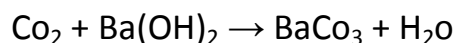


Ex/A 0.2011gm sample of an organic compound was burned in a stream of oxygen , and CO₂ produced in a solution of barium hydroxide.

Calculate the percentage of carbon in the sample if 0.500gm BaCo₃ was formed .



Solution/

$$1 \text{ mol BaCO}_3 \equiv 197.0\text{gm}$$

$$\text{No . mol BaCO}_3 = 0.500 \text{ gm BaCO}_3 \times \frac{1 \text{ mol BaCO}_3}{197 \text{ gm BaCO}_3} = 2.538 \times 10^{-3}$$

mol BaCO₃

$$1 \text{ mol BaCO}_3 \equiv 1 \text{ mol CO}_2 = 2.538 \times 10^{-3} \text{ mol}$$

$$1 \text{ mol CO}_2 \equiv 44.0\text{gm}$$

$$\text{Mass CO}_2 = 2.538 \times 10^{-3} \text{ mol CO}_2 \times \frac{44.0 \text{ gm}}{1 \text{ mol CO}_2}$$

$$= 0.1116 \text{ gm CO}_2$$

$$\text{M CO}_2\% = \frac{m \text{ CO}_2}{m \text{ comp}} \times 100$$

$$\% = \frac{0.1116}{0.2011} \times 100 = 55.49 \%$$

Ex/A 0.3516 gm sample of a commercial phosphate detergent was ignited at a red heat to destroy the organic matter . the residue was then taken up in hot HCl, which converted the P to H₃po₄. The phosphate was precipitated as MgNH₄Po₄.6H₂o by addition of Mg⁺² followed by aqueous NH₃ . after being filtered and washed. The precipitate was converted to Mg₂p₂o₇(222.57gm mol) by ignition at 1000 C⁰ this residue weighed 0.2161gm. Calculate the percent P(30.974gm mol) in the sample.

Solution/

$$\text{No. mol MP} = 0.2161 \text{ gm MP} \times \frac{1 \text{ mol MP}}{222.57 \text{ gm MP}} = 9.71 \times 10^{-4} \text{ mol P}$$

1 mol MP \equiv 2 mol P

No. mol P = 30.774 gm

$$\text{Mass P} = 9.71 \times 10^{-4} \text{ mol} \times \frac{30.974 \text{ gm}}{\text{mol}} = 0.03007 \text{ gm P}$$

$$\text{Percent P\%} = \frac{0.03007}{0.1736} \times 100 = 17.32\% = 17.11\%$$

Ex/A 0.7151 gm sample of impure $\text{Al}_2(\text{CO}_3)_3$ decomposed with HCl. The liberated CO_2 was collected on calcium oxide and found to weight 0.0621 gm the percentage of (Al) in the sample ?

Solution/

1 mol $\text{CO}_2 \approx 44.0 \text{ gm}$

$$\text{No. mol } \text{CO}_2 = 0.0621 \text{ gm } \text{CO}_2 \times \frac{1 \text{ mol } \text{CO}_2}{44.0 \text{ gm } \text{CO}_2} = 1.411 \times 10^{-3} \text{ mol}$$

1 mol $\text{Al}_2(\text{CO}_3)_3 \equiv 3 \text{ mol}(\text{CO}_2)$

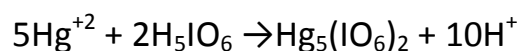
1 mol $\text{Al}_2(\text{CO}_3)_3 \equiv 2 \text{ mol}(\text{Al})$

$$\text{No. mol Al} = 1.411 \times 10^{-3} \text{ mol } \text{CO}_2 \times \frac{2 \text{ mol Al}}{3 \text{ mol } \text{CO}_2} = 0.941 \times 10^{-3} \text{ mol Al}$$

$$\text{Mass. Al} = 0.941 \times 10^{-3} \text{ mol} \times \frac{27 \text{ gm}}{1 \text{ mol}} = 0.0254 \text{ gm Al}$$

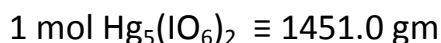
$$\% \text{ Al} = \frac{0.0254 \text{ gm}}{0.7151 \text{ gm}} \times 100 = 3.56\%$$

Ex / the mercury in a 0.8142 gm sample was precipitated with an excess of paraper iodic acid H_2IO_6



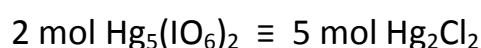
The precipitate was filtered free of precipitating agent, dried and weighed 0.4114 gm was recovered. Calculate the percentage of Hg_2Cl_2 in the sample.

