1. CH110 Sessional Examination, December 2011. Question Seven (20 Marks).

(a) Ethylene glycol used in automobile antifreeze, contains only carbon, hydrogen and oxygen. Combustion analysis of a 23.46 mg sample yields 20.42 mg of H_2O and 33.27 mg CO_2. What is the **empirical formula** and **molecular formula** of the ethylene glycol if it has a molecular mass of 62.0 ***amu*** [6]

Answer:

If we represent the combustion of ethylene glycol in excess oxygen using the reaction

$$C_xH_yO_z + O_2 \rightarrow CO_2 + H_2O \qquad (1)$$

interpretation of information of this equation give the table below

Descriptor	***$C_xH_yO_z$***	***CO_2***	***H_2O***	***C***	***H***
Mass of reagent (mg), m	*23.46*	*33.27*	*20.42*		
Molar mass (amu), M	*62.0*	*44.0*	*18.0*	*12.0*	*1.0*

Tabulated data for CO_2 and H_20 can be used to get the mass of carbon and hydrogen in ethylene glycol that underwent combustion.We do this by using the formulae

$$\frac{M_C}{M_{CO_2}} = \frac{m_C}{m_{CO_2}} \qquad (2) \qquad \text{and} \quad \frac{M_H}{M_{H_2O}} = \frac{m_H}{m_{H_2O}} \qquad (3)$$

These formulae simply mean that the fractional composition of an element in a compound can be obtained from ratio of the molar mass of an element to the molar mass of compound containing it or the mass of the element in a sample of known mass (that is, an aliquot) of the compound.

Determining the mass of carbon (m_C) and hydrogen (m_H) consumed during combustion of ethylene glycol simply means making these terms the subject of formula 2 and 3 followed by substituting the relevant values which gives

$$m_C = \frac{M_C \times m_{CO_2}}{M_{CO_2}} = \frac{12.0\ amu \times 33.27\ mg}{44.0\ amu} = \underline{\underline{9.07\ mg}} \qquad [1] \qquad \text{and}$$

$$m_H = \frac{M_H \times m_{H_2O}}{M_{H_2O}} = \frac{1.0\ amu \times 20.42\ mg}{18.0\ amu} = \underline{\underline{2.15\ mg}} \qquad [1]$$

The mass of oxygen (m_O) in $C_xH_yO_z$ is obtained from the equation (4) by making m_O the subject of the formula

$$m_{C_xH_yO_z} = m_C + m_H + m_O \qquad (4)$$

$$m_O = m_{C_xH_yO_z} - m_C + m_H = 23.46 - (9.07 + 2.15) = \underline{\underline{12.24\ mg}} \qquad [1]$$

Now that the masses each component element of ethylene glycol is known, we can proceed to calculating the number of moles of each component to determine the empirical formula as shown in the table below.

Parameter	***Element***		
	C	***H***	***O***
mass in mg (m) &	*9.07*	*2.15*	*12.24*
mass in g	*0.00907*	*0.00215*	*0.01224*
malar mass in g (M)	*12.0*	*1.0*	*16.0*
No. of moles $n = \frac{m}{M}$	$\frac{0.00907}{12.0} = 7.55\times10^{-4}$	$\frac{0.00215}{1.0} = 2.13\times10^{-3}$	$\frac{0.01224}{16.0} = 7.65\times10^{-4}$
Dividing by smallest no. of moles gives	$\frac{7.55\times10^{-4}}{7.55\times10^{-4}} = 1.00$	$\frac{2.13\times10^{-3}}{7.55\times10^{-4}} = 2.82$	$\frac{7.65\times10^{-4}}{7.55\times10^{-4}} = 1.01$
Integral mole ratio	*1*	*3*	*1*

The empirical formula is obtained using the integral mole ratio as a subscript of each element in the table, that is, $\underline{\underline{CH_3O}}$. *[2]*

The molar mass of the empirical formula of ethylene glycol (M_{EF}) is given by equation 5.

$$M_{EF} = M_C + (3\times M_H) + M_O \qquad (5)$$

$$M_{EF} = 12.01 + (3\times1.008) + 16.00 = 31.03 \approx \underline{\underline{31.0\ amu}}$$

Since the molar mass of the molecular formula of ethylene glycol (M_{MF}) is given as 62.0 amu, the ratio of the molar masses of the two molar masses (M_{MF} and M_{EF}) is

$$\frac{M_{MF}}{M_{EF}} = \frac{62.0}{31.0} = 2$$

so that the molecular formula of ethylene glycol is that of the empirical formula with subscripts multiplied by 2. Therefore, molecular formula is $(CH_3O)_{\times2} = \underline{\underline{C_2H_6O_2}}$ *[1]*

(b) Consider the following unbalanced equation: $H_{2(g)} + CO_{(g)} \leftrightarrow CH_3OH_{(g)}$

(i) Explain what is meant by ***stoichiometric equation***. On what **law is *stoichiometry*** based? [2]

Answer:

A stoichiometric equation is a chemical equation for which the simplest ratio of the coefficients of reacting quantities combine exactly to yield products in simplest ratio of their coefficients. It is a balanced chemical equation. *[1]*

Stoichiometry is based on the Law of Conservation of Mass. *[1]*

(ii) Suppose 1.7125 kg $CO_{(g)}$ is reacted with 0.2125 kg $H_{2(g)}$. Calculate the theoretical yield of methanol. To what volume (in dm^3) does this yield translate at stp? [6]

Answer:

We start solving this problem by writing a balanced chemical equation for the reaction by putting 2 as the coefficient of hydrogen on the reactants as shown below.

$$2H_2 + CO \rightarrow CH_3OH \qquad (6) \qquad [1]$$

Since the theoretical yield depends on the limiting reagent, we need to determine the limiting reagent. We do this by first determining the number of moles of reactants followed by comparing mole ratios with the understanding that equation 6 requires that 2 moles of hydrogen molecules (H_2) react with a mole of carbon dioxide (CO_2) to produce a mole of methanol (CH_3OH).

