{ mol Zr}}{1\text{ mol ZrSiO}_4} \times \frac{1\text{ mol ZrSiO}_4}{183.31\text{ g ZrSiO}_4} \times \frac{91.224\text{ g Zr}}{1\text{ mol Zr}} \times 100\% = 49.765\%\text{ Zr}$$

(b) $$\%\text{Be} = \frac{3\text{ mol Fe}}{1\text{ mol Be}_3\text{Al}_2\text{Si}_6\text{O}_{18}} \times \frac{1\text{ mol Be}_3\text{Al}_2\text{Si}_6\text{O}_{18}}{537.502\text{ g Be}_3\text{Al}_2\text{Si}_6\text{O}_{18}} \times \frac{9.01218\text{ g Be}}{1\text{ mol Be}} \times 100\%$$

% Be = 5.03004% Be

(c) $$\%\text{Fe} = \frac{3\text{ mol Fe}}{1\text{ mol Fe}_3\text{Al}_2\text{Si}_3\text{O}_{12}} \times \frac{1\text{ mol Fe}_3\text{Al}_2\text{Si}_3\text{O}_{12}}{497.753\text{ g Fe}_3\text{Al}_2\text{Si}_3\text{O}_{12}} \times \frac{55.845\text{ g Fe}}{1\text{ mol Fe}} \times 100\%$$

%Fe = 33.658% Fe

(d) $$\%\text{S} = \frac{1\text{ mol S}}{1\text{ mol Na}_4\text{SSi}_3\text{Al}_3\text{O}_{12}} \times \frac{1\text{ mol Na}_4\text{SSi}_3\text{Al}_3\text{O}_{12}}{481.219\text{ g Na}_4\text{SSi}_3\text{Al}_3\text{O}_{12}} \times \frac{32.06\text{ g S}}{1\text{ mol S}} \times 100\%$$

%S = 6.6622% S

27. **(M)** Oxide with the largest %Cr will have the largest number of moles of Cr per mole of oxygen.

CrO: $\dfrac{1\text{ mol Cr}}{1\text{ mol O}} = 1\text{ mol Cr/mol O}$ Cr_2O_3: $\dfrac{2\text{ mol Cr}}{3\text{ mol O}} = 0.667\text{ mol Cr/mol O}$

CrO_2: $\dfrac{1\text{ mol Cr}}{2\text{ mol O}} = 0.500\text{ mol Cr/mol O}$ CrO_3: $\dfrac{1\text{ mol Cr}}{3\text{ mol O}} = 0.333\text{ mol Cr/mol O}$

Arranged in order of increasing %Cr: $CrO_3 < CrO_2 < Cr_2O_3 < CrO$

28. **(M)** For SO_2 and Na_2S, a mole of each contains a mole of S and two moles of another element; in the case of SO_2, the other element (oxygen) has a smaller atomic mass than the other element in Na_2S (Na), causing SO_2 to have a higher mass percent sulfur. For S_2Cl_2 and $Na_2S_2O_3$, a mole of each contains two moles of S; for S_2Cl_2, the rest of the mole has a mass of 71.0 g; while for $Na_2S_2O_3$, it would be $(2 \times 23) + (3 \times 16) = 94$ g. Sulfur makes up the greater proportion of the mass in S_2Cl_2, giving it the larger percent of S. Now we compare

SO_2 and S_2Cl_2: SO_2 contains one mole of S (32.1 g) and S_2Cl_2 contains two moles of S (64.2 g). In S_2Cl_2 the remainder of a mole has a mass of 71.0 g, while in SO_2 the remainder of a mole would be $2 \times 16.0 = 32.0$ g. Thus, SO_2 has the highest percent of S so far. For CH_3CH_2SH compared to SO_2, we see that both compounds have one S atom, SO_2 has two O atoms (each with a molar mass of ~16 g mol^{-1}), and CH_3CH_2SH effectively has two CH_3 groups (each CH_3 group with a mass of ~15 g mol^{-1}). Thus, CH_3CH_2SH has the highest percentage sulfur by mass of the compounds listed.

Chemical Formulas from Percent Composition

<u>29</u>. **(M)** SO_3 (40.05% S) and S_2O (80.0% S) (2 O atoms $\approx$ 1 S atom in terms of atomic masses) Note the molar masses are quite close (within 0.05 g/mol).

30. **(M)** The element chromium has an atomic mass of 52.0 u. Thus, there can only be one chromium atom per formula unit of the compound. (Two atoms of chromium have a mass of 104 u, more than the formula mass of the compound.) Three of the four remaining atoms in the formula unit must be oxygen. Thus, the oxide is CrO_3, chromium(VI) oxide.

<u>**31.**</u> **(M)** Determine the % oxygen by difference.

$\%\text{O} = 100.00\% - 45.27\%\ \text{C} - 9.50\%\ \text{H} = 45.23\%\ \text{O}$

The following calculations are based on a 100.00 g sample.

$$\text{mol O} = 45.23\ \text{g} \times \frac{1\ \text{mol O}}{15.999\ \text{g O}} = 2.827\ \text{mol O} \quad \div 2.827 \rightarrow 1.000\ \text{mol O}$$

$$\text{mol C} = 45.27\ \text{g C} \times \frac{1\ \text{mol C}}{12.011\ \text{g C}} = 3.769\ \text{mol C} \quad \div 2.827 \rightarrow 1.333\ \text{mol C}$$

$$\text{mol H} = 9.50\ \text{g H} \times \frac{1\ \text{mol H}}{1.008\ \text{g H}} = 9.42\ \text{mol H} \quad \div 2.827 \rightarrow 3.33\ \text{mol H}$$

Multiply all amounts by 3 to obtain integers. Empirical formula is $C_4H_{10}O_3$.

32. **(M)** We base our calculation on 100.0 g of monosodium glutamate.

$$13.6\ \text{g Na} \times \frac{1\ \text{mol Na}}{22.99\ \text{g Na}} = 0.592\ \text{mol Na} \quad \div 0.592 \rightarrow 1.00\ \text{mol Na}$$

$$35.5\ \text{g C} \times \frac{1\ \text{mol C}}{12.01\ \text{g C}} = 2.96\ \text{mol C} \quad \div 0.592 \rightarrow 5.00\ \text{mol C}$$

$$4.8\ \text{g H} \times \frac{1\ \text{mol H}}{1.01\ \text{g H}} = 4.8\ \text{mol H} \quad \div 0.592 \rightarrow 8.1\ \text{mol H}$$

$$8.3\ \text{g N} \times \frac{1\ \text{mol N}}{14.0\ \text{g N}} = 0.59\ \text{mol N} \quad \div 0.592 \rightarrow 1.0\ \text{mol N}$$

$$37.8\ \text{g O} \times \frac{1\ \text{mol O}}{16.00\ \text{g O}} = 2.36\ \text{mol O} \quad \div 0.592 \rightarrow 3.99\ \text{mol O}$$

Empirical formula : $NaC_5H_8NO_4$

33. **(M)**

(a) $74.01\text{g C} \times \dfrac{1 \text{ mol C}}{12.011 \text{ g C}} = 6.162 \text{ mol C} \qquad \div 1.298 \rightarrow 4.747 \text{ mol C}$

$5.23 \text{ g H} \times \dfrac{1 \text{ mol H}}{1.008 \text{ g H}} = 5.19 \text{ mol H} \qquad \div 1.298 \rightarrow 4.00 \text{ mol H}$

$20.76 \text{ g O} \times \dfrac{1 \text{ mol O}}{15.999 \text{ g O}} = 1.298 \text{ mol O} \qquad \div 1.298 \rightarrow 1.000 \text{ mol O}$

Multiply each of the mole numbers by 4 to obtain an empirical formula of $C_{19}H_{16}O_4$.

(b)

$$\left.\begin{array}{l}
39.98 \text{ g C} \times \dfrac{1 \text{ mol C}}{12.011 \text{ g C}} = 3.328\underline{6} \text{ mol C} \quad \div 0.7399 \rightarrow 4.499 \text{ mol C} \\
3.73 \text{ g H} \times \dfrac{1 \text{ mol H}}{1.008 \text{ g H}} = 3.70 \text{ mol H} \quad \div 0.7399 \rightarrow 5.00 \text{ mol H} \\
20.73 \text{ g N} \times \dfrac{1 \text{ mol N}}{14.007 \text{ g N}} = 1.480 \text{ mol N} \quad \div 0.7399 \rightarrow 2.000 \text{ mol N} \\
11.84 \text{ g O} \times \dfrac{1 \text{ mol O}}{15.999 \text{ g O}} = 0.7400 \text{ mol O} \quad \div 0.7399 \rightarrow 1.000 \text{ mol O} \\
23.72 \text{ g S} \times \dfrac{1 \text{ mol S}}{32.06 \text{ g S}} = 0.7399 \text{ mol S} \quad \div 0.7399 \rightarrow 1.000 \text{ mol S}
\end{array}\right\}$$

Multiply by 2 to obtain the empirical formula

$C_9H_{10}N_4O_2S_2$

34. **(M)**

(a) $95.21 \text{ g C} \times \dfrac{1 \text{ mol C}}{12.011 \text{ g C}} = 7.927 \text{ mol C} \quad \div 4.74 \rightarrow 1.67 \text{ mol C}$

$4.79 \text{ g H} \times \dfrac{1 \text{ mol H}}{1.008 \text{ g H}} = 4.75 \text{ mol H} \quad \div 4.75 \rightarrow 1.00 \text{ mol H}$

Multiply each of the mole numbers by 3 to obtain an empirical formula of C_5H_3.

(b) Each percent is numerically equal to the mass of that element present in 100.00 g of the compound. These masses then are converted to amounts of the elements, in moles.

$\text{amount C} = 38.37 \text{g C} \times \dfrac{1 \text{ mol C}}{12.011 \text{g C}} = 3.195 \text{ mol C} \qquad \div 0.491 \rightarrow 6.51 \text{ mol C}$

$\text{amount H} = 1.49 \text{ g H} \times \dfrac{1 \text{ mol H}}{1.008 \text{ g H}} = 1.48 \text{ mol H} \qquad \div 0.491 \rightarrow 3.01 \text{ mol H}$

$\text{amount Cl} = 52.28 \text{ g Cl} \times \dfrac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} = 1.475 \text{ mol Cl} \quad \div 0.491 \rightarrow 3.004 \text{ mol Cl}$

$\text{amount O} = 7.86 \text{ g O} \times \dfrac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.491 \text{ mol O} \qquad \div 0.491 \rightarrow 1.00 \text{ mol O}$

Multiply each number of moles by 2 to obtain the empirical formula: $C_{13}H_6Cl_6O_2$.

35. Convert each percentage into the mass in 100.00 g, and then to the moles of that element.

$$94.34 \text{ g C} \times \frac{1 \text{ mol C}}{12.011 \text{ g C}} = 7.854 \text{ mol C} \qquad \div 5.62 = 1.40 \text{ mol C} \times 5 = 7.00$$

$$5.66 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 5.62 \text{ mol H} \qquad \div 5.62 = 1.00 \text{ mol H} \times 5 = 5.00$$

Multiply by 5 to achieve whole number ratios. The empirical formula is C_5H_7, and the formula mass [(7 × 12.011 g C) + (5 × 1.008 g H)] = 89.117 u. Since this empirical molar mass is one-half of the 178 u, the correct molecular mass, the molecular formula must be twice the empirical formula. Molecular formula: $C_{14}H_{10}$

36. **(M)** The percent of selenium in each oxide is found by difference.

First oxide: $\%\text{Se} = 100.0\% - 28.8\% \quad \text{O} = 71.2\% \text{ Se}$

A 100.0 gram sample would contain 28.8 g O and 71.2 g Se

$$28.8 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 1.80 \text{ mol O} \qquad \div 0.902 \quad \rightarrow 2.00 \text{ mol O}$$

$$71.2 \text{ g Se} \times \frac{1 \text{ mol Se}}{78.96 \text{ g Se}} = 0.902 \text{ mol Se} \qquad \div 0.902 \quad \rightarrow 1.00 \text{ mol Se}$$

The empirical formula is SeO_2. An appropriate name is selenium dioxide.

Second oxide: $\%\text{Se} = 100.0\% - 37.8\%\text{O} = 62.2\% \text{ Se}$

A 100.0 gram sample would contain 37.8 g O and 62.2 g Se

$$37.8 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.36 \text{ mol O} \qquad \div 0.788 \quad \rightarrow 2.99 \text{ mol O}$$

$$62.2 \text{ g Se} \times \frac{1 \text{ mol Se}}{78.96 \text{ g Se}} = 0.788 \text{ mol Se} \qquad \div 0.788 \quad \rightarrow 1.00 \text{ mol Se}$$

The empirical formula is SeO_3. An appropriate name is selenium trioxide.

37. **(M)** Determine the mass of oxygen by difference. Then convert all masses to amounts in moles. oxygen $\text{mass} = 100.00 \text{ g} - 73.27 \text{ g C} - 3.84 \text{ g H} - 10.68 \text{ g N} = 12.21 \text{ g O}$

$$\text{amount C} = 73.27 \text{ g C} \times \frac{1 \text{ mol C}}{12.011 \text{ g C}} = 6.100 \text{ mol C} \qquad \div 0.7625 \quad \rightarrow 8.000 \text{ mol C}$$

$$\text{amount H} = 3.84 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 3.81 \text{ mol H} \qquad \div 0.7625 \quad \rightarrow 5.00 \text{ mol H}$$

$$\text{amount N} = 10.68 \text{ g N} \times \frac{1 \text{ mol N}}{14.007 \text{ g N}} = 0.7625 \text{ mol N} \qquad \div 0.7625 \quad \rightarrow 1.000 \text{ mol N}$$

$$\text{amount O} = 12.21 \text{ g O} \times \frac{1 \text{ mol O}}{15.999 \text{ g O}} = 0.7632 \text{ mol O} \qquad \div 0.7625 \quad \rightarrow 1.001 \text{ mol O}$$

The empirical formula is C_8H_5NO, which has an empirical mass of 131 u. This is almost exactly half the molecular mass of 262.3 u. Thus, the molecular formula is twice the empirical formula and is $C_{16}H_{10}N_2O_2$.

