ANSWERS TO "TRY YOURSELF" PROBLEMS FROM STUDY SECTION 2.8

Try Yourself 2.18

What is the percent composition of Cl in $Pt(NH_3)_2Cl_2$?

Answer:

You will need the molar masses of Cl and that of the whole compound.

M_{Cl} = 35.45 $g.mol^{-1}$ and $M_{Pt(NH3)Cl2}$ = 300.04 $g.mol^{-1}$

Mass % of Cl in $Pt(NH_3)_2Cl_2$ = [2(M_{Cl}) / $M_{Pt(NH3)2Cl2}$] x 100/1 (Remember there are 2 Cl atoms).

= (70.9 $g.mol^{-1}$ / 300.04 $g.mol^{-1}$) x 100/1 = 23.63 % Cl in the compound.

So, you can say that 23.63% of the compound is chlorine, the rest will be platinum, nitrogen and hydrogen. You can also say that for every 100 grams of the compound, 23.63 grams will be chlorine.

Try Yourself 2.19

What mass of lead is present in 10.0 g of PbS?

Answers:

You will need the molar mass of PbS as well as S.

M_S = 32.1 $g.mol^{-1}$ and M_{PbS} = 239.3 $g.mol^{-1}$

% S in PbS = (32.1 $g.mol^{-1}$ / 239.3 $g.mol^{-1}$) x 100/1 = 13.41 % S in PbS.

Therefore: In 10.0 g of PbS there will be 13.41 % (or 1.34 gram) of Sulphur, the rest (8.66 gram) will be lead (Pb).

Try Yourself 2.20

Eugenol is the major component in oil of cloves. It has a molar mass of 164.2 g/mol and is 73.14% C and 7.37% H; the remainder is oxygen. Calculate the empirical and molecular formulas of eugenol.

Answer:

73.14 % C = 73.14 g C = 73.14 g / 12 $g.mol^{-1}$ = 6.10 mol C

7.37 % H = 7.37 g H = 7.37 g / 1.01 $g.mol^{-1}$ = 7.30 mol H

19.49 % O = 19.49 g O = 19.49 g / 16 $g.mol^{-1}$ = 1.22 mol O

Calculate the mol ratios between all the elements:

6.10 mol C / 1.22 mol O : 7.30 mol H / 1.22 mol O : 1.22 mol O / 1.22 mol O

= 5 mol C : 5.98 mol H : 1 mol O

= 5C : 6H : 1O

Empirical formula = C_5H_6O with a molar mass of 82.06 $g.mol^{-1}$

The given molar mass is 164.2 $g.mol^{-1}$ which is two times the molar mass of the empirical formula.

Therefore: **The molecular formula** will be 2x the empirical formula, which will be **$C_{10}H_{12}O_2$**

Try Yourself 2.21

Formula from mass (data from lab experiments)

Tin metal (Sn) and purple iodine (I_2) combine to form orange, solid tin iodide with an unknown formula.

$Sn(s) + I_2\,(s) \rightarrow Sn_xI_y\,(s)$

Mass of Sn reacted = 0.455 g

Mass of I_2 reacted = 1.947 g

Calculate die values of x and y (in other words calculate the formula of the compound.

Answer:

$M_{Sn} = 118.71\ g.mol^{-1}$

$M_I = 126.90\ g.mol^{-1}$

$\Rightarrow n_{Sn} = \frac{m}{M} = \frac{0.455\,g}{118.71\,g \cdot mol^{-1}} = 0.00383\ mol\ Sn$

$(3.83 \times 10^{-3}\ mol\ Sn)$

$\Rightarrow$ For mono-atomic Iodine (I)

$\Rightarrow n_I = \frac{m}{M} = \frac{1.947\,g}{126.90\,g \cdot mol^{-1}} = 0.0153\ mol\ I$

$(1.53 \times 10^{-2}\ mol\ I)$

mol Ratio between Sn and I

$\frac{0.00383\ mol\ Sn}{0.00383\ mol\ Sn} : \frac{0.0153\ mol\ I}{0.00383\ mol\ Sn}$