Descriptor	H_2	***CO***	CH_3OH	***C***	***H***
Mass of reagent (kg), m	*0.2125*	*1.7125*			
Mass of reagent (g), m	*212.5*	*1712.5*			
Molar mass (g), M	*2.0*	*28.0*	*32.0*	*12.0*	*1.0*
No. of moles $n = \frac{m}{M}$	$\frac{212.5}{2.0} = 106.25$	$\frac{1712.5}{28.0} = 61.16$			

To determine the limiting reagent we have to consider the feasibility of the following statements:

(a) If all the hydrogen in the reaction vessel reacted (106.5 moles), we would need $\frac{106.5}{2.0} = 53.25$ *moles of CO or 53.25mol×28 g/mol = 1491 g or 1.491 kg of CO.*

Since the amount of CO available for the reaction is 1.7125 kg, this reaction is feasible in that 1.491 kg < 1.7125 kg.

(b) If all the CO in the reaction vessel reacted (61.16 moles), we would need $2\times 61.16 = 122.32$ moles of H_2 or 122.32 mol $\times$ 2 g/mol = 244.64 g or 0.2446 kg of H_2. Since the amount of H_2 available for the reaction is 0.2125 Kg, this reaction is not feasible in that we have too little H_2 from the fact that 0.2446 kg >0.2125 kg.

Comparisons in (a) and (b) show that H_2 is the limiting reagent. [2]

The theoretical yield will be calculated from consumption of all H_2 moles of the feasible reaction that is, 106.5 moles that produced 53.35 moles of the methanol. The amount of methanol produced is therefore given by the equation

$$m_{CH_3OH} = n_{CH_3OH} \times M_{CH_3OH} \qquad (7)$$

$$m_{CH_3OH} = 53.25 \text{ mol} \times 32.0 \text{ g/mol} = 1704 \text{ g or } \underline{\underline{1.704 \text{ kg}}}. \qquad [2]$$

The corresponding volume of air at STP is given by equation (8) below.

$$\frac{m}{M_r} = \frac{V}{V_r} \qquad (8)$$

where V_r and M_r are the molar volume of a gas at STP (22.4 dm^3/mol) and its molar mass while V is the volume of gas of mass m at STP. Making V the subject of the above equation gives

$V = \dfrac{V_r \times m}{M_r}$ *and substituting the values of variables on the RHS gives*

$$V = \frac{22.4\ dm^3/mol\ \times 1704\ g}{32.0\ \ g/mol} = \underline{\underline{1193\ dm^3}} \qquad [1]$$

(iii) If 1.785 x 10^3 g $CH_3OH_{(g)}$ is actually produced, what is the percentage yield of methanol? [2]

The percentage yield of methanol is given by equation (9) below.

$$\%\ yield = \frac{actual\ mass}{theoretical\ mass} \times 100 \qquad (9)$$

Substituting the values determined earlier for the actual and theoretical mass gives

$$\%\ yield = \frac{1785}{1704} \times 100 = 104.75\% \approx \underline{\underline{105\%}} \qquad [2]$$

(iv) If 1.785 x 10^3 g $CH_3OH_{(g)}$ is actually produced and dissolved in a drum (200 dm^3) of water, find the concentration ofthe $CH_3OH_{(g)}$ solution? [2]

The molar concentration of the solution is given by equation 10.

$$M_{conc} = \frac{n_{CH_3OH}}{V} \qquad (10)$$

Where $n_{CH_3OH} = \dfrac{m_{CH_3OH}}{M_{CH_3OH}}$ *is the moles of methanol and V is 200 dm³.*

Recalling that 1 dm³= 1 L and substituting the appropriate values we have

$$M_{conc} = \frac{n_{CH_3OH}}{V}$$

$$M_{conc} = \frac{1}{V}\times\frac{m_{CH_3OH}}{M_{CH_3OH}} = \frac{1785\ g}{32\ g/mol}\times\frac{1}{200\ L} = \frac{1785}{32\times 200} = \underline{\underline{02789\ M}}$$ *[2]*

(v) What is the percentage composition of methanol? [2]

The percentage composition of methanol is obtained by giving the percentage fraction of each element (C, H and O) as shown in the table below:

% of element = $\dfrac{total\ \ atoms\ \ of\ \ element\times molar\ \ mass\ \ of\ \ element}{molar\ \ mass\ \ of\ \ compound}\times 100\% = \dfrac{M_{element}}{M_{CH_3OH}}\times 100\%$

	C	*H*	*O*
Atoms in CH_3OH	*1*	*4*	*1*
Element's molar mass (M)	$M_C = 12.01$	$M_H = 1.008$	$M_O = 16.00$
Total mass of element in a molecule ($M_{element}$ *)*	$12.01\times 1 = 12.01$	$1.008\times 4 = 4.032$	$16.00\times 1 = 16.00$
M_{CH_3OH}	*32.04*	*32.04*	*32.04*
% composition *[2]*	$\dfrac{12.01\times 100}{32.04} = 37.5\%$	$\dfrac{4.032\times 100}{32.04} = 12.5\%$	$\dfrac{16.00\times 100}{32.04} = 50.0\%$

2. CH110 DeferredSessional Examination, December 2011. Question Seven (20 Marks).

(a) A compound X contains 72% Mn and 28% O by mass. When X is burnt in oxygen gas a new compound Y containing 63.3% Mn and 36.7 O is formed.

(i) Determine the empirical formulas of X and Y [6]

Answer:

To determine the empirical formula, we start by assuming we are starting with a 100 g of the compound, that is, make each percentage to mass of each element in grams. Next we convert mass of element to moles. We then divide the moles of each compound by the smallest number of moles. We convert these mole ratios to integral values to get subscripts to be used in the empirical formula. As tabulated below.