38. **(M)** Convert each percentage into the mass in 100.00 g, and then to the moles of that element.

$$\text{amount C} = 44.45\text{ g C} \times \frac{1\text{mol C}}{12.011\text{ g C}} = 3.701\text{ mol C} \quad \div 3.70 \quad \rightarrow 1.00\text{ mol C}$$

$$\text{amount H} = 3.73\text{ g H} \times \frac{1\text{ mol H}}{1.008\text{ g H}} = 3.70\text{ mol H} \quad \div 3.70 \quad \rightarrow 1.00\text{ mol H}$$

$$\text{amount N} = 51.82\text{ g N} \times \frac{1\text{ mol N}}{14.007\text{ g N}} = 3.700\text{ mol N} \quad \div 3.70 \quad \rightarrow 1.00\text{ mol N}$$

The empirical formula is CHN, which has an empirical mass of 27.026 u. This is one fifth the molecular mass of 135.14 u. Thus, the molecular formula is five times greater than the empirical formula and is $C_5H_5N_5$.

39. **(M)** The molar mass of element X has the units of grams per mole. We can determine the amount, in moles of Cl, and convert that to the amount of X, equivalent to 25.0 g of X.

$$\text{molar mass} = \frac{25.0\text{ g X}}{75.0\text{ g Cl}} \times \frac{35.45\text{ g Cl}}{1\text{ mol Cl}} \times \frac{4\text{ mol Cl}}{1\text{ mol X}} = \frac{47.3\text{ g X}}{1\text{ mol X}}$$

The atomic mass is 47.3 u. This atomic mass is close to that of the element titanium, which therefore is identified as element X.

40. **(M)** The molar mass of element X has the units of grams per mole. We can determine the amount, relative to the mass percent Cl. Assume 1 mole of compound. This contains 15.9994 g O and 70.905 g Cl. The following relation must hold true.

$$\text{mass percent Cl} = 59.6\% = 0.596 = \frac{\text{mass Cl in one mole}}{\text{mass of one mole of XOCl}_2} = \frac{70.90\text{ g}}{(\text{X} + 15.999 + 70.90)\text{ g}}$$

$$0.596 = \frac{70.90}{(\text{X} + 86.89)} \quad \text{or} \quad 0.596\text{X} + 51.79 = 70.90 \qquad 0.596\text{X} = 19.11$$

$$\text{Hence, X} = \frac{19.11}{0.596} = 32.0\underline{6} \qquad \text{The atomic mass of X} = 32.1\text{ g mol}^{-1} \qquad \text{X is the element sulfur}$$

41. **(M)** Consider 100 g of chlorophyll, which contains 2.72 g of Mg. To answer this problem, we must take note of the fact that 1 mole of chlorophyll contains 1 mole of Mg.

$$\frac{100\text{ g chlorophyll}}{2.72\text{ g Mg}} \times \frac{24.305\text{ g Mg}}{1\text{ mol Mg}} \times \frac{1\text{ mol Mg}}{1\text{ mol chlorophyll}} = 894\text{ g mol}^{-1}$$

Therefore, the molecular mass of chlorophyll is 894 u.

42. **(D)** Compound I has a molecular mass of 137 u. We are told that chlorine constitutes 77.5% of the mass, so the mass of chlorine in each molecule is $137\text{ u} \times \frac{77.5}{100} = 106\text{ u}$.

This corresponds to three chlorine atoms (106 u ÷ 35.45 u/Cl atom = 2.99 or 3 Cl atoms). The remaining 31 u, (137 u – 106 u), is the mass for element X in one molecule of Compound I. Compound II has 85.1% chlorine by mass, so the mass of chlorine in each molecule of Compound II is $208\text{ u} \times \frac{85.1}{100} = 177\text{ u}$.

This corresponds to five Cl atoms (177 u ÷ 35.45 u/Cl atom = 4.99 ~5 chlorine atoms). The remaining mass is 31 u (208 u – 177 u), which is very close to the mass of X found in each molecule of Compound I. Thus, we have two compounds: X_nCl_3, which has a molecular mass of 137 u, and X_nCl_5, which has a molecular mass of 208 u.

(We also know that the mass of X in both molecular species is ~31 u). If we assume that $n = 1$ in the formulas above, then element X must be phosphorus (30.974 u) and the formulas for the compounds are PCl_3 (Compound I) and PCl_5 (Compound II).

Combustion Analysis

43. **(M)**

(a) First we determine the mass of carbon and of hydrogen present in the sample. Remember that a hydrocarbon contains only hydrogen and carbon.

$$0.6260\,\text{g CO}_2 \times \frac{1\,\text{mol CO}_2}{44.009\,\text{g CO}_2} \times \frac{1\,\text{mol C}}{1\,\text{mol CO}_2} = 0.01422\,\text{mol C} \times \frac{12.011\,\text{g C}}{1\,\text{mol C}} = 0.1708\,\text{g C}$$

$$0.1602\,\text{g H}_2\text{O} \times \frac{1\,\text{mol H}_2\text{O}}{18.015\,\text{g H}_2\text{O}} \times \frac{2\,\text{mol H}}{1\,\text{mol H}_2\text{O}} = 0.01779\,\text{mol H} \times \frac{1.008\,\text{g H}}{1\,\text{mol H}} = 0.01793\,\text{g H}$$

Then the % C and % H are found.

$$\%\text{C} = \frac{0.1708}{0.1888\,\text{g cmpd}} \times 100\% = 90.47\%\,\text{C} \quad \%\text{H} = \frac{0.01793\,\text{g H}}{0.1888\,\text{g cmpd}} \times 100\% = 9.497\%\,\text{H}$$

(b) Use the moles of C and H from part (a), and divide both by the smallest value, namely 0.01422 mol. Thus 0.01422 mol C ÷ 0.01422 mol = 1.000 mol H;
0.01779 mol H ÷ 0.01422 mol = 1.250 mol H.

The empirical formula is obtained by multiplying these mole numbers by 4. It is C_4H_5.

(c) The molar mass of the empirical formula C_4H_5 is [4 × 12.011 + 5 × 1.008] = 53.084 g/mol. This value is 1/2 of the actual molar mass. The molecular formula is twice the empirical formula. ∴ Molecular formula: C_8H_{10}.

44. **(M)** Determine the mass of carbon and of hydrogen present in the sample.

$$1.1518\,\text{g CO}_2 \times \frac{1\,\text{mol CO}_2}{44.090\,\text{g CO}_2} \times \frac{1\,\text{mol C}}{1\,\text{mol CO}_2} = 0.026172\,\text{mol C} \times \frac{12.011\,\text{g C}}{1\,\text{mol C}} = 0.3144\,\text{g C}$$

$$0.2694\,\text{g H}_2\text{O} \times \frac{1\,\text{mol H}_2\text{O}}{18.015\,\text{g H}_2\text{O}} \times \frac{2\,\text{mol H}}{1\,\text{mol H}_2\text{O}} = 0.029908\,\text{mol H} \times \frac{1.008\,\text{g H}}{1\,\text{mol H}} = 0.03015\,\text{g H}$$

(a) The percent composition can be determined using the masses of C and H.

$$\%\text{C} = \frac{0.3144\,\text{g C}}{0.4039\,\text{g cmpd}} \times 100\% = 77.83\%\,\text{C} \quad \%\text{H} = \frac{0.030148\,\text{g H}}{0.4039\,\text{g cmpd}} \times 100\% = 7.4641\%\,\text{H}$$

% O = 100% – 77.84% – 7.4641% = 14.71%

These percents can be used in determining the empirical formula if one wishes.

(b) To find the empirical formula, determine the mass of oxygen by difference, and its amount in moles. Mass O = 0.04039 g – 0.3144 g – 0.0302 g = 0.0593 g

$$\left.\begin{array}{l} 0.0593 \text{ g O} \times \dfrac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.00371 \text{ mol O} \div 0.00371 \rightarrow 1.00 \text{ mol O} \\ \qquad\qquad\qquad\qquad 0.026172 \text{ mol C} \div 0.00371 \rightarrow 7.05 \text{ mol C} \\ \qquad\qquad\qquad\qquad 0.029908 \text{ mol H} \div 0.00371 \rightarrow 8.06 \text{ mol H} \end{array}\right\}$$

Empirical formula is C_7H_8O.

(c) The molecular formula is found by realizing that a mole of empirical units has a mass of $(7 \times 12.0 \text{ g C} + 8 \times 1.0 \text{ g H} + 16.0 \text{ g O}) = 108.0$ g. Since this agrees with the molecular mass, the molecular formula is the same as the empirical formula: C_7H_8O.

45. **(M)** First, determine the mass of carbon and hydrogen present in the sample.

$$0.458 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.01 \text{ g } CO_2} \times \frac{1 \text{ mol C}}{1 \text{ mol } CO_2} = 0.0104 \text{ mol C} \times \frac{12.011 \text{ g C}}{1 \text{ mol C}} = 0.125 \text{ g C}$$

$$0.374 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.015 \text{ g } H_2O} \times \frac{2 \text{ mol H}}{1 \text{ mol } H_2O} = 0.0415 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 0.0418 \text{ g H}$$

Then, the mass of N that this sample would have produced is determined.
(Note that this is also the mass of N_2 produced in the reaction.)

$$0.226 \text{ g } N_2 \times \frac{0.312 \text{ g 1st sample}}{0.486 \text{ g 2nd sample}} = 0.145 \text{ g } N_2$$

From which we can calculate the mass of N in the sample.

$$0.145 \text{ g } N_2 \times \frac{1 \text{ mol } N_2}{28.014 \text{ g } N_2} \times \frac{2 \text{ mol N}}{1 \text{ mol } N_2} \times \frac{14.007 \text{ g N}}{1 \text{ mol N}} = 0.145 \text{ g N}$$

We may alternatively determine the mass of N by difference:
0.312 g – 0.125 g C – 0.0418 g H = 0.145 g N

Then, we can calculate the relative number of moles of each element.

$$\left.\begin{array}{ll} 0.145 \text{ g N} \times \dfrac{1 \text{ mol N}}{14.007 \text{ g N}} = & 0.0104 \text{ mol N} \div 0.0104 \rightarrow 1.00 \text{ mol N} \\ & 0.0104 \text{ mol C} \div 0.0104 \rightarrow 1.00 \text{ mol C} \\ & 0.0415 \text{ mol H} \div 0.0104 \rightarrow 4.01 \text{ mol H} \end{array}\right\}$$

Thus, the empirical formula is CH_4N

46. **(M)** Thiophene contains only carbon, hydrogen, and sulfur, so there is no need to determine the mass of oxygen by difference. We simply determine the amount of each element from the mass of its combustion product.

$$2.7224 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.009 \text{ g } CO_2} \times \frac{1 \text{ mol C}}{1 \text{ mol } CO_2} = 0.061859 \text{ mol C} \div 0.01548 \rightarrow 3.996 \text{ mol C}$$

$$0.5575 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.015 \text{ g } H_2O} \times \frac{2 \text{ mol H}}{1 \text{ mol } H_2O} = 0.06189 \text{ mol H} \div 0.01548 \rightarrow 3.999 \text{ mol H}$$

$$0.9915 \text{ g } SO_2 \times \frac{1 \text{ mol } SO_2}{64.058 \text{ g } SO_2} \times \frac{1 \text{ mol S}}{1 \text{ mol } SO_2} = 0.01548 \text{ mol S} \div 0.01548 \rightarrow 1.00 \text{ mol S}$$

The empirical formula of thiophene is C_4H_4S.

47. **(M)** Each mole of CO_2 is produced from a mole of C. Therefore, the compound with the largest number of moles of C per mole of the compound will produce the largest amount of CO_2 and, thus, also the largest mass of CO_2. Of the compounds listed, namely CH_4, C_2H_5OH, $C_{10}H_8$, and C_6H_5OH, $C_{10}H_8$ has the largest number of moles of C per mole of the compound and will produce the greatest mass of CO_2 per mole on complete combustion.

48. **(M)** The compound that produces the largest mass of water per gram of the compound will have the largest amount of hydrogen per gram of the compound. Thus, we need to compare the ratios of amount of hydrogen per mole to the molar mass for each compound. Note that C_2H_5OH has as much H per mole as does C_6H_5OH, but C_6H_5OH has a higher molar mass. Thus, C_2H_5OH produces more H_2O per gram than does C_6H_5OH. Notice also that CH_4 has 4 H's per C, while $C_{10}H_8$ has 8 H's per 10 C's or 0.8 H per C. Thus CH_4 will produce more H_2O than will $C_{10}H_8$. Thus, we are left with comparing CH_4 to C_2H_5OH. The O in the second compound has about the same mass (16 u) as does C (12 u). Thus, in CH_4 there are 4 H's per C, while in C_2H_5OH there are about 2 H's per C. Thus CH_4 will produce the most water per gram on combustion, of all four compounds.