$= 1\ mol\ Sn : \underbrace{3.99\ mol\ I}_{\rightarrow 4I}$

$\therefore$ Formula $= SnI_4$

* You can also do the calculation for di-atomic iodine $\Rightarrow I_2$

$n_{I_2} = \frac{m}{M} = \frac{1.947\,g}{2(126.90\,g \cdot mol^{-1})} = \frac{1.947\,g}{253.8\,g \cdot mol^{-1}}$

$= 0.00767\ mol\ I_2$

Ratio:

$\frac{0.00383\ mol\ Sn}{0.00383\ mol\ Sn} : \frac{0.00767\ mol\ I_2}{0.00383\ mol\ Sn}$

$= 1\ mol\ Sn : 2.00\ mol\ I_2$

$\therefore 2(I_2) = 4I$

$\therefore$ Formula $= SnI_4$

Both methods are accepted.

Try Yourself 2.22

Elemental sulfur (1.256 g) is combined with fluorine, F_2, to give a compound with the formula SF_x, a very stable, colorless gas. If you have isolated 5.722 g of SF_x, what is the value of x?

Answer:

1.256 g S and $M_S = 32.1\ g \cdot mol^{-1}$

5.722 g SF_x isolated

Using the Law of mass conservation we can deduce the following:

5.722 g of SF_x − 1.256 g of S = 4.466 g of F

∴ 4.466 g of Fluorine were used in the reaction.

* mol amount of S: $n = \frac{m}{M} = \frac{1.256\ g}{32.1\ g \cdot mol^{-1}}$ $= 0.0391$ mol S

* Using mono-atomic fluorine (F).

$n_F = \frac{m}{M} = \frac{4.466\ g}{18.99\ g \cdot mol^{-1}} = 0.235$ mol F

Ratio between S and F

$\frac{0.0391\ mol\ S}{0.0391\ mol\ S} : \frac{0.235\ mol\ F}{0.0391}$

1 S : 6 F

⇒ Formula is SF_6 ($x = 6$)

* Using diatomic fluorine (F_2)

$n_{F_2} = \frac{m}{M} = \frac{4.466\ g}{2(18.99\ g \cdot mol^{-1})} = \frac{4.466\ g}{37.98\ g \cdot mol^{-1}}$

$= 0.118$ mol F_2

Ratio between S and F_2

$\frac{0.0391\ S}{0.0391\ S} : \frac{0.118\ F_2}{0.0391} = 1\ S : 3\ F_2$

in total there are 6 F-atoms

∴ SF_6 ($x = 6$)

<u>Try Yourself 2.23</u>

Given $RuCl_3.xH_2O$

If you heat 1.056 g of the hydrated salt and find that only 0.838 g of $RuCl_3$ remains when all of the water has been driven off. Calculate the value of x from this information.

<u>Answer:</u>

You will need the following molar masses: $M_{RuCl3} = 207.42$ g.mol^{-1} ; $M_{H2O} = 18.02$ g.mol^{-1}

<u>Calculate the mol amount of water lost during heating:</u>

Mass of water lost during heating = 1.056 g – 0.838 g = 0.218 g H_2O

$n_{water} = m_{water} / M_{water}$ = 0.218 g / 18.02 g.mol^{-1} = 0.0121 mol H_2O (1.21 x 10^{-2} mol H_2O)

<u>Calculate the mol amount of anhydrous $RuCl_3$ left over after heating:</u>

$n_{RuCl3} = m_{RuCl3} / M_{RuCl3}$ = 0.838 g / 207.42 g.mol^{-1} = 0.00404 mol $RuCl_3$ (4.04 x 10^{-3} mol $RuCl_3$)

<u>Calculate the mol ratio between water and anhydrous $RuCl_3$:</u>

Ratio of Anhydrous $RuCl_3$: H_2O = 0.0121 mol / 0.00404 mol = 2.99 H_2O : 1 $RuCl_3$ = 3 H_2O : 1 $RuCl_3$

(Remember for a ratio calculation you will divide the smallest mol amount into the larger mol amount(s)).

<u>Conclution:</u>

X = 3; therefore the formula is: $RuCl_3.3H_2O$

(Rutinium(III) chloride trihidraat. / Ruthenium(III) chloride trihydrate).