Compound	X		Y	
Element	Mn	O	Mn	O
Element %	72	28	63.3	36.7
Element mass (m) in g	72	28	63.3	36.7
Element molar mass (M_r), g/mol	54.94	16.00	54.94	16.00
Element No. of moles ($n = \frac{m}{M_r}$)[2]	$\frac{74 \ g}{54.94 \ g/mol}$ $= 1.31$	$\frac{28 \ g}{16 \ g/mol}$ $= 1.75$	$\frac{63.3 \ g}{54.94 \ g/mol}$ $= 1.15$	$\frac{36.7 \ g}{16 \ g/mol}$ $= 2.29$
Mole ratio for compound[2]	$\frac{1.31}{1.31} = 1$	$\frac{1.75}{1.31} = 1.33$	$\frac{1.15}{1.15} = 1$	$\frac{2.29}{1.15} = 2$
Further conversion of mole ratios to integral value	$1 \times 3 = 3$	$1.33 \times 3 = 4$		
Empirical Formula [2]	Mn_3O_4		MnO_2	

(ii) The combustion of X yields Y. Write the balanced equation. [2]

Answer:

The equation is $Mn_3O_4 \ + \ O_2 \ \rightarrow \ 3MnO_4$ *(11)* *[2]*

(iii) Calculate the number of moles of X used and the mass of Y formed knowing that 108 ml (1.08 dL) of oxygen gas were used in this conversion at room temperature and pressure (rtp – 25^{o}C and 1 atmosphere). [4]

Answer:

In this reaction, the stoichiometry is that 1 mole of X (Mn_3O_4) reacts 1 mole of oxygen (O_2). Since the molar volume of a gas at 24.0 dm^3 or 24000 mL we know that the number of moles of oxygen used in the reaction is given by equation (12)

$$n_{O_2} = \frac{V}{V_r} \qquad (12)$$

where V=108 mL and V_r=24000 therefore

$$n_{O_2} = \frac{108}{24000} = 0.0045 \ mol$$

Since from the stoichiometry, $n_{O_2} = n_{Mn_3O_4}$*, the number of moles X (Mn_3O_4) reacted is 0.0045 and number of moles of Y (MnO_2) produced is 3×0.0045=0.0135 moles. [2]*
The mass of Y (MnO2) produced is given by equation (13)

$$m_Y \ or \ m_{MnO_2} = n_{MnO_2} \times M_{MnO_2} \qquad (13)$$

Since $M_{MnO_2} = 54.94 + 2 \times 16.00 = 86.94 \ g$ *. Putting this value and that of* $n_{MnO_2} (0.0135 \ mol)$ *into equation (13) gives*

$$m_Y \ or \ m_{MnO_2} = 0.0135 \times 86.94 \ g = \underline{\underline{1.17 \ g}} \qquad [2]$$

(b) The volume of 0.08 M KCl solution is reduced from 300 ml to 200 ml by boiling off some of the water. What is the number of moles, the molarity and the molality of the new solution? Convert molarity into ppm,that is µg/ml. [6]

Answer

Let M_i (0.08 M) and V_i (300 mL) be the concentration and volume of the solution before concentration reduction by boiling and M_f and V_f (200 mL) be the concentration and volume of the solution after concentration reduction.

Since the Law of Conservation of Mass requires that the number of moles before and after concentration reduction be the same, equation holds (14)

$$n = M_i \times V_i = M_f \times V_f \qquad (14)$$

where the volumes V_i and V_f are in litres and the concentrations M_i and M_f are expressed in molarity.

Therefore the number of moles is given as

$$n = M_i \times V_i = \frac{0.08 \ mol/L \ \times 300 \ mL}{1000 \ mL/L} = \underline{\underline{0.024 \ mol}} \qquad [1]$$

Making M_f the subject of equation 14 allows us to determine the molarity concentration of the solution with reduced volume provided the final volume is litres (L)

$$M_f = \frac{n = M_i \times V_i}{V_f} = \frac{0.024 \ mol}{0.200 \ L} = \underline{\underline{0.12 \ M \ \ (or \ mol/L)}} \qquad [1]$$

Molality is defined as moles per kilogram of solution as given in equation 15 below.

$$molality \ (m) \ = \frac{number \ moles \ of \ solute}{mass \ of \ solution \ in \ kg} \text{ or}$$

$$m_{molality} = \frac{n_{solute}}{solution\ mass\ in\ kg} \qquad (15)$$

$$m_{molality} = \frac{0.024\ mol}{0.200\ kg} = \underline{\underline{0.12\ m\ (or\ mol/kg)}} \qquad [1]$$

Converting molarity to ppm (or µg/mL) calculate the mass of the solute (0.024 moles) of KCl and divide by the volume which is 200 mL. The molar mass of KCl is $M_{KCl}=M_K+M_{Cl}$ = 35.45+39.10 = 74.55 g/mol. The mass of solute in the solution is therefore $m_{KCl} = n_{KCl} \times M_{KCl}$ = 0.024 × 74.55=1.7892 g. This sample in µg is 1.7892 × 1000 µg or 1789.2 µg. [1]

The concentration in ppm is given by the equation

$$M_{conc} = \frac{m}{V} = \frac{1789.2}{200} = 8.946\ ppm \approx \underline{\underline{8.9\ ppm}} \qquad [1]$$

(c) Consider the equation 2A + B → A_2B. If 1.0 mol of A and 1.0 mol of B is reacted to produce the maximum amount of A_2B, how many moles of each substance will be present at the end of the reaction? [2]

Answer:
The table below shows the information about this reaction.

Equation	2A +	B →	A_2B
Present before reaction	1 mole	1 mole	0 mole
Required for reaction	1 mole	0.5 mole	1 mole
Excess or end of reaction	0 mole	0.5 mole	1mole

A is the limiting reagent in this reaction and it was all consumed in the reaction while 0.5 mole and 1 mole of B and A_2B are present at the end of reaction. [2]

3. CH110 Test 1, 2007. Question Four.

(a) A compound contains only carbon, hydrogen, nitrogen and oxygen. Combustion of 0.157 g of the compound produced 0.213 g CO_2 and 0.0310 g H_2O. In another experiment, it is found that 0.103 g of the compound produces 0.0230 g NH_3. What is the empirical formula of the compound? Hint: combustion involves the use of excess oxygen. Assume that all the carbon ends up in CO_2 and all the hydrogen ends up in H_2O. Also assume that all the nitrogen ends up in NH_3 [5]

Answer:

This problem is similar to 1(a) above. If we represent the combustion of unknown compound $C_wH_xO_yN_z$ in excess oxygen using the reaction

$$C_wH_xO_yN_z + O_2 \rightarrow CO_2 + H_2O \qquad (1a)$$

The reaction that produces ammonia is represented as

$$C_wH_xO_yN_z + O_2 \rightarrow NH_3 \qquad (1b)$$

interpretation of information of equation (1a) gives the table below

Descriptor	***$C_wH_xO_yN_z$***	***CO_2***	***H_2O***	***C***	***H***
Mass of reagent (g), m	*0.157*	*0.213*	*0.0310*		
Molar mass (g), M_r	*unknown*	*44.0*	*18.0*	*12.0*	*1.0*
Element mass (g), m_{el}				*0.0581*	*0.00172*
$\%\ composition = \frac{m_{el}}{m_{C_wH_xO_yN_z}} \times 100$				$= \frac{0.0581}{0.157} \times 100 = 37.0$	$= \frac{0.00172}{0.157} \times 100 = 1.10$