49. **(M)** The molecular formula for $CH_3CH(OH)CH_2CH_3$ is $C_4H_{10}O$. Here we will use the fact that $C_4H_{10}O$ has a molar mass of 74.123 g/mol to calculate the masses of CO_2 and H_2O:

Mass of CO_2:

Conversion pathway approach:

$$1.562 \text{ g } C_4H_{10}O \times \frac{1 \text{ mol } C_4H_{10}O}{74.123 \text{ g } C_4H_{10}O} \times \frac{4 \text{ mol C}}{1 \text{ mol } C_4H_{10}O} \times \frac{1 \text{ mol } CO_2}{1 \text{ mol C}} \times \frac{44.009 \text{ g } CO_2}{1 \text{ mol } CO_2} = 3.710 \text{ g } CO_2$$

Stepwise approach:

$$1.562\text{ g }C_4H_{10}O \times \frac{1\text{ mol }C_4H_{10}O}{74.123\text{ g }C_4H_{10}O} = 0.02107\text{ mol }C_4H_{10}O \times \frac{4\text{ mol C}}{1\text{ mol }C_4H_{10}O} = 0.08429\text{ mol C}$$

$$0.08429\text{ mol C} \times \frac{1\text{ mol }CO_2}{1\text{ mol C}} = 0.08429\text{ mol }CO_2 \times \frac{44.009\text{ g }CO_2}{1\text{ mol }CO_2} = 3.710\text{ g }CO_2$$

Mass of H_2O:

Conversion pathway approach:

$$1.562\text{ g }C_4H_{10}O \times \frac{1\text{ mol }C_4H_{10}O}{74.123\text{ g }C_4H_{10}O} \times \frac{10\text{ mol H}}{1\text{ mol }C_4H_{10}O} \times \frac{1\text{ mol }H_2O}{2\text{ mol H}} \times \frac{18.015\text{ g }H_2O}{1\text{ mol }H_2O} = 1.898\text{ g }H_2O$$

Stepwise approach:

$$1.562\text{ g }C_4H_{10}O \times \frac{1\text{ mol }C_4H_{10}O}{74.123\text{ g }C_4H_{10}O} = 0.02107\text{ mol }C_4H_{10}O$$

$$0.02107\text{ mol }C_4H_{10}O \times \frac{10\text{ mol H}}{1\text{ mol }C_4H_{10}O} = 0.2107\text{ mol H}$$

$$0.2107\text{ mol H} \times \frac{1\text{ mol }H_2O}{2\text{ mol H}} = 0.1054\text{ mol }H_2O \times \frac{18.015\text{ g }H_2O}{1\text{ mol }H_2O} = 1.898\text{ g }H_2O$$

50. **(D)** moles of $C_2H_6S = 3.15\text{ mL} \times \frac{0.84\text{ g }C_2H_6S}{1\text{ mL }C_2H_6S} \times \frac{1\text{ mol }C_2H_6S}{62.13\text{ g }C_2H_6S} = 0.042\underline{6}\text{ mol }C_2H_6S$

Thus, the mass of CO_2 expected is

$$= 0.042\underline{6}\text{ mol }C_2H_6S \times \frac{2\text{ mol }CO_2}{1\text{ mol }C_2H_6S} \times \frac{44.009\text{ g }CO_2}{1\text{ mol }CO_2} = 3.7\underline{5}\text{ g of }CO_2(g)$$

The mass of $SO_2(g)$ expected from the complete combustion is

$$= 0.042\underline{6}\text{ mol }C_2H_6S \times \frac{1\text{ mol }SO_2}{1\text{ mol }C_2H_6S} \times \frac{64.058\text{ g }SO_2}{1\text{ mol }H_2O} = 2.7\underline{3}\text{ g of }SO_2(g)$$

The mass of $H_2O(l)$ expected from the complete combustion is

$$= 0.042\underline{6}\text{ mol }C_2H_6S \times \frac{6\text{ mol H}}{1\text{ mol }C_2H_6S} \times \frac{1\text{ mol }H_2O}{2\text{ mol H}} \times \frac{18.015\text{ g }H_2O}{1\text{ mol }H_2O} = 2.3\text{ g of }H_2O(l)$$

Oxidation States

51. **(E)** The oxidation state (O.S.) is given first, followed by the explanation for its assignment.

(a)	$C = -4$ in CH_4	H has an oxidation state of $+1$ in its nonmetal compounds. (Remember that the sum of the oxidation states in a neutral compound equals 0.)
(b)	$S = +4$ in SF_4	F has $O.S. = -1$ in its compounds.
(c)	$O = -1$ in Na_2O_2	Na has $O.S. = +1$ in its compounds.
(d)	$C = 0$ in $C_2H_3O_2^-$	H has $O.S. = +1$ in its nonmetal compounds; that of $O = -2$ (usually). (Remember that the sum of the oxidation states in a polyatomic ion equals the charge on that ion.)
(e)	$Fe = +6$ in FeO_4^{2-}	O has $O.S. = -2$ in most of its compounds (especially metal containing compounds).

52. **(E)** The oxidation state of sulfur in each species is determined below. Remember that the oxidation state of O is -2 in its compounds. And the sum of the oxidation states in an ion equals the charge on that ion.

(a) $S = +4$ in SO_3^{2-} **(b)** $S = +2$ in $S_2O_3^{2-}$ **(c)** $S = +7$ in $S_2O_8^{2-}$

(d) $S = +6$ in HSO_4^- **(e)** $S = -2.5$ in $S_4O_6^{2-}$

53. **(E)** Remember that the oxidation state of oxygen is usually -2 in its compounds. Cr^{3+} and O^{2-} form Cr_2O_3, chromium(III) oxide. Cr^{4+} and O^{2-} form CrO_2, chromium(IV) oxide. Cr^{6+} and O^{2-} form CrO_3, chromium(VI) oxide.

54. **(E)** Remember that oxygen usually has an oxidation state of -2 in its compounds.

$N = +1$ in N_2O, dinitrogen monoxide

$N = +2$ in NO, nitric oxide or nitrogen monoxide

$N = +3$ in N_2O_3, dinitrogen trioxide

$N = +4$ in NO_2, nitrogen dioxide

$N = +5$ in N_2O_5, dinitrogen pentoxide

55. **(E)**

(a)	$O = +2$ in OF_2	F has an oxidation state of -1 in its compounds.
(b)	$O = +1$ in O_2F_2	F has $O.S. = -1$ in its compounds.
(c)	$O = \frac{-1}{2}$ in CsO_2	Cs has $O.S. = +1$ in its compounds.
(d)	$O = -1$ in BaO_2	Ba has $O.S. = +2$ in its compounds.

56. **(E)**

(a)	MgH_2	Mg = +2	H = –1	
(b)	CsO_3	Cs = +1	O = –1/3	
(c)	HOF	H = +1	F = –1	O = 0
(d)	$NaAlH_4$	Na = +1	H = –1*	Al = +3

*Note: in metal hydrides, H has an oxidation number of –1.

Nomenclature

57. **(E)**

(a)	SrO	strontium oxide	**(b)**	ZnS	zinc sulfide
(c)	K_2CrO_4	potassium chromate	**(d)**	Cs_2SO_4	cesium sulfate
(e)	Cr_2O_3	chromium(III) oxide	**(f)**	$Fe_2(SO_4)_3$	iron(III) sulfate
(g)	$Mg(HCO_3)_2$	magnesium hydrogen carbonate or magnesium bicarbonate	**(h)**	$(NH_4)_2HPO_4$	ammonium hydrogen phosphate
(i)	$Ca(HSO_3)_2$	calcium hydrogen sulfite	**(j)**	$Cu(OH)_2$	copper(II) hydroxide
(k)	HNO_3	nitric acid	**(l)**	$KClO_4$	potassium perchlorate
(m)	$HBrO_3$	bromic acid	**(n)**	H_3PO_3	phosphorous acid

58. **(E)**

(a)	$Ba(NO_3)_2$	barium nitrate	**(b)**	HNO_2	nitrous acid
(c)	CrO_2	chromium(IV) oxide	**(d)**	KIO_3	potassium iodate
(e)	LiCN	lithium cyanide	**(f)**	KIO	potassium hypoiodite
(g)	$Fe(OH)_2$	iron(II) hydroxide	**(h)**	$Ca(H_2PO_4)_2$	calcium dihydrogen phosphate
(i)	H_3PO_4	phosphoric acid	**(j)**	$NaHSO_4$	sodium hydrogen sulfate
(k)	$Na_2Cr_2O_7$	sodium dichromate	**(l)**	$NH_4C_2H_3O_2$	ammonium acetate
(m)	MgC_2O_4	magnesium oxalate	**(n)**	$Na_2C_2O_4$	sodium oxalate

59. **(E)**

(a)	CS_2	carbon disulfide	**(b)**	SiF_4	silicon tetrafluoride
(c)	ClF_5	chlorine pentafluoride	**(d)**	N_2O_5	dinitrogen pentoxide
(e)	SF_6	sulfur hexafluoride	**(f)**	I_2Cl_6	diiodine hexachloride

60. (E)

(a)	ICl	iodine monochloride	**(b)**	ClF_3	chlorine trifluoride
(c)	SiF_4	silicon tetrafluoride	**(d)**	PF_5	phosphorus pentafluoride
(e)	NO_2	nitrogen dioxide	**(f)**	S_4N_4	tetrasulfur tetranitride

61. (E)

(a)	$Al_2(SO_4)_3$	aluminum sulfate	**(b)**	$(NH_4)_2Cr_2O_7$	ammonium dichromate
(c)	SiF_4	silicon tetrafluoride	**(d)**	Fe_2O_3	iron(III) oxide
(e)	C_3S_2	tricarbon disulfide	**(f)**	$Co(NO_3)_2$	cobalt(II) nitrate
(g)	$Sr(NO_2)_2$	strontium nitrite	**(h)**	HBr(aq)	hydrobromic acid
(i)	HIO_3	iodic acid	**(j)**	PCl_2F_3	phosphorus dichloride trifluoride

62. (E)

(a)	$Mg(ClO_4)_2$	magnesium perchlorate	**(b)**	$Pb(C_2H_3O_2)_2$	lead(II) acetate
(c)	SnO_2	tin(IV) oxide	**(d)**	HI(aq)	hydroiodic acid
(e)	$HClO_2$	chlorous acid	**(f)**	$NaHSO_3$	sodium hydrogen sulfite
(g)	$Ca(H_2PO_4)_2$	calcium dihydrogen phosphate	**(h)**	$AlPO_4$	aluminum phosphate
(i)	N_2O_4	dinitrogen tetroxide	**(j)**	S_2Cl_2	disulfur dichloride

63. (E)

(a) Ti^{4+} and Cl^{-} produce $TiCl_4$ **(b)** Fe^{3+} and SO_4^{2-} produce $Fe_2(SO_4)_3$

(c) Cl^{7+} and O^{2-} produce Cl_2O_7 **(d)** S^{7+} and O^{2-} produce $S_2O_8^{2-}$

64. (E)

(a) N^{5+} and O^{2-} produce N_2O_5 **(b)** N^{3+}, O^{2-} and H^{+} produce HNO_2

(c) $C^{+4/3}$ and O^{2-} produce C_3O_2 **(d)** $S^{+2.5}$ and O^{2-} produce $S_4O_6^{2-}$

65. (E)

(a)	$HClO_2$	chlorous acid	**(b)**	H_2SO_3	sulfurous acid
(c)	H_2Se	hydroselenic acid	**(d)**	HNO_2	nitrous acid

66. (E)

(a)	HF(aq)	hydrofluoric acid	**(b)**	HNO_3	nitric acid
(c)	H_3PO_3	phosphorous acid	**(d)**	H_2SO_4	sulfuric acid

67. (E)

(a) OF_2 oxygen difluoride **(b)** XeF_2 xenon difluoride

(c) $CuSO_3$ copper(II) sulfite **(d)** $(NH_4)_2HPO_4$ ammonium hydrogen phosphate

Both (c) and (d) are ionic compounds.

68. (E)

(a) KNO_2 potassium nitrite **(b)** BrF_3 bromine trifluoride

(c) S_2Cl_2 disulfur dichloride **(d)** $Mg(ClO)_2$ magnesium hypochlorite

(e) Cl_2O dichlorine monoxide

Both (a) and (d) are ionic compounds.

Hydrates

69. (E) The hydrate with the greatest mass percent H_2O is the one that gives the largest result for the number of moles of water in the hydrate's empirical formula, divided by the mass of one mole of the anhydrous salt for the hydrate.

$$\frac{5\ H_2O}{CuSO_4} = \frac{5\text{ mol }H_2O}{159.6\text{ g}} = 0.03133 \qquad \frac{6\ H_2O}{MgCl_2} = \frac{6\text{ mol }H_2O}{95.2\text{ g}} = 0.0630$$

$$\frac{18\ H_2O}{Cr_2(SO_4)_3} = \frac{18\text{ mol }H_2O}{392.3\text{ g}} = 0.04588 \qquad \frac{2\ H_2O}{LiC_2H_3O_2} = \frac{2\text{ mol }H_2O}{66.0\text{ g}} = 0.0303$$

The hydrate with the greatest % H_2O therefore is $MgCl_2 \cdot 6\ H_2O$

70. (E) A mole of this hydrate will contain about the same mass of H_2O and of Na_2SO_3.

molar mass $Na_2SO_3 = (2\times 23.0\text{ g Na}) + 32.1\text{ g S} + (3\times 16.0\text{ g O}) = 126.1\text{ g/mol}$

$$\text{number of mol }H_2O = 126.1\text{ g}\times\frac{1\text{ mol }H_2O}{18.0\text{ g }H_2O} = 7.01\text{ mol }H_2O$$

Thus, the formula of the hydrate is $Na_2SO_3 \cdot 7\ H_2O$.