Tabulated data for CO_2 and H_20 can be used to get the mass of carbon and hydrogen in $C_wH_xO_yN_z$ that underwent combustion. We do this by using the formulae

$$\frac{M_C}{M_{CO_2}} = \frac{m_C}{m_{CO_2}} \quad (2a) \quad \text{and} \quad \frac{M_H}{M_{H_2O}} = \frac{m_H}{m_{H_2O}} \quad (3a)$$

These formulae simply mean that the fractional composition of an element in a compound can be obtained from ratio of the molar mass of an element to the molar mass of compound containing it or the mass of the element in a sample of known mass (that is, an aliquot) of the compound.

Determining the mass of carbon (m_C) and hydrogen (m_H) consumed during combustion of $C_wH_xO_yN_z$ simply means making these terms the subject of formula 2a and 3a followed by substituting the relevant values which gives

$$m_C = \frac{M_C \times m_{CO_2}}{M_{CO_2}} = \frac{12.0\ g \times 0.213\ g}{44.0\ g} = \underline{\underline{0.05809\ g}} \qquad [1] \qquad \text{and}$$

$$m_H = \frac{M_H \times m_{H_2O}}{M_{H_2O}} = \frac{1.0\ g \times 0.0310\ g}{18.0\ g} = \underline{\underline{0.00172\ g}} \qquad [1]$$

interpretation of information of equation (1b) gives the table below

Descriptor	***$C_wH_xO_yN_z$***	***NH_3***	***N***	***H***
Mass of reagent (g), m	*0.103*	*0.0230*		

Molar mass (g), M	*unknown*	*17.0*	*14.0*	*1.0*
Element mass (g), m_{el}			*0.0189*	
$\% \ composition = \frac{m_{el}}{m_{C_wH_xO_yN_z}} \times 100$			$= \frac{0.0189}{0.103} \times 100 = 18.4$	

The mass of nitrogen is obtained from the formula

$$\frac{M_N}{M_{NH_3}} = \frac{m_N^{'}}{m_{NH_3}^{'}} \qquad (4a)$$

Determining the mass of nitrogen (m_N) consumed during production of ammonia from $C_wH_xO_yN_z$ simply means making these terms the subject of formula 4a followed by substituting the relevant values which gives

$$m_N^{'} = \frac{M_N \times m_{NH_3}^{'}}{M_{NH_3}} = \frac{14.0\ g \times 0.0230\ g}{17.0\ g} = \underline{\underline{0.0189\ g}} \qquad [1]$$

The percentage composition of nitrogen of $C_wH_xO_yN_z$calculated in the above table can be used to calculate the mass of nitrogen in a sample of mass 0.157 g (that is, m_N) according to equation (5a)

$$m_N = 0.157 \times \frac{nitrogen\ percentage\ composition\ in\ C_wH_xO_yN_z}{100} \qquad (5a)$$

Substituting the relevant value of nitrogen percentage composition gives

$$m_N = 0.157 \times \frac{18.4}{100} = \underline{\underline{0.02881\ g}}$$

The mass of oxygen (m_O) in $C_wH_xO_yN_z$ is obtained from the equation (6a) by making m_O the subject of the formula

$$m_{C_wH_xO_yN_z} = m_C + m_H + m_O + m_N \qquad (6a)$$

$$m_O = m_{C_wH_xO_yN_z} - (m_C + m_H + m_N)$$

Substituting the relevant values gives

$$m_O = 0.157 - (0.05809 + 0.00172 + 0.02881)\ g$$

$$m_O = 0.157 - 0.8862g = \underline{\underline{0.06838\ g}} \qquad [1]$$

Now that the masses each component elements of $C_wH_xO_yN_z$ is known, we can proceed to calculating the number of moles of each component to determine the empirical formula as shown in the table below.

Parameter	***Element***			
	C	***H***	***O***	***N***
mass in g	*0.05809*	*0.001722*	*0.06838*	*0.02881*
malar mass in g (M)	*12.0*	*1.0*	*16.0*	*14.0*
No. of moles $n = \frac{m}{M}$	$\frac{0.05809}{12.0} = 4.84 \times 10^{-3}$	$\frac{0.001722}{1.0} = 1.72 \times 10^{-3}$	$\frac{0.06838}{16.0} = 4.27 \times 10^{-4}$	$\frac{0.02881}{14.0} = 2.06 \times 10^{-3}$
Dividing by smallest no. of moles gives	$\frac{4.84 \times 10^{-3}}{1.72 \times 10^{-3}} = 2.81$	$\frac{1.72 \times 10^{-3}}{1.72 \times 10^{-3}} = 1.0$	$\frac{4.27 \times 10^{-3}}{1.72 \times 10^{-3}} = 2.48$	$\frac{2.06 \times 10^{-3}}{1.72 \times 10^{-3}} = 1.20$
Rounded mole ratio	*3*	*1*	*2.5*	*1*
Integral mole ratio	*3x2=6*	*2x1=2*	*2.5x2=5*	*2x1=2*

The empirical formula is obtained using the integral mole ratio as a subscript of each element in the table, that is, $\underline{\underline{C_6H_2O_5N_2}}$. [1]

(b) A compound contains only carbon, hydrogen and oxygen. Combustion of 10.68 mg of the compound yields 16.01 mg CO_2 and 4.37 mg H_2O. The molar mass of the compound is 176.1 g/mol. What are the empirical and molecular formulas of the compounds [10]

Answer:

This question is similar to question to question 1(a)

If we represent the combustion of $C_xH_yO_z$ *in excess oxygen using the reaction*

$$C_xH_yO_z + O_2 \rightarrow CO_2 + H_2O \qquad (1c)$$

interpretation of information of this equation give the table below

Descriptor	$C_xH_yO_z$	CO_2	H_2O	**C**	**H**
Mass of reagent (mg), m	*10.68*	*16.01*	*4.37*		
Molar mass (g/mol), M	*176.1*	*44.0*	*18.0*	*12.0*	*1.0*