71. (M)

molar mass $CuSO_4 = 63.546\text{ g Cu} + 32.066\text{ g S} + (4\times 15.9994\text{ g O}) = 159.61\text{ g }CuSO_4\text{/mol}$.

Note that each $CuSO_4$ will pick up 5 equivalents of H_2O to give $CuSO_4 \cdot 5H_2O$.

Conversion pathway approach:

$$\text{mass of required }CuSO_4 = 12.6\text{ g }H_2O\times\frac{1\text{ mol }H_2O}{18.015\text{ g }H_2O}\times\frac{1\text{ mol }CuSO_4}{5\text{ mol }H_2O}\times\frac{159.61\text{ g }CuSO_4}{1\text{ mol }CuSO_4}$$

$= 22.3\text{ g }CuSO_4$ is the minimum amount required to remove all the water

Stepwise approach:

$$12.6\text{ g H}_2\text{O}\times\frac{1\text{ mol H}_2\text{O}}{18.0153\text{ g H}_2\text{O}}=0.699\text{ mol H}_2\text{O}\times\frac{1\text{ mol CuSO}_4}{5\text{ mol H}_2\text{O}}=0.140\text{ mol CuSO}_4$$

$$0.140\text{ mol CuSO}_4\times\frac{159.60\text{ g CuSO}_4}{1\text{ mol CuSO}_4}=22.3\text{ g CuSO}_4$$

= is the minimum amount required to remove all the water

72. **(M)** The increase in mass of the solid is the result of each mole of the solid absorbing 10 moles of water.

$$\text{increase in mass}=36.15\text{ g Na}_2\text{SO}_4\times\frac{1\text{ mol Na}_2\text{SO}_4}{142.036\text{ g Na}_2\text{SO}_4}\times\frac{10\text{ mol H}_2\text{O added}}{1\text{ mol Na}_2\text{SO}_4}\times\frac{18.0153\text{ g H}_2\text{O}}{1\text{ mol H}_2\text{O}}$$

$$=45.85\text{ g H}_2\text{O added}$$

73. **(M)** We start by converting to molar amounts for each element based on 100.0 g:

$$20.3\text{ g Cu}\times\frac{1\text{ mol Cu}}{63.546\text{ g Cu}}=0.319\text{ mol Cu}\quad\div 0.319\ \rightarrow\ 1.00\text{ mol Cu}$$

$$8.95\text{ g Si}\times\frac{1\text{ mol Si}}{28.085\text{ g Si}}=0.319\text{ mol Si}\quad\div 0.319\ \rightarrow\ 1.00\text{ mol Si}$$

$$36.3\text{ g F}\times\frac{1\text{ mol F}}{18.998\text{ g F}}=1.91\text{ mol F}\quad\div 0.319\ \rightarrow\ 5.99\text{ mol F}$$

$$34.5\text{ g H}_2\text{O}\times\frac{1\text{ mol H}_2\text{O}}{18.015\text{ g}}=1.91\underline{5}\text{ mol H}_2\text{O}\quad\div 0.319\ \rightarrow\ 6.00\text{ mol H}_2\text{O}$$

Thus the empirical formula for the hydrate is $CuSiF_6 \cdot 6\ H_2O$.

74. **(M)** Let's start by looking at the data provided.

mass of anhydrous compound = 3.967 g

mass of water = 8.129 g – 3.967 g = 4.162 g

$$\text{moles of anhydrous compound}=3.967\text{ g MgSO}_4\times\frac{1\text{ mol MgSO}_4}{120.37\text{ g}}=0.03296\text{ mol}$$

$$\text{moles of H}_2\text{O}=4.162\text{ g}\times\frac{1\text{ mol H}_2\text{O}}{18.015\text{ g H}_2\text{O}}=0.2310\text{ mol H}_2\text{O}$$

$$\text{setting up proportions}\quad\frac{0.2310\text{ mol H}_2\text{O}}{0.03296\text{ mol anhydrous compound}}=\frac{x\text{ mol H}_2\text{O}}{1.00\text{ mol anhydrous compound}}$$

$x = 7.009$ Thus, the formula of the hydrate is $MgSO_4 \cdot 7H_2O$.

Organic Compounds and Organic Nomenclature

75. **(E)** Answer is (b), butan-2-ol is the most appropriate name for this molecule. It has a four carbon atom chain with a hydroxyl group on the carbon second from the end.

76. **(E)** Answer (c), butanoic acid is the most appropriate name for this molecule. It has a four carbon atom chain with an acid group on the first carbon (terminal carbon atom)

77. **(E)** Molecules (a), (b), (c), and (d) are structural isomers. They share a common formula, namely $C_5H_{12}O$, but have different molecular structures. Molecule (e) has a different chemical formula ($C_6H_{14}O$) and hence cannot be classified as an isomer. It should be pointed out that molecules (a) and (c) are identical as well as being isomers of (b).

78. **(E)** Molecules (a), (b), and (c) are structural isomers. They share a common formula, namely $C_5H_{11}Cl$, but have different molecular structures. Molecule (d) has a different chemical formula ($C_6H_{13}Cl$) and hence cannot be classified as an isomer.

79. **(E)**

(a) $CH_3(CH_2)_5CH_3$ **(b)** $CH_3CH_2CO_2H$

(c) $CH_3CH_2CH_2CH(CH_3)CH_2OH$ **(d)** CH_3CH_2F

80. **(E)**

(a) $CH_3(CH_2)_6CH_3$ **(b)** $CH_3(CH_2)_5CO_2H$

(c) $CH_3(CH_2)_2CH(OH)CH_2CH_3$ **(d)** $CH_3CHClCH_2CH_3$

81. **(M)**

(a)	methanol; CH_3OH	Molecular mass = 32.04 u
(b)	2-chlorohexane; $CH_3(CH_2)_3CHClCH_3$	Molecular mass = 120.6 u
(c)	pentanoic acid; $CH_3(CH_2)_3CO_2H$	Molecular mass = 102.1 u
(d)	2-methylprooan-1-ol $CH_3CH(CH_3)CH_2OH$	Molecular mass = 74.12 u

82. **(M)**

(a)	pentan-2-ol; $CH_3CH_2CH_2CH(OH)CH_3$	Molecular mass = 88.15 u
(b)	propanoic acid; $CH_3CH_2CO_2H$	Molecular mass = 74.08 u
(c)	1-bromobutane; $CH_3(CH_2)_2CH_2Br$	Molecular mass = 137.0 u
(d)	3-chlorobutanoic acid; $CH_3CHClCH_2CO_2H$	Molecular mass = 122.6 u

INTEGRATIVE AND ADVANCED EXERCISES

83. (M)

molar mass = $(1\times 6.94 \text{ g Li}) + (1\times 26.982 \text{ g Al}) + (2\times 28.05 \text{ g Si}) + (6\times 15.999 \text{ g O}) = 186.02 \text{ g/mol}$

Conversion pathway approch:

$$\text{number of Li}-6\text{ atoms} = 518 \text{ g spodumene} \times \frac{1 \text{ mol spodumene}}{186.02 \text{ g spodumene}} \times \frac{1 \text{ mol Li}}{1 \text{ mol spodumene}} \times \frac{7.40 \text{ mol Li-6}}{100.00 \text{ mol total Li}}$$

$$\times \frac{6.022\times 10^{23} \text{ Li-6 atoms}}{1 \text{ mol Li-6}} = 1.24\times 10^{23} \text{ Li-6 atoms}$$

Stepwise approch:

$$518 \text{ g spodumene } \times \frac{1 \text{ mol spodumene}}{186.02 \text{ g spodumene}} = 2.78 \text{ mol spodumene}$$

$$2.78 \text{ mol spodumene} \times \frac{1 \text{ mol Li}}{1 \text{ mol spodumene}} = 2.78 \text{ mol Li}$$

$$2.78 \text{ mol Li} \times \frac{7.40 \text{ mol Li-6}}{100.00 \text{ mol total Li}} = 0.206 \text{ mol Li-6}$$

$$0.206 \text{ mol Li-6} \times \frac{6.022\times 10^{23} \text{ Li-6 atoms}}{1 \text{ mol Li-6}} = 1.24\times 10^{23} \text{ Li-6 atoms}$$

84. (M) Determine the mass of each element in the sample.

$$\text{mass Sn} = 0.245 \text{ g SnO}_2 \times \frac{1 \text{ mol SnO}_2}{150.71 \text{ g SnO}_2} \times \frac{1 \text{ mol Sn}}{1 \text{ mol SnO}_2} \times \frac{118.71 \text{ g Sn}}{1 \text{ mol Sn}} = 0.193 \text{ g Sn}$$

$$\text{mass Pb} = 0.115 \text{ g PbSO}_4 \times \frac{1 \text{ mol PbSO}_4}{303.26 \text{ g PbSO}_4} \times \frac{1 \text{ mol Pb}}{1 \text{ mol PbSO}_4} \times \frac{207.2 \text{ g Pb}}{1 \text{ mol Pb}} = 0.0786 \text{ g Pb}$$

$$\text{mass Zn} = 0.246 \text{ g Zn}_2\text{P}_2\text{O}_7 \times \frac{1 \text{ mol Zn}_2\text{P}_2\text{O}_7}{304.72 \text{ g Zn}_2\text{P}_2\text{O}_7} \times \frac{2 \text{ mol Zn}}{1 \text{ mol Zn}_2\text{P}_2\text{O}_7} \times \frac{65.38 \text{ g Zn}}{1 \text{ mol Zn}} = 0.106 \text{ g Zn}$$

Then determine the % of each element in the sample.

$$\%\text{ Sn} = \frac{0.193 \text{ g Sn}}{1.1713 \text{ g brass}} \times 100\% = 16.5\% \text{ Sn} \quad \%\text{ Pb} = \frac{0.0786 \text{ g Pb}}{1.1713 \text{ g brass}} \times 100\% = 6.71\% \text{ Pb}$$

$$\%\text{ Zn} = \frac{0.106 \text{ g Zn}}{1.1713 \text{ g brass}} \times 100\% = 9.05\% \text{ Zn}$$

The % Cu is found by difference.

$$\%\text{Cu} = 100\% - 16.5\% \text{ Sn} - 6.71\% \text{ Pb} - 9.05\% \text{ Zn} = 67.7\% \text{ Cu}$$

85. **(M)** 1 lb = 16 oz = 453.59237 g or 1 oz = 28.35 g

$$3.50 \text{ oz meat} \times \frac{28.35 \text{ g meat}}{1 \text{ oz meat}} \times \frac{0.12 \text{ g } C_6H_5COONa}{100 \text{ g meat}} \times \frac{22.990 \text{ g Na}}{144.105 \text{ g } C_6H_5COONa} \times \frac{1000 \text{ mg Na}}{1 \text{ g Na}}$$
$$= 19.0 \text{ mg Na}$$

<u>86.</u> **(M)** First, we determine the amount of each mineral necessary to obtain 1 kg or 1000 g of boron.

$$1000 \text{ g B} \times \frac{1 \text{ mol B}}{10.81 \text{ g B}} \times \frac{1 \text{ mol } Na_2B_4O_7 \cdot 4\,H_2O}{4 \text{ mol B}} \times \frac{273.28 \text{ g } Na_2B_4O_7 \cdot 4\,H_2O}{1 \text{ mol } Na_2B_4O_7 \cdot 4\,H_2O}$$
$$= 6{,}319.\underline{5} \text{ g } Na_2B_4O_7 \cdot 4\,H_2O$$

$$1000 \text{ g B} \times \frac{1 \text{ mol B}}{10.81 \text{ g B}} \times \frac{1 \text{ mol } Na_2B_4O_7 \cdot 4\,H_2O}{4 \text{ mol B}} \times \frac{381.372 \text{ g } Na_2B_4O_7 \cdot 10\,H_2O}{1 \text{ mol } Na_2B_4O_7 \cdot 4\,H_2O}$$
$$= 8{,}819.\underline{1} \text{ g } Na_2B_4O_7 \cdot 10\,H_2O$$

The difference between these two masses is the required additional mass. Hence,

8819.<u>1</u> g – 6319.<u>5</u> g = 2499.<u>6</u> g. Thus, an additional 2.500 kg mass is required.

87. **(M)** $N_A = \dfrac{9.64853415 \times 10^4 \text{ C}}{1 \text{ mol Ag}} \times \dfrac{1 \text{mol Ag}}{1 \text{mol e}^-} \times \dfrac{1 \text{ e}^-}{1.602176462 \times 10^{-19} \text{C}} = \dfrac{6.0221422 \times 10^{23} \text{ e}^-}{\text{mole e}^-}$

88. **(M)** First, determine the formula of the compound. The compound is 26.58% K, 35.45% Cr and 37.97% O. Assuming 100 g of compound, 26.58 g are potassium, 35.45 g are chromium, and 37.97 g are oxygen.

$$26.58 \text{ g K} \times \frac{1 \text{ mol K}}{39.10 \text{ g K}} = 0.6798 \text{ mol K} \quad \div 0.6798 \text{ mol} \rightarrow 1.000 \text{ mol K}$$

$$35.45 \text{ g Cr} \times \frac{1 \text{ mol Cr}}{52.00 \text{ g Cr}} = 0.6818 \text{ mol Cr} \quad \div 0.6798 \text{ mol} \rightarrow 1.000 \text{ mol Cr}$$

$$37.97 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 2.373 \text{ mol O} \quad \div 0.6798 \text{ mol} \rightarrow 3.491 \text{ mol O}$$

$2 \times KCrO_{3.5} = K_2Cr_2O_7$ which is the formula of the compound

The oxidation state of Cr is +6 and the compound $K_2Cr_2O_7$ is named potassium dichromate.