Tabulated data for CO_2 *and* H_20 *can be used to get the mass of carbon and hydrogen in ethylene glyThe molar mass of the empirical formula of* $C_xH_yO_z$ *(*M_{EF}*) is given by equation 5c.*

$$M_{EF} = (3 \times M_C) + (2 \times M_H) + (3 \times M_O) \qquad (5c)$$

$$M_{EF} = (3 \times 12.01) + (2 \times 1.008) + (3 \times 16.00) = 86.05 \approx \underline{\underline{86.1\ amu}}$$

Since the molar mass of the molecular formula of $C_xH_yO_z$ (M_{MF}) is given as 176.1 amu, the ratio of the molar masses of the two molar masses (M_{MF} and M_{EF}) is

$$\frac{M_{MF}}{M_{EF}} = \frac{176.1}{86.1} = 2.045 \approx 2 \qquad [1]$$

so that the molecular formula of $C_xH_yO_z$ is that of the empirical formula with subscripts multiplied by 2. Therefore, molecular formula is $\left(C_3H_2O_3\right)_{\times 2} = \underline{\underline{C_6H_4O_6}}$ [1]

col that underwent combustion. We do this by using the formulae

$$\frac{M_C}{M_{CO_2}} = \frac{m_C}{m_{CO_2}} \quad (2c) \qquad \text{and} \qquad \frac{M_H}{M_{H_2O}} = \frac{m_H}{m_{H_2O}} \quad (3c)$$

These formulae simply mean that the fractional composition of an element in a compound can be obtained from ratio of the molar mass of an element to the molar mass of compound containing it or the mass of the element in a sample of known mass (that is, an aliquot) of the compound.

Determining the mass of carbon (m_C) and hydrogen (m_H) consumed during combustion of $C_xH_yO_z$simply means making these terms the subject of formula 2c and 3c followed by substituting the relevant values which gives

$$m_C = \frac{M_C \times m_{CO_2}}{M_{CO_2}} = \frac{12.0\ g \times 16.01\ mg}{44.0\ g} = \underline{\underline{4.37\ mg}} \qquad [1] \qquad \text{and}$$

$$m_H = \frac{M_H \times m_{H_2O}}{M_{H_2O}} = \frac{1.0\ g \times 4.37\ mg}{18.0\ g} = \underline{\underline{0.243\ mg}} \qquad [1]$$

The mass of oxygen (m_O) in $C_xH_yO_z$ is obtained from the equation (4c) by making m_O the subject of the formula

$$m_{C_xH_yO_z} = m_C + m_H + m_O \qquad (4c)$$

$$m_O = m_{C_xH_yO_z} - m_C + m_H = 10.68 - (4.37 + 0.243) = \underline{\underline{6.067\ mg}} \qquad [1]$$

Now that the masses each component element of $C_xH_yO_z$is known, we can proceed to calculating the number of moles of each component to determine the empirical formula as shown in the table below.

Parameter	**Element**		
	C	**H**	**O**
mass in mg (m) &	*4.37*	*0.243*	*6.067*
mass in g	*0.00437*	*0.000243*	*0.006067*
malar mass in g (M)	*12.0*	*1.0*	*16.0*
No. of moles $n=\frac{m}{M}$	$\frac{0.00437}{12.0}=3.64\times10^{-4}$ *[1]*	$\frac{0.000243}{1.0}=2.43\times10^{-4}$ *[1]*	$\frac{0.006067}{16.0}=3.79\times10^{-4}$ *[1]*
Dividing by smallest no. of moles gives	$\frac{3.64\times10^{-4}}{2.43\times10^{-4}}=1.50$	$\frac{2.43\times10^{-4}}{2.43\times10^{-4}}=1.00$	$\frac{3.79\times10^{-4}}{2.43\times10^{-4}}=1.56$
Integral mole ratio	*1.50 x2=3*	*1x2=2*	*1.56x2=3.13≈3*

The empirical formula is obtained using the integral mole ratio as a subscript of each element in the table, that is, $\underline{\underline{C_3H_2O_3}}$. *[1]*

The molar mass of the empirical formula of $C_xH_yO_z(M_{EF})$ *is given by equation 5c.*

$$M_{EF}=(3\times M_C)+(2\times M_H)+(3\times M_O) \qquad (5c)$$

$$M_{EF}=(3\times12.01)+(2\times1.008)+(3\times16.00)=86.05 \approx \underline{\underline{86.1\ amu}}$$

Since the molar mass of the molecular formula of $C_xH_yO_z(M_{MF})$ *is given as 176.1amu, the ratio of the molar masses of the two molar masses (*M_{MF} *and* M_{EF}*) is*

$$\frac{M_{MF}}{M_{EF}}=\frac{176.1}{86.1}=2.045\approx2 \qquad [1]$$

so that the molecular formula of $C_xH_yO_z$ *is that of the empirical formula with subscripts multiplied by 2. Therefore, molecular formula is* $(C_3H_2O_3)_{\times2}=\underline{\underline{C_6H_4O_6}}$ *[1]*

(c) A 6.20 mg sample of an organic compound containing carbon, hydrogen and oxygen only was burned in an excess of oxygen, yielding 8.80 mg of CO_2 and 5.40 mg of H_2O. Calculate the empirical formula of the compound. [5]

Answer:

This question is similar to question to question 1(a)

If we represent the combustion of $C_xH_yO_z$ *in excess oxygen using the reaction*

$$C_xH_yO_z + O_2 \rightarrow CO_2 + H_2O \qquad (1c)$$

interpretation of information of this equation give the table below

Descriptor	$C_xH_yO_z$	CO_2	H_2O	**C**	**H**
Mass of reagent (mg), m	*6.20*	*8.80*	*5.40*		
Molar mass (g/mol), M		*44.0*	*18.0*	*12.0*	*1.0*

Tabulated data for CO_2 and H_20 can be used to get the mass of carbon and hydrogen in $C_xH_yO_z$that underwent combustion. We do this by using the formulae

$$\frac{M_C}{M_{CO_2}} = \frac{m_C}{m_{CO_2}} \qquad (2c) \qquad and \qquad \frac{M_H}{M_{H_2O}} = \frac{m_H}{m_{H_2O}} \qquad (3c)$$

These formulae simply mean that the fractional composition of an element in a compound can be obtained from ratio of the molar mass of an element to the molar mass of compound containing it or the mass of the element in a sample of known mass (that is, an aliquot) of the compound.