89. **(M)** It is not possible to have less than 1 molecule of S_8. In order to determine whether it is possible to have 1.00×10^{-23} g of S_8, determine how many molecules that number is equivalent to.

$$1.00\times10^{-23}\text{ g }S_8\times\frac{1\text{ mol }S_8}{256.48\text{ g }S_8}\times\frac{6.022\times10^{23}\text{ molecules }S_8}{1\text{ mol }S_8}=1.0235\text{ molecules }S_8$$

Therefore it is **not** possible to have 1.00×10^{-23} g of S_8.

$$1\text{ molecule }S_8\times\frac{1\text{ mol }S_8}{6.022\times10^{23}\text{ molecules }S_8}\times\frac{256.48\text{ g S}}{1\text{ mol }S_8}=4.26\times10^{-22}\text{ g S}$$

$$4.26\times10^{-22}\text{ g S}\times\frac{1\text{ yg S}}{10^{-24}\text{ g S}}=426\text{ yoctograms S}$$

90. **(E)** A hydrocarbon with a double bond has the molecular formula C_nH_{2n}.

While the moles of H_2O will be greater than the moles of CO_2 as with any other hydrocarbon, the mass of H_2O will never exceed that of CO_2. To prove this, we consider the fact that a hydrocarbon with the greatest ratio of the amount of hydrogen to the amount of carbon will produce the greatest mass of H_2O per gram of CO_2. This hydrocarbon is CH_4, and it produces 2 mol H_2O for every mole of CO_2. From this information, we can determine the maximum ratio of mass H_2O/mass CO_2.

$$\frac{\text{mass }H_2O}{\text{mass }CO_2}=\frac{2\text{ mol }H_2O}{1\text{ mol }CO_2}\times\frac{18.02\text{ g }H_2O}{1\text{ mol }H_2O}\times\frac{1\text{ mol }CO_2}{44.01\text{ g }CO_2}=0.8189\text{ g }H_2O/\text{g }CO_2$$

Thus no hydrocarbon exists that yields a greater mass of H_2O than of CO_2.

91. **(M)** We determine the masses of CO_2 and H_2O produced by burning the C_3H_8.

$$\text{mass}_{CO_2}=6.00\text{ g }C_3H_8\times\frac{1\text{ mol }C_3H_8}{44.097\text{ g }C_3H_8}\times\frac{3\text{ mol C}}{1\text{ mol }C_3H_8}\times\frac{1\text{ mol }CO_2}{1\text{ mol C}}\times\frac{44.01\text{ g }CO_2}{1\text{ mol }CO_2}$$
$$=17.9\underline{6}\text{ g }CO_2$$

$$\text{mass}_{H_2O}=6.00\text{ g }C_3H_8\times\frac{1\text{ mol }C_3H_8}{44.097\text{ g }C_3H_8}\times\frac{8\text{ mol H}}{1\text{ mol }C_3H_8}\times\frac{1\text{ mol }H_2O}{2\text{ mol H}}\times\frac{18.015\text{ g }H_2O}{1\text{ mol }H_2O}$$
$$=9.80\underline{5}\text{ g }H_2O$$

Then, from the masses of CO_2 and H_2O in the unknown compound, we determine the amounts of C and H in that compound and finally its empirical formula.

$$\text{amount C}=(29.0-17.9\underline{6})\text{ g }CO_2\times\frac{1\text{ mol }CO_2}{44.01\text{ g }CO_2}\times\frac{1\text{ mol C}}{1\text{ mol }CO_2}=0.251\text{ mol C}$$

$$\text{amount H}=(18.8-9.80\underline{5})\text{ g }H_2O\times\frac{1\text{ mol }H_2O}{18.015\text{ g }H_2O}\times\frac{2\text{ mol H}}{1\text{ mol }H_2O}=0.998\underline{6}\text{ mol H}$$

The empirical formula of the unknown compound is CH_4. The C:H ratio is 0.9986/0.251 = 3.98. The molecular formula can be calculated by knowing that we have 0.251 moles, which accounts for the 4.00 g of hydrocarbon (40% of 10.0 g). This gives a molar mass of 4.00 ÷ 0.251 = 15.9 g/mol. This is nearly the same as the molar mass of the empirical formula CH_4 (16.04 g/mol).

92. **(M)**

(a) We determine the mass of CO_2 produced from the mixture, with *x* representing the mass of CH_4, and then solve for *x*.

$$n_{\text{carbon}} = x\text{ g CH}_4 \times \frac{1\text{ mol CH}_4}{16.043\text{ g CH}_4} \times \frac{1\text{ mol C}}{1\text{ mol CH}_4} + (0.732 - x)\text{g C}_2\text{H}_6 \times \frac{1\text{ mol C}_2\text{H}_6}{30.070\text{ g C}_2\text{H}_6} \times \frac{2\text{ mol C}}{1\text{ mol C}_2\text{H}_6}$$

$$2.064\text{ CO}_2 = \left(\frac{x}{16.043} + \frac{2(0.732-x)}{30.070}\right)\text{mol C} \times \frac{1\text{mol CO}_2}{1\text{ mol C}} \times \frac{44.009\text{ g CO}_2}{1\text{ mol CO}_2}$$

$$2.064\text{ CO}_2 = 2.7433x + 2.142 - 2.9272x = -0.1839x + 2.142$$

$$x = \frac{2.142 - 2.064}{0.1839} = 0.4\underline{2}\text{ g CH}_4$$

$$\%\text{CH}_4 = \frac{0.4\underline{2}\text{ g CH}_4}{0.732\text{ g mixture}} \times 100\% = 57\%\text{ CH}_4 \approx 60\%\text{ CH}_4 \text{ and } 4\underline{3}\%\text{ C}_2\text{H}_6 \sim 40\%\text{ C}_2\text{H}_6$$

(b) In 100 g of mixture there are the following amounts.

$$\text{amount CH}_4 = 5\underline{7}\text{ g CH}_4 \times \frac{1\text{ mol CH}_4}{16.043\text{ g CH}_4} = 3.\underline{6}\text{ mol CH}_4$$

$$\text{amount C}_2\text{H}_6 = 4\underline{3}\text{ g C}_2\text{H}_6 \times \frac{1\text{ mol C}_2\text{H}_6}{30.070\text{ g C}_2\text{H}_6} = 1.\underline{4}\text{ mol C}_2\text{H}_6$$

$$\text{mol \% CH}_4 = \frac{3\underline{6}\text{ mol CH}_4}{5\underline{0}\text{ mol total}} \times 100\% = 7\underline{2}\text{ mol\% CH}_4 \text{ and } 2\underline{8}\text{ mol\% C}_2\text{H}_6$$

<u>93</u>. **(M)** Since the compound is composed of H_2SO_4 and H_2O, we will need to determine the percent composition of both H_2SO_4 and water.

$$\%\text{ H}_2\text{SO}_4 = \frac{\#\text{ grams H}_2\text{SO}_4}{\text{total mass in grams}} \times 100\%$$

$$\#\text{ g H}_2\text{SO}_4 = 65.2\text{ g (NH}_4)_2\text{SO}_4 \times \frac{1\text{ mol (NH}_4)_2\text{SO}_4}{132.15\text{ g (NH}_4)_2\text{SO}_4} \times \frac{1\text{ mol H}_2\text{SO}_4}{1\text{ mol (NH}_4)_2\text{SO}_4} \times \frac{98.08\text{ g H}_2\text{SO}_4}{1\text{ mol H}_2\text{SO}_4} = 48.4\text{ g}$$

$$\text{total mass} = 32.0\text{ mL mixture} \times \frac{1.78\text{ g mixture}}{1\text{ mL mixture}} = 57.0\text{ g mixture}$$

$$\%\text{ H}_2\text{SO}_4 = \frac{48.4\text{ g}}{57.0\text{ g}} \times 100\% = 85.0\%$$

$\%\ H_2O = 100.0 - 85.0 = 15.0\%$

94. (E) $\% \text{ Ag} = \dfrac{\text{mass Ag}}{\text{mass sample}} \times 100\%$

Using 35.446 g/mol as the molar mass of Cl and 107.87 g/mol as the molar mass of Ag, we find that the molar mass of AgCl is (107.87 + 35.446) g/mol = 143.316 g/mol = 143.32 g/mol. (The number of decimal places in this result is determined by the number of decimal places in the molar mass of Ag.) Using 143.32 g/mol for the molar mass of AgCl, we have

$$31.56 \text{ g AgCl} \times \frac{1 \text{ mol AgCl}}{143.32 \text{ g AgCl}} \times \frac{1 \text{ mol Ag}}{1 \text{ mol AgCl}} \times \frac{107.87 \text{ g Ag}}{1 \text{ mol Ag}} = 23.75 \text{ g Ag}$$

$$\% \text{ Ag} = \frac{23.75 \text{ g}}{26.39 \text{ g}} \times 100\% = 90.00\%$$

Now, using 35.457 g/mol as the molar mass of Cl, we find that the molar mass of AgCl is (107.87 + 35.457) g/mol = 143.327 g/mol = 143.33 g/mol and

$$31.56 \text{ g AgCl} \times \frac{1 \text{ mol AgCl}}{143.33 \text{ g AgCl}} \times \frac{1 \text{ mol Ag}}{1 \text{ mol AgCl}} \times \frac{107.87 \text{ g Ag}}{1 \text{ mol Ag}} = 23.75 \text{ g Ag}$$

We see from this calculation that the mass of Ag obtained is the same (to four significant figures). Therefore, the corresponding value for the % Ag will also be 90.00%.

In this case, the fact that the atomic mass of Cl is given as an interval has little effect on the calculated final result.

<u>95.</u> (D)

$$9.0 \times 10^{-4} \frac{\mu \text{mol } C_2H_6S}{\text{m}^3 \text{ air}} \times \frac{1 \times 10^{-6} \text{ mol } C_2H_6S}{1 \, \mu \text{mol } C_2H_6S} \times \frac{62.13 \text{ g } C_2H_6S}{1 \text{ mol } C_2H_6S} \times \frac{1 \text{ m}^3}{(100)^3 \text{cm}^3} \times$$

$$\frac{1 \text{ cm}^3}{1 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} \times \frac{1 \text{ L air}}{1.2 \text{ g air}} = 4.7 \times 10^{-11} \frac{\text{g } C_2H_6S}{\text{g air}}$$

$$4.7 \times 10^{-11} \frac{\text{g } C_2H_6S}{\text{g air}} \times \frac{1 \times 10^9 \text{ g}}{1 \text{ billion grams}} = 0.0466 \text{ ppb} = 0.05 \text{ ppb}$$

<u>96.</u> (D)

(a) If we have one mole of entities, then we must have 0.7808 mol N_2, 0.2095 mol O_2, 0.0093 mol Ar, and 0.0004 mol CO_2.

$$0.7808 \text{ mol } N_2 \times \frac{28.02 \text{ g } N_2}{1 \text{ mol } N_2} = 21.88 \text{ g } N_2$$

$$0.2095 \text{ mol } O_2 \times \frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2} = 6.704 \text{ g } O_2$$

$$0.0004 \text{ mol } CO_2 \times \frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2} = 0.0176 \text{ g } CO_2$$

$$0.0093 \text{ mol Ar} \times \frac{39.948 \text{ g Ar}}{1 \text{ mol Ar}} = 0.3715 \text{ g Ar}$$

mass of air sample = 21.88g N_2 + 6.704 g O_2 + 0.0176 g CO_2 + 0.3715 g Ar = 28.97 g

(b) $1\text{ m}^3 \times \frac{(100)^3\text{ cm}^3}{1\text{ m}^3} \times \frac{1\text{ mL}}{1\text{ cm}^3} \times \frac{1\text{ L}}{1000\text{ mL}} \times \frac{1.2\text{ g}}{1\text{ L}} = 1200\text{ g dry air}$

$$1200\text{ g} \times \frac{1\text{ mol entities}}{28.97\text{ g}} \times \frac{1.14 \times 10^{-4}}{100} = 4.72 \times 10^{-5}\text{ mol Kr}$$

$$4.72 \times 10^{-5}\text{ mol Kr} \times \frac{83.80\text{ g}}{1\text{ mol}} = 3.96 \times 10^{-3}\text{ g Kr} = 4.0\text{ mg Kr}$$

97. **(M)**

(a) Make the assumption that 1 mL H_2O = 1 g H_2O.

$$350\text{ g H}_2\text{O} \times \frac{0.8\text{ g CHCl}_3}{1{,}000{,}000{,}000\text{ g H}_2\text{O}} \times \frac{1\text{ mol CHCl}_3}{119.377\text{ g CHCl}_3} \times \frac{6.022 \times 10^{23}\text{ CHCl}_3}{1\text{ mol CHCl}_3} = 1.41 \times 10^{15}\text{ CHCl}_3$$

(b) $1.14 \times 10^{15}\text{ CHCl}_3 \times \frac{1\text{ mol CHCl}_3}{6.022 \times 10^{23}\text{ CHCl}_3} \times \frac{119.377\text{ g CHCl}_3}{1\text{ mol CHCl}_3} = 2.25 \times 10^{-7}\text{ g CHCl}_3$

Alternatively, $350\text{ g H}_2\text{O} \times \frac{1\text{ g CHCl}_3}{1{,}000{,}000{,}000\text{ g H}_2\text{O}} = 3.50 \times 10^{-7}\text{ g CHCl}_3$

This amount would not be detected with an ordinary analytical balance. You would require something that was at least 300 times more sensitive.

98. **(M).**We can determine both the number of moles of M and the mass of M in 0.1131 g MSO_4. Their quotient is the atomic mass of M.

$$\text{mol M}^{2+} = 0.2193\text{ g BaSO}_4 \times \frac{1\text{ mol BaSO}_4}{233.39\text{ g BaSO}_4} \times \frac{1\text{ mol SO}_4^{2-}}{1\text{ mol BaSO}_4} \times \frac{1\text{ mol M}^{2+}}{1\text{ mol SO}_4^{2-}} = 0.0009396\text{ mol M}^{2+}$$

$$\text{mass SO}_4^{2-} = 0.0009396\text{ mol M}^{2+} \times \frac{1\text{ mol SO}_4^{2-}}{1\text{ mol M}^{2+}} \times \frac{96.064\text{ g SO}_4^{2-}}{1\text{ mol SO}_4^{2-}} = 0.09026\text{ g SO}_4^{2-}$$

$$\text{mass M} = \text{mass MSO}_4 - \text{mass SO}_4^{2-} = 0.1131\text{ g MSO}_4 - 0.09026\text{ g SO}_4^{2-} = 0.0228\text{ g M}$$

$$\text{atomic mass M} = \frac{\text{mass M}}{\text{moles M}} = \frac{0.0228\text{ g M}}{0.0009396\text{ mol M}} = 24.3\text{ g M/mol}$$

M is the element magnesium.

<u>99</u>. **(M)** $\text{mass SO}_4^{2-} = 1.511\text{ g BaSO}_4 \times \frac{1\text{ mol BaSO}_4}{233.39\text{ g BaSO}_4} \times \frac{1\text{ mol SO}_4^{2-}}{1\text{ mol BaSO}_4} = 0.006474\text{ mol SO}_4^{2-}$

$$0.006474\text{ mol SO}_4^{2-} \times \frac{96.064\text{ g SO}_4^{2-}}{1\text{ mol SO}_4^{2-}} = 0.6219\text{ g SO}_4^{2-}$$

$$\text{amount M} = 0.006474\text{ mol SO}_4^{2-} \times \frac{2\text{ mol M}^{3+}}{3\text{ mol SO}_4^{2-}} = 0.004316\text{ mol M}^{3+}$$

$$\text{mass M} = 0.738\text{ g M}_2(\text{SO}_4)_2 - 0.6219\text{ g SO}_4^{2-} = 0.116\text{ g M}$$

$\text{atomic mass of M} = \frac{0.116\text{ g M}}{0.004316\text{ mol M}} = 26.9\text{ g M/mol}$ M is the element aluminum.

100. **(D)** Set up an equation in the usual conversion-factor format, to determine what mass of MS can be obtained from the given mass of M_2O_3. Of course, the mass of MS is not unknown; it is 0.685 g. What is unknown is the atomic mass of the element M; let's call this x and solve for x.

$$0.685 \text{ g MS} = 0.622 \text{ g } M_2O_3 \times \frac{1 \text{ mol } M_2O_3}{[2x+(3\times16.0)] \text{ g } M_2O_3} \times \frac{2 \text{ mol M}}{1 \text{ mol } M_2O_3} \times \frac{1 \text{ mol MS}}{1 \text{ mol M}} \times \frac{(x+32.1) \text{ g MS}}{1 \text{ mol MS}}$$

$$0.685 \text{ g MS} = \frac{0.622\times2\times(x+32.1)}{2x+48.0} \qquad 1.24\underline{4}x+39.9 = 1.37x+32.9 \quad 0.126x = 7.0$$

$x = 56$ is the atomic mass of M M is element 56, iron (Fe)

101. **(D)** Let $x = \text{mol } MgCl_2$ in the sample, and $y = \text{mol NaCl}$. Set up two equations.

$$0.6110 = \frac{2x\times35.45 \text{ g Cl}+35.45y \text{ g Cl}}{0.5200 \text{ g sample}} \qquad 0.5200 \text{ g} = 95.205x \text{ g } MgCl_2 + 58.44\,y \text{ g NaCl}$$

We then solve these two equations for x.

$$0.6110\times0.5200 = 0.3177 = 70.9x+35.45y \quad \frac{0.3177}{35.45} = 0.008962 = 2x+y$$

$$y = 0.008962-2x$$

$$0.5200 = 95.205x+58.44(0.008962-2x)$$

$$= 95.205x+0.5237-116.88x$$

$$0.5200-0.5237 = -0.0037 = (95.205-116.88)x = -21.675x$$

$$-0.0037 = -21.675x$$

$$x = \frac{0.0037}{21.675}$$

$$x = 1.7\times10^{-4} \text{ mol } MgCl_2$$

Then determine the value of y.

$$y = 0.008962-2x = 0.008962-2\ \times0.00017 = 0.00862 \text{ mol NaCl}$$

$$\text{mass } MgCl_2 = 1.7\times10^{-4} \text{ mol } MgCl_2 \times \frac{95.21 \text{ g } MgCl_2}{1 \text{ mol } MgCl_2} = 0.016 \text{ g } MgCl_2$$

$$\text{mass NaCl} = 0.00862 \text{ mol NaCl}\times\frac{58.44 \text{ g NaCl}}{1 \text{ mol NaCl}} = 0.504 \text{ g NaCl}$$

$$\%MgCl_2 = \frac{0.016 \text{ g } MgCl_2}{0.5200 \text{ g sample}}\times100\% = 3.1\% \text{ } MgCl_2 \qquad \% \text{ NaCl} = \frac{0.504 \text{ g NaCl}}{0.5200 \text{ g sample}}\times100\% = 96.9\% \text{ NaCl}$$

The precision of the calculation is poor because there is only a small % $MgCl_2$ in the mixture. When calculating the number of moles of $MgCl_2$, 0.5237 is subtracted from 0.5200 as shown above. The number of significant figures is reduced by this subtraction, which has the effect of reducing the precision in subsequent calculations.

102. (M) First, we determine the mass of Pb in 2.750 g Pb_3O_4.

$$\text{mass Pb} = 2.750 \text{ g } Pb_3O_4 \times \frac{1 \text{ mol } Pb_3O_4}{685.596 \text{ g } Pb_3O_4} \times \frac{3 \text{ mol Pb}}{1 \text{ mol } Pb_3O_4} \times \frac{207.2 \text{ g Pb}}{1 \text{ mol Pb}}$$

$$= 2.493 \text{ g Pb}$$

Then, we determine the amounts of O and Pb in the second oxide. From these, we determine the empirical formula of the second oxide.

$$\text{amount O} = (2.686 \text{ g} - 2.493 \text{ g}) \text{ O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.0121 \text{ mol O}$$

$$\text{amount Pb} = 2.493 \text{ g Pb} \times \frac{1 \text{ mol Pb}}{207.2 \text{ g Pb}} = 0.0120 \text{ mol Pb}$$

Thus, the empirical formula of the second oxide is PbO.

<u>103</u>. (M) If we determine the mass of anhydrous $ZnSO_4$ in the hydrate, we then can determine the mass of water, and the formula of the hydrate.

$$\text{mass } ZnSO_4 = 0.8223 \text{ g } BaSO_4 \times \frac{1 \text{ mol } BaSO_4}{233.386 \text{ g}} \times \frac{1 \text{ mol } ZnSO_4}{1 \text{ mol } BaSO_4} \times \frac{161.454 \text{ g } ZnSO_4}{1 \text{ mol } ZnSO_4}$$

$$= 0.5688 \text{ g } ZnSO_4$$

The water present in the hydrate is obtained by difference.

$\text{mass } H_2O = 1.013 \text{ g hydrate} - 0.5688 \text{ g } ZnSO_4 = 0.444 \text{ g } H_2O$

The hydrate's formula is determined by a method similar to that for obtaining an empirical formula.

$$\text{amt. } ZnSO_4 = 0.5688 \text{ g} \times \frac{1 \text{ mol } ZnSO_4}{161.436 \text{ mol } ZnSO_4} = 0.003523 \text{ mol } ZnSO_4 \div 0.003523 \longrightarrow 1.00 \text{ mol } ZnSO_4$$

$$\text{amt. } H_2O = 0.444 \text{ g} \times \frac{1 \text{ mol } H_2O}{18.015 \text{ g } H_2O} = 0.0246\underline{5} \text{ mol } H_2O \div 0.003523 \longrightarrow 7.00 \text{ mol } H_2O$$

Thus, the formula of the hydrate is $ZnSO_4 \cdot 7\, H_2O$.

104. (M) $\frac{1.552 \text{ g MI}}{1.186 \text{ g I}} \times \frac{126.904 \text{ g I}}{1 \text{ mol I}} \times \frac{1 \text{ mol I}}{1 \text{ mol MI}} = \frac{166.1 \text{ g MI}}{1 \text{ mol MI}}$ (This is the molar mass of MI.)

Subtract the mass of 1 mol of I to obtain the molar mass of M.

molar mass M = $(166.1 - 126.904) \text{ g mol}^{-1} = 39.2 \text{ g mol}^{-1}$

The cation is probably K ($39.098 \text{ g mol}^{-1}$).

Alternatively, find mass of M in sample: 1.552 g MI – 1.186 g I = 0.366 g M

$$\frac{0.366 \text{ g M}}{1.186 \text{ g I}} \times \frac{126.90 \text{ g I}}{1 \text{ mol I}} \times \frac{1 \text{ mol I}}{1 \text{ mol M}} = 39.2 \text{ g mol}^{-1}$$

105. (M)

$$13 \text{ atoms} \times \frac{15.38 \text{ atoms E}}{100 \text{ atoms in formula unit}} = \frac{1.999 \text{ atoms E}}{\text{formula unit}} \quad \therefore \ H_xE_2O_z \text{ (Note: } x + z = 11)$$

34.80% E by mass, hence, 65.20% H and O by mass

178 u × 0.3480 E = 61.944 u for 2 atoms of E, ∴ E = 30.972 u Probably P (30.974 u)

H and O in formula unit = 178 u – 30.972 u = 116 u

$x + z = 11$ or $x = 11 - z$ and $x(1.008 \text{ u}) + z(15.999 \text{ u}) = 116 \text{ u}$

Substitute and solve for z: $(11 - z)(1.008 \text{ u}) + z(15.999 \text{ u}) = 116 \text{ u}$

11.08734 u – 1.008 u(z) + 15.999 u(z) = 116 u Divide through by u and collect terms

105 = 14.9915(z) or $z = 7$ and $x = 11 - z = 11 - 7 = 4$.

Therefore, the formula is $H_4P_2O_7$ (as a check, 13 atoms and 177.975 u ~ 178 u).

106. (M) First find the mass of carbon, hydrogen, chlorine, and oxygen. From the molar ratios, we determine the molecular formula.

$$2.094 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.009 \text{ g } CO_2} \times \frac{1 \text{ mol C}}{1 \text{ mol } CO_2} = 0.04759 \text{ mol C} \times \frac{12.011 \text{ g C}}{1 \text{ mol C}} = 0.5716 \text{ g C}$$

$$0.286 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.015 \text{ g } H_2O} \times \frac{2 \text{ mol H}}{1 \text{ mol } H_2O} = 0.0317\underline{5} \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 0.0320 \text{ g H}$$

$$\text{moles of chlorine} = \frac{\text{mol C}}{2} = \frac{0.04759}{2} = 0.02380 \text{ mol Cl}$$

$$\text{mass of Cl} = 0.02380 \text{ mol Cl} \times \frac{35.45 \text{ g Cl}}{1 \text{ mol Cl}} = 0.8436 \text{ g Cl}$$

mass of oxygen obtained by difference: 1.510 g – 0.8436 g – 0.5716 g – 0.0320 g = 0.063 g O

$$\text{moles of oxygen} = 0.063 \text{ g O} \times \frac{1 \text{ mol O}}{15.999 \text{ g O}} = 0.0039\underline{4} \text{ mol O}$$

Divide the number of moles of each element by 0.0039$\underline{4}$ to give an empirical formula of $C_{12.\underline{1}}H_{8.0\underline{6}}Cl_{6.0\underline{4}}O_{1.00}$ owing to the fact that the oxygen mass is obtained by difference, and it has only two significant digits and thus a higher degree of uncertainty.

The empirical formula is $C_{12}H_8Cl_6O$, which with a molecular mass of 381 u has the same molecular mass as the molecular formula. Hence, this empirical formula is also the molecular formula.

107. (M) 1.271 g Na_2SO_4 absorbs 0.387 g H_2O

$$\text{mass}_{Na_2SO_4\cdot 10\,H_2O} = 0.387\text{ g }H_2O \times \frac{1\text{ mol }H_2O}{18.015\text{ g }H_2O} \times \frac{1\text{ mol }Na_2SO_4\cdot 10\,H_2O}{10\text{ mol }H_2O} \times \frac{322.186\text{ g }Na_2SO_4\cdot 10\,H_2O}{1\text{ mol }Na_2SO_4\cdot 10\,H_2O}$$

$$= 0.692\text{ g }Na_2SO_4\cdot 10\,H_2O$$

$$\text{mass percent }Na_2SO_4\cdot 10\,H_2O = \frac{0.692\text{ g}}{(1.271\text{ g} + 0.387\text{ g})} \times 100\% = 41.7\%$$

108. (D) Let X = molar mass of Bi and Y = moles of Bi_2O_3

molar mass of $Bi(C_6H_5)_3$ = X + (18 × 12.011 g mol^{-1} + 15 × 1.008 g mol^{-1})
= X + 231.318 g mol^{-1}

molar mass of Bi_2O_3 = 2X + (3 × 15.999 g mol^{-1} = 2X + 47.997 g mol^{-1}

Consider the reaction: 2 $Bi(C_6H_5)_3 \rightarrow Bi_2O_3$

If Y = moles of Bi_2O_3, then 2Y = moles of $Bi(C_6H_5)_3$.