2.

Determining the mass of carbon (m_C) and hydrogen (m_H) consumed during combustion of$C_xH_yO_z$ simply means making these terms the subject of formula 2c and 3c followed by substituting the relevant values which gives

$$m_C = \frac{M_C \times m_{CO_2}}{M_{CO_2}} = \frac{12.0 \ g \times 8.80 \ mg}{44.0 \ g} = \underline{\underline{2.40 \ mg}} \quad [1] \qquad and$$

$$m_H = \frac{M_H \times m_{H_2O}}{M_{H_2O}} = \frac{1.0 \ g \times 5.40 \ mg}{18.0 \ g} = \underline{\underline{0.300 \ mg}} \quad [1]$$

The mass of oxygen (m_O) in $C_xH_yO_z$ is obtained from the equation (4c) by making m_O the subject of the formula

$$m_{C_xH_yO_z} = m_C + m_H + m_O \qquad (4c)$$

$$m_O = m_{C_xH_yO_z} - m_C + m_H = 6.20 - (2.40 + 0.300) = \underline{\underline{3.50 \ mg}} \quad [1]$$

Now that the masses each component element of $C_xH_yO_z$ is known, we can proceed to calculating the number of moles of each component to determine the empirical formula as shown in the table below.

Parameter	***Element***		
	C	***H***	***O***
mass in mg (m) &	*2.40*	*0.300*	*3.50*
mass in g	*0.00240*	*0.000300*	*0.00350*
malar mass in g (M)	*12.0*	*1.0*	*16.0*
No. of moles $n=\frac{m}{M}$ *[1]*	$\frac{0.00240}{12.0}=2.0\times10^{-4}$	$\frac{0.000300}{1.0}=3.00\times10^{-4}$	$\frac{0.00350}{16.0}=2.19\times10^{-4}$
Dividing by smallest no. of moles gives	$\frac{2.0\times10^{-4}}{2.0\times10^{-4}}=1.00$	$\frac{3.00\times10^{-4}}{2.0\times10^{-4}}=1.50$	$\frac{2.19\times10^{-4}}{2.0\times10^{-4}}=1.1$
Integral mole ratio	*1.00 x2=2*	*1.5x2=3*	*1.1x2=2.2* $\approx$ *2*

The empirical formula is obtained using the integral mole ratio as a subscript of each element in the table, that is, $\underline{\underline{C_2H_3O_2}}$. *[1]*

(d) A certain compound consists of 47.4 % S and 52.6 % Cl. Its molecular mass is approximately 135 amu. Determine the molecular formula? [5]

Appendix: Atomic masses of some elements in the test

Element	**Atomic Mass**
Carbon	12.01
Oxygen	16.00
Nitrogen	14.01
Hydrogen	1.008
Sulphur	32.07
Chlorine	35.5

Answer:

To determine the empirical formula, we start by assuming we are starting with a 100 g of the compound, that is, make each percentage to mass of each element in grams. Next we convert mass of element to moles. We then divide the moles of each compound by the smallest number of moles. We convert these mole ratios to integral values to get subscripts to be used in the empirical formula. As tabulated below.

Compound	S_xCl_y	
Element	S	Cl
Element %	47.4	52.6
Element mass (m) in g	47.4	52.6
Element molar mass (M_r), amu	32.07	35.5
Element No. of moles ($n = \frac{m}{M_r}$) [1]	$\frac{47.4\ g}{32.07\ g/mol} = 1.478$	$\frac{52.6\ g}{35.5\ g/mol} = 1.481$
Mole ratio for compound [1]	$\frac{1.478}{1.478} = 1$	$\frac{1.481}{1.481} = 1$
Empirical Formula [1]	SCl	

The molar mass of the empirical formula of S_xCl_y(M_{EF}) is given by equation 5.

$$M_{EF} = M_S + M_{Cl} \qquad (5)$$

$$M_{EF} = 32.07 + 35.5 = \underline{\underline{67.57\ amu}}$$

Since the molar mass of the molecular formula of S_xCl_y(M_{MF}) is given as 135 amu, the ratio of the molar masses of the two molar masses (M_{MF} and M_{EF}) is

$$\frac{M_{MF}}{M_{EF}} = \frac{135}{67.57} = 2 \qquad [1]$$

so that the molecular formula of S_xCl_y is that of the empirical formula with subscripts multiplied by 2. Therefore, molecular formula is $(SCl)_{\times 2} = \underline{\underline{S_2Cl_2}}$ *[1]*

4. **FO 130 Test 1, June 2006. Question One.**

If 15 g of H_2 react with 96 g of O_2 to produce the maximum possible amount of water, how many grams of each substance will be present at the end of the reaction?

Answer:

The balanced equation of the chemical reaction is $2H_2 + O_2 \rightarrow H_2O$. The table below shows that at the end of the reaction 3 g, 0 g and 108 g of H_2, O_2 and H_2O at the end of the reaction, respectively.

Equation	$2H_2$ +	$O_2 \rightarrow$	$2H_20$
Present before reaction(g), m	15	96	0
Molar mass (g), M_r	2.0	32.0	18.0
Moles before reaction, n=m/M_r	15/2 = 7.5	96/32 = 3.0	
Moles required for reaction, n_{req}	6 mole	3 mole	6 mole
Excess or moles @ end of reaction,n_{ex}=n-n_{req}	1.5 mole	0 mole	6 mole
Amount (g) at end of reaction, m_{end}=$n_{ex}\times M_r$	1.5x2=$\underline{\underline{3.0}}$	0	18x6=$\underline{\underline{108.0}}$

5. FO 130 Test 1, June 2006. Question Two.

The concentration of glucose ($C_6H_{12}O_6$) in human blood ranges from about 30 mg/dL before meals up to 120 mg/dL after eating.

(a) Find the molarity of glucose in blood before and after eating?