For both the reactant and the product, mass = MM × moles. Therefore,

5.610 g = 2Y(X + 231.318 g mol^{-1}) = 2XY + (2Y) 231.318 g mol^{-1}, and
2.969 g = Y(2X + 47.997 g mol^{-1}) = 2XY + (Y) 47.997 g mol^{-1}

Rearrange 5.610 g = 2XY + (Y) 231.318 g mol^{-1} to 2XY
= 5.610 g – (2Y) 231.318 g mol^{-1}

Substitute for 2XY in 2.969 = 2XY + (Y)47.997 g mol^{-1}
2.969 g = 5.610 g – (2Y) 231.318 g mol^{-1} + (Y) 47.997 g mol^{-1}
= 5.610 g – (Y) 414.639 g mol^{-1}

Collect terms and solve for Y.
5.610 g –2.969 g = (Y) 414.639 g mol^{-1} = 2.641 g
Y = 2.641 g ÷ 414.639 g mol^{-1} = 0.0063694 mol

Substitute Y in 2.969g = 2XY + (Y) 47.997 g mol^{-1} and solve for X, the molar mass of Bi
2.969g = 2X(0.0063694 mol) + (0.0063694 mol) 47.997 g mol^{-1}

$$X = \frac{2.969\text{g} - (0.0063694\text{ mol})\ 47.997\text{ g mol}^{-1}}{(2)0.0063694\text{ mol}} = 209.1\text{g mol}^{-1}\ \text{(Actually it is 208.98 g mol}^{-1}\text{)}$$

109. (D)

$$\text{volume of Au} = 0.25\text{ mm} \times 15\text{ mm} \times 15\text{ mm} = 56.25\text{ mm}^3$$

$$56.25\text{ mm}^3 \times \frac{1\text{ cm}^3}{(10)^3\text{ mm}^3} \times \frac{19.3\text{ g Au}}{\text{cm}^3} = 1.086\text{ g Au}$$

$$1.086\text{ g Au} \times \frac{1\text{ mol Au}}{196.97\text{ g Au}} = 5.51\times10^{-3}\text{ mol Au}$$

$$1.400\text{ g} - 1.086\text{ g} = 0.314\text{ g F} \times \frac{1\text{ mol F}}{18.998\text{ g F}} = 0.0165\text{ mol F}$$

$$\frac{5.51\times10^{-3}}{5.51\times10^{-3}} = 1\text{ mol Au}$$

$$\frac{0.0165}{5.51\times10^{-3}} = 3\text{ mol F}$$

The formula is therefore AuF_3, which is gold(III) fluoride.

110. (M)

Calculate the mass of chlorine: 0.244 L × 2.898 g/L = 0.707 g chlorine

Calculate the mass of iodine: 1.553 g − 0.707 g = 0.846 g iodine

Calculate the moles of chlorine: 0.707 g/35.45 g/mol = 0.0199 mol chlorine

Calculate the moles of iodine: 0.846 g/126.90 g/mol = 0.00667 mol iodine

Calculate the mole ratio: 0.0199:0.00667 = 1:2.98 ≈ 1:3

Calculate the empirical molar mass: (126.90 + 3 × 35.45) g/mol = 233.25 g/mol

Because 467/233.25 ≈ 2, the molecular formula is I_2Cl_6.

111. (M)

The coating is a compound of copper and iodine. The mass of iodine reacting with the copper strip is 0.733 g – 0.725 g = 0.008 g.

The mass of the copper reacted is 0.725 g – 0.721 g = 0.004 g.

The masses of copper and iodine that reacted are known with very low precision (only one significant figure). Therefore, in the following calculations, we are justified in rounding the molar masses of I and Cu (126.90 and 63.546 g/mol, respectively) to two significant figures and the final result to one significant figure.

The amount of iodine is $0.008\text{ g} \times \dfrac{1}{130\text{ g mol}^{-1}} = 6\times10^{-5}\text{ mol}$

The amount of copper is $0.004\text{ g} \times \dfrac{1}{64\text{ g mol}^{-1}} = 6\times10^{-5}\text{ mol}$

The compound contains equal amounts (moles) of Cu and I. Therefore, the empirical formula is CuI.

FEATURE PROBLEMS

112. (D)

(a) "5-10-5" fertilizer contains 5.00 g N (that is, 5.00% N), 10.00 g P_2O_5, and 5.00 g K_2O in 100.00 g fertilizer. We convert the last two numbers into masses of the two elements.

(1) $$\%\text{P} = 10.00\%\ \text{P}_2\text{O}_5 \times \frac{1\text{ mol P}_2\text{O}_5}{141.9\text{ g P}_2\text{O}_5} \times \frac{2\text{ mol P}}{1\text{ mol P}_2\text{O}_5} \times \frac{30.97\text{ g P}}{1\text{ mol P}} = 4.37\%\ \text{P}$$

(2) $$\%\text{K} = 5.00\%\ \text{K}_2\text{O} \times \frac{1\text{ mol K}_2\text{O}}{94.20\text{ g K}_2\text{O}} \times \frac{2\text{ mol K}}{1\text{ mol K}_2\text{O}} \times \frac{39.10\text{ g K}}{1\text{ mol K}} = 4.15\%\ \text{K}$$

(b) First, we determine %P and then convert it to $\% P_2O_5$, given that 10.0% P_2O_5 is equivalent to 4.37% P.

(1) $$\% \ P_2O_5 = \frac{2 \text{ mol P}}{1 \text{ mol } Ca(H_2PO_4)_2} \times \frac{30.97 \text{ g P}}{1 \text{ mol P}} \times \frac{1 \text{ mol } Ca(H_2PO_4)_2}{234.05 \text{ g } Ca(H_2PO_4)_2} \times 100\% \times \frac{10.0\% \ P_2O_5}{4.37\% \ P} = 60.6\% \ P_2O_5$$

(2) $$\% \ P_2O_5 = \frac{1 \text{ mol P}}{1 \text{ mol} (NH_4)_2 HPO_4} \times \frac{30.97 \text{ g P}}{1 \text{ mol P}} \times \frac{1 \text{ mol} (NH_4)_2 HPO_4}{132.06 \text{ g} (NH_4)_2 HPO_4} \times 100\% \times \frac{10.0\% \ P_2O_5}{4.37\% \ P} = 53.7\% \ P_2O_5$$

(c) If the mass ratio of $(NH_4)_2HPO_4$ to KCl is set at 5.00:1.00, then for every 5.00 g of $(NH_4)_2HPO_4$ in the mixture there must be 1.00 g of KCl. Let's start by finding the %N, %P, and %K for the fertilizer mixture.

$$\%N \text{ (by mass)} = \frac{2 \text{ mol N}}{1 \text{ mol} (NH_4)_2 HPO_4} \times \frac{1 \text{ mol} (NH_4)_2 HPO_4}{132.06 \text{ g} (NH_4)_2 PO_4} \times \frac{14.007 \text{ g N}}{1 \text{ mol N}} \times \frac{5.00 \text{ g} (NH_4)_2 HPO_4}{6.00 \text{ g mixture}} \times 100\% = 17.7\% \text{ N}$$

$$\%P \text{ (by mass)} = \frac{1 \text{ mol P}}{1 \text{ mol } (NH_4)_2HPO_4} \times \frac{1 \text{ mol } (NH_4)_2HPO_4}{132.06 \text{ g } (NH_4)_2HPO_4} \times \frac{30.9738 \text{ g P}}{1 \text{ mol P}} \times \frac{5.00 \text{ g } (NH_4)_2HPO_4}{6.00 \text{ g of mixture}} \times 100\% = 19.5\% \text{ P}$$

$$\%K \text{ (by mass)} = \frac{1 \text{ mol K}}{1 \text{ mol KCl}} \times \frac{1 \text{ mol KCl}}{74.55 \text{ g KCl}} \times \frac{39.0983 \text{ g K}}{1 \text{ mol K}} \times \frac{1.00 \text{ g KCl}}{6.00 \text{ g mixture}} \times 100\% = 8.74\% \text{ K}$$

Next, we convert %P to $\%P_2O_5$ and %K to $\%K_2O$.

$$\%P_2O_5 = 19.5\% \text{ P} \times \frac{10.0 \ \%P_2O_5}{4.37 \ \%P} = 44.\underline{6}\% \ P_2O_5$$

$$\%K_2O = 8.\underline{74}\% \text{ K} \times \frac{5.00\% \ K_2O}{4.15\% \text{ K}} = 1\underline{0.5}\% \ K_2O$$

Thus, the combination of 5.00 g $(NH_4)_2HPO_4$ with 1.00 g KCl affords a "17.7-44.6-10.5" fertilizer, that is, 17.7% N, a percentage of phosphorus expressed as 44.6% P_2O_5, and a percentage of potassium expressed as 10.5% K_2O.

(d) A "5-10-5" fertilizer must possess the mass ratio 5.00 g N: 4.37 g P: 4.15 g K per 100 g of fertilizer. Thus a "5-10-5" fertilizer requires an N:P relative mass ratio of 5.00 g N:4.37 g P = 1.00 g N:0.874 g P. Note specifically that the fertilizer has a somewhat *greater* mass of N than of P.

If all of the N and P in the fertilizer comes solely from $(NH_4)_2HPO_4$, then the atom ratio of N relative to P will remain fixed at 2 N:1 P. Whether or not an inert non-fertilizing filler is present in the mix is immaterial. The relative N:P mass ratio is (2×14.01) g N:30.97 g P, that is, 0.905 g N:1.00 g P. Note specifically that $(NH_4)_2HPO_4$ has a somewhat *lesser* mass of N than of P. Clearly, it is impossible to make a "5-10-5" fertilizer if the only fertilizing components are $(NH_4)_2HPO_4$ and KCl.

113. (D)

(a) First, calculate the mass of water that was present in the hydrate prior to heating.
mass of H_2O = 2.574 g $CuSO_4 \cdot x\ H_2O$ – 1.647 g $CuSO_4$ = 0.927 g H_2O
Next, we need to find the number of moles of anhydrous copper(II) sulfate and water that were initially present together in the original hydrate sample.

$$\text{moles of } CuSO_4 = 1.647 \text{ g } CuSO_4 \times \frac{1 \text{ mol } CuSO_4}{159.6 \text{ g } CuSO_4} = 0.01032 \text{ moles } CuSO_4$$

$$\text{moles of } H_2O = 0.927 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.015 \text{ g } H_2O} = 0.0514\underline{6} \text{ moles of water}$$

The empirical formula is obtained by dividing the number of moles of water by the number of moles of $CuSO_4$ (x = ratio of moles of water to moles of $CuSO_4$)

$$x = \frac{0.05146 \text{ moles } H_2O}{0.01032 \text{ moles } CuSO_4} = 4.99 \sim 5$$ The empirical formula is $CuSO_4 \cdot 5\ H_2O$.

(b) mass of water present in hydrate = 2.574 g – 1.833 g = 0.741 g H_2O

$$\text{moles of water} = 0.741 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.015 \text{ g } H_2O} = 0.0411 \text{ moles of water}$$

mass of $CuSO_4$ present in hydrate = 1.833 g $CuSO_4$

$$\text{moles of } CuSO_4 = 1.833 \text{ g } CuSO_4 \times \frac{1 \text{ mol } CuSO_4}{159.602 \text{ g } CuSO_4} = 0.0115 \text{ mol } CuSO_4$$

The empirical formula is obtained by dividing the number of moles of water by the number of moles of $CuSO_4$ (x = ratio of moles of water to moles of $CuSO_4$).

$$x = \frac{0.0411 \text{ moles } H_2O}{0.0115 \text{ moles } CuSO_4} = 3.58 \sim 4.$$

Since the hydrate has not been completely dehydrated, there is no problem with obtaining non-integer “garbage” values.

So, the empirical formula is $CuSO_4 \cdot 4\ H_2O$.

(c) When copper(II) sulfate is strongly heated, it decomposes to give $SO_3(g)$ and $CuO(s)$. The black residue formed at 1000°C in this experiment is probably CuO. The empirical formula for copper(II) oxide is CuO. Let’s calculate the percentages of Cu and O by mass for CuO:

$$\text{mass percent copper} = \frac{63.546 \text{ g Cu}}{79.545 \text{ g CuO}} \times 100\% = 79.89\% \text{ by mass Cu}$$

$$\text{mass percent oxygen} = \frac{15.999 \text{ g O}}{79.545 \text{ g CuO}} \times 100\% = 20.11\% \text{ by mass O}$$

The number of moles of CuO formed (by reheating to 1000 °C)

$= 0.812 \text{ g CuO} \times \dfrac{1 \text{ mol CuO}}{79.545 \text{ g CuO}} = 0.0102 \text{ moles of CuO}$

This is very close to the number of moles of anhydrous $CuSO_4$ formed at 400. °C. Thus, it would appear that upon heating to 1000 °C, the sample of $CuSO_4$ was essentially completely converted to CuO.