Answer:

1 dL =100 mL or 0.1L;

Molar mass of carbon, hydrogen and oxygen atoms are 12 g, 1 g, and 16 g respectively.
$\therefore$ molar mass of glucose is 12$\times$6+12$\times$1+16$\times$6=72+12+96=84+96=180 g/mol.

30 mg is 0.030 g or $\left(\frac{0.030\ g}{180\ g/mol}\right) = 1.667\times10^{-4}\,mol$ *and*

120 mg is 0.120 g or $\left(\frac{0.120\ g}{180\ g/mol}\right) = 6.667\times10^{-4}\,mol$

Therefore, 30 mg/dL =1.667$\times10^{-4}$ *mol/0.1L=1.667*$\times10^{-3}$ *mol/L =1.667*$\times10^{-3}$ *M*

Similarly, 120 mg/dL = 6.667$\times10^{-3}$ *M*

Thus the concentration of glucose before and after eating is $\underline{\underline{1.667\times10^{-3}M}}$ *and* $\underline{\underline{6.667\times10^{-3}M}}$ *, respectively.*

(b) Express the above concentration in ppm?

Answer:

Expressing the above concentrations in ppm requires us to recall that 1 ppm is 1µg/mL.
Since we know that 1 mg = 1 000 µg and 1 dL = 100 mL.

We can therefore calculate that 30 mg/dL= $\frac{30\times1000\ \mu g}{100\ mL} = 300\mu g/mL = \underline{\underline{300\ \ ppm}}$.

Similarly, 120 mg/dL = $\frac{120\times 1000\ \mu g}{100\ mL} = 1200\mu g/mL = \underline{\underline{1200\ ppm}}$

6. CH110Sessional examination 2005. Question Six.

(a) The molecular mass of citric acid is 192.13 and the compound is 37.51% C, 58.29% O and 4.20% H. Determine the empirical formula and the molecular formula of citric acid?

Answer

To determine the empirical formula, we start by assuming we are starting with a 100 g of the compound, that is, make each percentage to mass of each element in grams. Next we convert mass of element to moles. We then divide the moles of each compound by the smallest number of moles. We convert these mole ratios to integral values to get subscripts to be used in the empirical formula. As tabulated below.

Compound	Citric acid		
Element	C	H	O
Element %	37.51	4.20	58.29
Element mass (m) in g	37.51	4.20	58.29
Element molar mass (M_r), amu	12.01	1.008	16.00
Element No. of moles ($n=\frac{m}{M_r}$)	$\frac{37.51\ g}{12.01\ g/mol} = 3.123$	$\frac{4.20\ g}{1.008\ g/mol} = 4.167$	$\frac{58.29\ g}{16.00\ g/mol} = 3.643$
Mole ratio for compound	$\frac{3.123}{3.123}=1$	$\frac{4.167}{3.123}=1.33$	$\frac{3.643}{3.123}=1.167$
Mole ratio x3	3X1=3	1.33x3=4	1.167x3=3.50
Mole ratio x3x2	3x2=6	4x2=8	3.5x2=7
Empirical Formula	$C_6H_8O_7$		

The molar mass of the empirical formula of citric acid (M_{EF}) is given by equation 5d.

$$M_{EF} = M_C + M_H + M_O \qquad (5d)$$

$$M_{EF} = (6\times 12.01) + (8\times 1.008) + (16\times 7) = \underline{\underline{192.124 \approx 192.12\ amu}}$$

Since the molar mass of the molecular formula of citric acid (M_{MF}) is given as 192.13amu, the ratio of the molar masses of the two molar masses (M_{MF} and M_{EF}) is

$$\frac{M_{MF}}{M_{EF}} = \frac{192.13}{192.12} = 1$$

so that the molecular formula of citric acid is that of the empirical formula with subscripts multiplied by 1. Therefore, molecular formula is $(C_6H_8O_7)_{\times 1} = \underline{\underline{C_6H_8O_7}}$

7. FO 130 Test 1, June 2005. Question One.

When iodine (I_2) reacts directly with fluorine (F_2), a compound containing 57.2% by mass of iodine is formed.

(a) Determine the empirical formula of the compound. [5]

Answer

To determine the empirical formula, we start by assuming we are starting with a 100 g of the compound, that is, make each percentage to mass of each element in grams. Next we convert mass of element to moles. We then divide the moles of each compound by the smallest number of moles. We convert these mole ratios to integral values to get subscripts to be used in the empirical formula. As tabulated below.

Compound	I_xF_y	
Element	*I*	*F*
Element %	*57.2*	*100-57.2=42.8*
Element mass (m) in g	*57.2*	*42.8*
Element molar mass (M_r), amu	*126.9*	*19.00*
Element No. of moles ($n = \frac{m}{M_r}$ *) [2]*	$\frac{57.2\ g}{126.9\ g/mol} = 0.451$	$\frac{42.8\ g}{19.00\ g/mol} = 2.253$
Mole ratio for compound [2]	$\frac{0.451}{0.451} = 1.0$	$\frac{2.253}{0.451} = 5.0$
Empirical Formula [1]	IF_5	

(b) The empirical formula of this compound is the same as the molecular formula. Write the balanced equation for the formation of this compound. [4]

Answer

The balanced chemical equation is $I_2 + 5F_2 \rightarrow 2IF_2$ *[4]*

(c) Why is it necessary to balance a chemical reaction? [4]

Answer:

It is necessary balance a chemical reaction because matter is conserved during a chemical reaction, that is, atoms are merely rearranged bond breaking and/or formation during a chemical reaction. [4]

(d) How many grams of iodine (I_2) are needed to produce 100 g of the product if the reaction has 90.09% yield? [6]

Answer:

%yield $\frac{\text{actual yield}}{\text{Theoretical yield}} \times 100$ [1]

Since % yield = 90.09 % and actual yield is 100 g, therefore theoretical yield is 100/0.9009=111 g. [2]

The iodine composition of the product is given as

% Iodine composition in $IF_5 = \frac{M_I}{M_{IF_5}} \times 100 = \frac{m_I}{m_{IF_5}} \times 100 = 57.2\%$

The amount of Iodine need to produce 111g of IF5 is obtained from making MI the subject of the above formula that is

$$m_I = \frac{57.2 \times m_{IF_5}}{100}$$ [1]

Substituting $m_{IF_5} = 111\ g$ *in the above equations gives*

$$m_I = \frac{57.2 \times 111}{100} = 63.492\ g \approx \underline{\underline{63.5\ g}}$$ [2]

(e) What mass of the compound contains exactly the same number of iodine atoms as 73.5 g of iodine (I_2)? [6]

Answer

From the equation $m_I = \frac{57.2 \times m_{IF_5}}{100}$*, we have to make* m_{IF_5} *the subject of the formula so that we have the equation* $m_{IF_5} = \frac{100 \times m_I}{57.2}$ [3]

On substituting $m_I = 73.5\ g$ *this equation gives* $m_{IF_5} = \frac{100 \times 73.5}{57.2} = \underline{\underline{128.5\ g}}$ [3]

8. FO 130 Test 1, June 2005. Question Two.

A stock solution is prepared by dissolving 10.0 g of fluoxymesterone ($C_{20}H_{29}FO_3$) in enough water to give a total volume of 500.0 mL. A 100.0 µm (microlitre) aliquot (portion) of this solution is diluted to a final volume of 100.0 mL.

(a) Calculate the concentration of the final solution in terms of molarity. [5]

Answer

The number of moles of the stock solution is given by the equation

$$n_{stock} = \frac{m}{M_r} \qquad [\tfrac{1}{2}]$$

where m =10.0 g of fluoxymesterone and

M_r is the molar mass of fluoxymesterone and

$$M_r = (20 \times M_C) + (29 \times M_H) + M_F + (3 \times M_O)$$

Substituting M_C=12.01 g, M_H=1.008 g, M_F=19.00 g, and M_O=16.00 g gives

$$M_r = (20 \times 12.01) + (29 \times 1.008) + 19.00 + (3 \times 16.00)$$

$$M_r = (240.2) + (29.232) + 19.00 + (48.00) = \underline{\underline{336.432\ g}} \qquad [\tfrac{1}{2}]$$

Therefore

$$n_{stock} = \frac{10.0\ g}{336.432\ g/mol} = \underline{\underline{0.02972\ mol}} \qquad [\tfrac{1}{2}]$$

The molarity concentration of the stock solution is given by the equation

$$M_{stock} = \frac{n_{stock}}{V_{stock}} \qquad [\tfrac{1}{2}]$$

whereV_{stock}= 500 mL that is 0.5L. Therefore

$$M_{stock} = \frac{0.02972\ mol}{0.5\ L} = 0.05944\ mol/L = \underline{\underline{0.05944\ M}} \qquad [\tfrac{1}{2}]$$

Let $V_{aliquot} = 100 \mu L = 1.00 \times 10^{-4} L$ so that the number of moles in the aliquot solution is

$$n_{aliquot} = V_{aliquot} \times M_{stock} = 0.05944 \times 1.0 \times 10^{-4} = \underline{\underline{5.944 \times 10^{-6}\ mol}} \qquad [\tfrac{1}{2}]$$

The final molarity concentration of the diluted solution is

$$M_{final} = \frac{n_{aliquot}}{V_{final}} \qquad [1]$$

whereV_{final} = 100 mL or 0.1L. Substituting this value in the above equation gives

$$M_{final} = \frac{5.944 \times 10^{-6}\ mol}{0.1\ L} = \underline{\underline{5.944 \times 10^{-5} M}} \qquad [1]$$

(b) Convert molarity into ppm (parts per million), that is, µg/ml. [6]

Answer:

Concentration in ppm (m_{conc}) isgiven by the equation

$$m_{conc} = \frac{mass\ of\ \ solute\ in\ \ \mu g}{volume\ of\ \ solution\ in\ mL} = \frac{n_{aliquot}\ mol \times M_r\ \ g/mol \times 10^6\ \ \mu g/g}{V_{final}\ mL}$$ *[3]*

Substituting the appropriate values gives

$$m_{conc} = \frac{5.944 \times 10^{-6}\ \ mol \times 336.432\ \ g/mol \times 10^6\ \ \mu g/g}{100\ \ mL}$$ *[2]*

$$m_{conc} = 5.944 \times 3.36432 = \underline{\underline{20.0\ \ ppm}}$$ *[1]*

(c) Give a qualitative comment (no calculations involved) on how both *the molarity* and *the number of moles change* (increases, remains same or decreases) from *stock solution to aliquot and, ** from aliquot to diluted solution. [4]

Answer

The stock solution and the aliquot have the same molarity but the aliquot has much smaller number of moles than the stock solution. *[2]*

The aliquot and the diluted solution have the same number of moles but the diluted solution has a much lower molarity. *[2]*

9. CH110Test 1, April 2002. Question Four.

The mass of a single atom of an element was found to be 3.9867 x 10^{-23} g.

(a) Name the element.

Answer

The molar mass this element = mass of atom $(m_a) \times N_A$ *where N_A is Avogadro's number*

= (3.9867 x 10^{-23} g) $\times \left(6.022 \times 10^{23}\ \ /mol\right)$

=3.9867$\times$6.022 g/mol = \underline{24.01 g/mol}

This element is sodium.

(b) Write the formula of the phosphate of the element identified

Answer:

The formula of the phosphate of this element is $\underline{\underline{Na_3PO_4}}$

10. CH110Test 1, April 2002. Question Five.

(a) The % composition by mass of chlorine in a chloride of iron is 55.91%. Calculate the empirical formula of the chloride of iron and name the compound.

(b) A compound X_2O_3 contains 68% by mass of element X. Calculate clearly what the element X is.

11. CH110Test 1, April 2002. Question Six.

(a) Calculate the empirical formula and the molecular formula of a compound given that its composition % by mass includes hydrogen 18.2%, carbon 81.3% and its relative molecular mass is 88.

(b) Calculate the volume of 2 M $NaOH_{(aq)}$ that would neutralize 12.5 cm^3 of 1 M H_2SO_4.

12. CH110Test 1, April 2002. Question Seven.

16.25 g of powdered zinc found to be 25% of a powdered substance labeled J was reacted with 500 cm^3 of 0.5 M HCl and 2.7 dm^3 of hydrogen gas was collected measured at rtp.

(a) Calculate the weight of the impure substance J.

(b) Determine the limiting reactant.

(c) Determine the reactant in excess.

(d) What is the practical yield?

(e) Calculate the theoretical yield and finally

(f) Calculate the % yield of the gas collected.