114. (D)

(a) The formula for stearic acid, obtained from the molecular model, is $CH_3(CH_2)_{16}CO_2H$. The number of moles of stearic acid in 10.0 grams is

$= 10.0 \text{ g stearic acid} \times \dfrac{1 \text{ mol stearic acid}}{284.48 \text{ g stearic acid}} = 3.51\underline{5} \times 10^{-2} \text{ mol of stearic acid.}$

The layer of stearic acid is one molecule thick. According to the figure provided with the question, each stearic acid molecule has a cross-sectional area of ~0.22 nm^2. In order to find the stearic acid coverage in square meters, we must multiply the total number of stearic acid molecules by the cross-sectional area for an individual stearic acid molecule. The number of stearic acid molecules is:

$= 3.51\underline{5} \times 10^{-2} \text{ mol of stearic acid} \times \dfrac{6.022\times10^{23} \text{ molecules}}{1 \text{ mol of stearic acid}} = 2.11\underline{7} \times 10^{22} \text{ molecules}$

$\text{area in m}^2 = 2.11\underline{7} \times 10^{22} \text{ molecules of stearic acid} \times \dfrac{0.22 \text{ nm}^2}{\text{molecule}} \times \dfrac{(1 \text{ m})^2}{(1\times10^9 \text{ nm})^2}$

The area in m^2 = 4657 m^2 or 4.7×10^3 m^2 (with correct number of sig. fig.)

(b) The density for stearic acid is 0.85 g cm^{-3}. Thus, 0.85 grams of stearic acid occupies 1 cm^3. Find the number of moles of stearic acid in 0.85 g of stearic acid

$= 0.85 \text{ grams of stearic acid} \times \dfrac{1 \text{ mol stearic acid}}{284.48 \text{ g stearic acid}} = 3.0 \times 10^{-3} \text{ mol of stearic}$

acid. This number of moles of acid occupies 1 cm^3 of space. So, the number of stearic acid molecules in 1 cm^3

$= 3.0 \times 10^{-3} \text{ mol of stearic acid} \times \dfrac{6.022\times10^{23} \text{ molecules}}{1 \text{ mol of stearic acid}}$

$= 1.8 \times 10^{21}$ stearic acid molecules.

Thus, the volume for a single stearic acid molecule in nm^3

$= 1 \text{ cm}^3 \times \dfrac{1}{1.8\times10^{21} \text{ molecules stearic acid}} \times \dfrac{(1.0\times10^7 \text{ nm})^3}{(1 \text{ cm})^3} = 0.55\underline{6} \text{ nm}^3$

The volume of a rectangular column is simply the area of its base multiplied by its height (i.e., V = area of base (in nm^2) × height (in nm)).

So, the average height of a stearic acid molecule $= \dfrac{0.556 \text{ nm}^3}{0.22 \text{ nm}^2} = 2.5 \text{ nm}$

(c) The density for oleic acid = 0.895 g mL^{-1}. So, the concentration for oleic acid is

$$= \frac{0.895 \text{ g acid}}{10.00 \text{ mL}} = 0.0895 \text{ g mL}^{-1} \text{ (solution 1)}$$

This solution is then divided by 10, three more times, to give a final concentration of $8.9\underline{5} \times 10^{-5}$ g mL^{-1}. A 0.10 mL sample of this solution contains:

$$= \frac{8.95 \times 10^{-5} \text{g acid}}{1.00 \text{ mL}} \times 0.10 \text{ mL} = 8.9\underline{5} \times 10^{-6} \text{ g of acid.}$$

$$\text{The number of acid molecules} = 85 \text{ cm}^2 \times \frac{1}{4.6 \times 10^{-15} \text{cm}^2 \text{ per molecule}}$$
$$= 1.8\underline{5} \times 10^{16} \text{ oleic acid molecules.}$$

So, 8.95×10^{-6} g of oleic acid corresponds to 1.85×10^{16} oleic acid molecules.

The molar mass for oleic acid, $C_{18}H_{34}O_2$, is 282.47 g mol^{-1}.

The number of moles of oleic acid is

$$= 8.9\underline{5} \times 10^{-6} \text{ g} \times \frac{1 \text{ mol oleic acid}}{282.47 \text{ g}} = 3.1\underline{7} \times 10^{-8} \text{ mol}$$

So, Avogadro's number here would be equal to:

$$= \frac{1.8\underline{5} \times 10^{16} \text{ oleic acid molecules}}{3.1\underline{7} \times 10^{-8} \text{ oleic acid moles}} = 5.8 \times 10^{23} \text{ molecules per mole of oleic acid.}$$

SELF-ASSESSMENT EXERCISES

115. (E)

(a) Formula unit: The smallest reducible ratio of atoms in a molecule or ionic compound
(b) P_4: An allotrope of the element phosphorus
(c) Molecular compound: A compound where bonds are formed by sharing of electrons between atoms
(d) Binary compound: A compound formed between two elements
(e) Hydrate: A molecular or ionic compound that is accompanied with a fixed number of water molecules as an adduct.

116. (E)

(a) Mole of compound: An amount of compound that contains 6.02×10^{23} molecules of that compound (or, a mass of compound equal to its molecular weight).
(b) Structural formula: The formula that shows in what order the various atoms in the molecule are connected to each other and what is the mode of their bonding
(c) Oxidation state: The apparent number of electrons from a particular atom which are involved in bonding with other atoms (either being lost, gained, or shared)
(d) C–H–O determination: A process by which C and H and O weight% in a compound are determined by careful combustion and measuring the amounts of the evolved products specific to combustion of C, H, and O.

117. (E)

(a) Molecular mass is the mass of one molecule of a compound, while molar mass is the mass of one mole of that compound (or the molecular mass multiplied by 6.02×10^{23}).
(b) Empirical formula is the simplest formula for the compound, and shows the types of atoms and their ratios, whereas a molecular formula lists the types and actual number of atoms in the formula (a molecular formula is the empirical formula times an integer).
(c) Systematic name is the name of a compound that follows established guidelines where the number and possibly oxidation state of each element is provided in the name, whereas a trivial name is a common name given to a compound that usually tells us nothing about its composition (like water or ammonia or red rust).
(d) Hydroxyl functional group is –OH, where the covalent bond to the rest of the molecule is made through the oxygen. Carboxyl functional group is –C(=O)O, where the covalent bond to the rest of the molecule is made through the carbon, and a proton can attach to one of the oxygens.

118. (E)

(a) Mass of one atom of nitrogen (in amu).
(b) Mass of one molecule of dinitrogen gas, N_2 (in amu)
(c) Mass of one mole of dinitrogen N_2 (in g/mol).

119. **(E)** The answer is (c), because 12.01 g of H_2O = 0.667 mol H_2O, which equates to 0.667 × 3 = 2.00 moles of atoms. One mole of Br_2 also has 2.00 moles of atoms.

120. **(E)** The answer is (b). N_2H_4 can be reduced further to an empirical formula of NH_2.

121. **(E)** The answer is (d), because total atomic mass is 14 for N and 7 for H.

122. **(E)** Answer is (a).
(a) 50.0 g N_2O × (1 mol N_2O/44.0 g N_2O) × (2 mol N/1 mol N_2O) = 2.27 mol
(b) 17.0 g NH_3 × (1 mol NH_3/17.0 g NH_3) × (1 mol N/1 mol NH_3) = 1.00 mol
(c) 150 mL C_5H_5N × (0.983 g/1 mL) × (1 mol Pyr/79.0 g Pyr) × (1 mol N/1 mol Pyr) = 1.87 mol
(d) 1 mol N_2 × (2 mol N/1 mol N_2) = 2.0

123. **(M)**

$$\frac{2.9\text{ g Fe}}{\text{total blood}}\times\frac{\text{total blood}}{2.6\times10^{13}\text{ red blood cells}}\times\frac{1\text{ mol Fe}}{55.8\text{ g Fe}}\times\frac{6.02\times10^{23}\text{ Fe atoms}}{\text{mol Fe}}$$

$$=1.2\times10^{9}\ \frac{\text{Fe atoms}}{\text{red blood cell}}$$

124. **(E)** Answer is (c).
Mass % of F = (19×3)/(X + 19×3) = 0.65
Solving for X, we get X = 30.7 or 31 u

125. **(E)** Answer is (c). Total formal charge on H: +4. Total charge on O: –12, and the ion has a negative charge. Therefore, oxidation state of I = –12 + 4 + 1 = 7.

126. **(E)** The correct answer is +6 (choice c). Magnesium has a +2 oxidation state, and oxygen is −2. Manganese would need to be +6 for the charge on the molecule to sum to zero.

127. **(M)** Choice (a) is hydrogen periodate and is therefore the correct answer. Na_2SO_3 is sodium sulfite, $KClO_2$ is potassium chlorite, HFO is hydrogen hypofluorite, and NO_2 is nitrogen dioxide.

128. **(E)** $Sr(HCO_3)_2$ is strontium bicarbonate (choice d).

129. **(E)** The answer is (b). Ca is a +2 ion and ClO_2^- is –1 anion.

130. **(E)** Li_3P has a molar mass of 51.79 g mol^{-1} (choice d).

131. **(E)** The answer is (d). Multiplying O atomic mass by 4 (64 u) is nearly the same as the atomic mass of Cu (63.55).

132. **(E)** The answer is (d). Answer (a) isn't correct. While having the correct number of atoms, it is not an isomer because it is only a molecular formula and gives no information on atom bonding. Answer (b) isn't correct, because it's the exact same molecule as stated in the question. Answer (c) isn't correct because it doesn't have enough atoms. Therefore, the answer is (d), because it has the correct number of atoms in a different configuration.

133. **(M)** First, find out the mass of Na_2SO_3, which is 126.0 g/mol. Then:
Mass H_2O (x) = 0.5 (x + Mass Na_2SO_3).
$x = 0.5x + 63$.
Solving for x, we obtain x = 126 g (mass of H_2O)
Since we have 126 g of water, the number of moles of H_2O is 126 g/18.0 g mol^{-1} = 7
Therefore, the formula is $Na_2SO_3 \cdot 7\ H_2O$.

134. **(M)**
(a) Based on this composition, molar mass of malachite is calculated to be 221.18 g/mol. Since there are two moles of Cu per mole of malachite, the %mass of Cu is:

$$1000 \text{ g malachite} \times \frac{1 \text{ mol mal.}}{221.18 \text{ g mal.}} \times \frac{2 \text{ mol Cu}}{1 \text{ mol mal.}} \times \frac{63.546 \text{ g Cu}}{1 \text{ mol Cu}} = 574.61 \text{ g Cu}$$

$$\% \text{ Cu} = \frac{574.61 \text{ g}}{1000 \text{ g}} \times 100 = 57.46\%$$

(b) The formula for copper(II) oxide is CuO. Therefore, for one mole of malachite, there are two moles of CuO. Therefore,

$$1000 \text{ g malachite} \times \frac{1 \text{ mol mal.}}{221.18 \text{ g mal.}} \times \frac{2 \text{ mol CuO}}{1 \text{ mol mal.}} \times \frac{79.545 \text{ g CuO}}{1 \text{ mol CuO}}$$

mass CuO = 719.5 g

135. **(D)** Molar mass of acetaminophen is 151.2 u, or 151.2 g/mol. To determine the molecular formula, calculate the moles of various constituting elements, as shown below:

mol C = 63.56 g C × (1 mol C/12.011 g C) = 5.292 mol C
mol H = 6.00 g H × (1 mol H/1.008 g H) = 5.95 mol H
mol N = 9.27 g N × (1 mol N/14.01 g N) = 0.662 mol N
mol O = 21.17 g O × (1 mol O/15.999 g O) = 1.323 mol O

Then, divide all values by the smallest to determine mole ratios:
5.92 mol C / 0.662 mol N → 7.99 mol C
5.95 mol H / 0.662 mol N → 8.99 mol H
0.662 mol N / 0.662 mol N → 1.00 mol N
1.323 mol O / 0.662 mol N → 2.00 mol C

The C:H:N:O ratio is 8:9:1:2. The empirical formula is therefore $C_8H_9NO_2$. The molar mass of this formula unit is 151.1, which is the same as the molar mass of acetaminophen. Therefore, the empirical formula obtained is also the same as the molecular formula.

136. **(D)** The first step is to determine the mass of C, H, and O.

$$\text{mol C} = 6.029 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.009 \text{ g } CO_2} \times \frac{1 \text{ mol C}}{1 \text{ mol } CO_2} = 0.1370 \text{ mol}$$

$$\text{mass of C} = 0.1370 \text{ mol C} \times (12.011 \text{ g C}/1 \text{ mol C}) = 1.646 \text{ g C}$$

$$\text{mol H} = 1.709 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.015 \text{ g } H_2O} \times \frac{2 \text{ mol H}}{1 \text{ mol } H_2O} = 0.1897 \text{ mol}$$

$$\text{mass of H} = 0.1897 \text{ mol H} \times (1.008 \text{ g H}/1 \text{ mol H}) = 0.1912 \text{ g H}$$

Mass of oxygen is obtained by difference: mass of O = 2.174 g – (1.646 + 0.1912) = 0.337 g

$$\text{mol O} = 0.337 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.0211 \text{ mol}$$

(a) % Composition:
1.646 g C / 2.174 g Ibo = 75.71% C
0.191 g H / 2.174 g Ibo = 8.79% H
0.337 g O / 2.174 g Ibo = 15.5% O

(b) To determine the empirical formula, divide all mole values by the lowest one:
0.1370 mol C / 0.0211 mol O $\rightarrow$ 6.49 mol C
0.1897 mol H / 0.0211 mol O $\rightarrow$ 8.99 mol H
0.0211 mol O / 0.0211 mol O $\rightarrow$ 1.00 mol O

The empirical formula is obtained by multiplying the above ratios by 2. The formula is $C_{13}H_{18}O_2$.

137. **(M)** To construct a concept map, one must first start with the most general concepts. These concepts contain or are defined by more specific terms and concepts discussed in those sections. In this chapter, the main themes are types of chemical compounds (3-1), the mole concept (3-2), the composition of chemical compounds (3-3), oxidation state (3-4), and naming compounds (3-5). Naming of inorganic compounds (3-6) and organic compounds (3-7) are sub-topics of section 3-6. Take a look at the subsection headings and problems for more refining of the general and specific concepts.