Solution /



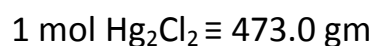
$$\text{No. mol Hg}_5(\text{IO}_6)_2 = 0.4114 \text{ gm Hg}_5(\text{IO}_6)_2 \times \frac{1 \text{ mol Hg}_5(\text{IO}_6)_2}{1451 \text{ gm Hg}_5(\text{IO}_6)_2}$$

$$= 2.84 \times 10^{-4} \text{ mol}$$



$$\text{No. mol Hg}_2\text{Cl}_2 = 2.84 \times 10^{-4} \text{ mol Hg}_5(\text{IO}_6)_2 \times \frac{5 \text{ mol Hg}_2\text{Cl}_2}{2 \text{ mol Hg}_5(\text{IO}_6)_2}$$

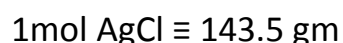
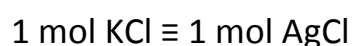
$$= 7.10 \times 10^{-4} \text{ mol}$$



$$\text{Mass of Hg}_2\text{Cl}_2 = 7.10 \times 10^{-4} \text{ mol} \times 473.0 \text{ gm / mol} = 0.3358 \text{ gm}$$

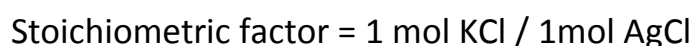
$$\% \text{ Hg}_2\text{Cl}_2 = \frac{0.3358 \text{ gm}}{0.8142 \text{ gm}} \times 100 = 41.24 \%$$

Ex / treatment of a 0.2500 gm sample of impure potassium chloride with an excess of AgNO_3 resulted in the formation of 0.2191 gm of AgCl . Calculate the percentage of KCl in the sample.



$$\text{No. mol AgCl} = 0.2191 \text{ gm AgCl} \times 1 \text{ mol AgCl} / 143.5 \text{ gm AgCl}$$

$$= 1.53 \times 10^{-3} \text{ mol AgCl}$$



$$\text{No. mol KCl} = 1.53 \times 10^{-3} \text{ mol AgCl} \times \frac{1 \text{ mol KCl}}{1 \text{ mol AgCl}} = 1.53 \times 10^{-3} \text{ mol KCl}$$

$$1 \text{ mol KCl} \equiv 74.5 \text{ gm KCl}$$

$$\text{Mass KCl} = 1.53 \times 10^{-3} \text{ mol KCl} \times 74.5 \text{ gm / mol KCl} = 0.114 \text{ gm KCl}$$

$$\% \text{ KCl} = 0.114 \text{ gm} / 0.2500 \text{ gm} \times 100 = 45.6 \%$$

Ex/ A sample of impure magnetite , Fe_3O_4 , weighing 1.542 gm is dissolved ; the iron is oxidized to Fe^+ and precipitated as $\text{Fe}(\text{OH})_3$. The precipitate is ignited to Fe_2O_3 , giving a weight of 1.485 gm .Calculate the percentage of Fe_3O_4 in the sample .

$$\text{No. mol Fe}_2\text{O}_3 = 1.485 \text{ gm Fe}_2\text{O}_3 \times \frac{\text{mol Fe}_2\text{O}_3}{159.7 \text{ gm Fe}_2\text{O}_3}$$

$$= 0.00930 \text{ mol Fe}_2\text{O}_3$$

This is equivalent to 2 / 3 as many moles of Fe_3O_4 , since 2 moles Fe_3O_4 will yield 3 moles Fe_2O_3 when oxidized.

$$2 \text{ moles Fe}_3\text{O}_4 \equiv 3 \text{ moles Fe}_2\text{O}_3$$

$$\text{No. mol Fe}_3\text{O}_4 = 0.00930 \text{ mole Fe}_2\text{O}_3 \times \frac{2 \text{ mole Fe}_3\text{O}_4}{3 \text{ mole Fe}_2\text{O}_3} = 0.00620 \text{ moles Fe}_3\text{O}_4$$

$$\text{Weight Fe}_3\text{O}_4 \text{ in sample} = 0.00620 \text{ mole} \times 231.55 \text{ gm / mole}$$

$$= 1.437 \text{ gm Fe}_3\text{O}_4$$

$$\text{Percentage Fe}_3\text{O}_4 = 1.437 \text{ gm} / 1.542 \text{ gm} \times 100 = 93.1 \%$$

Calculations involving concentrations of solutions

- 1- Physical methods :- the simplest ways of expressing the strength of a solution are in terms of the amount of solute present per unit amount of solvent or solution. Such methods are known as physical methods because they are based only on physical measurements of weight or volume and do not take into account the chemical reactions of the solute.

The more widely used physical methods are :

- 1- Grams solute per liter (or 1000ml) solution gm/L(solution)
gm/1000ml(solution).
- 2- Grams solute per liter (or 1000ml) solvent.
gm/L(solvent) gm/1000ml (solution).
- 3- Grams solute per unit weight of solution gm/gm (solution)
- 4- Grams solute per unit weight of solvent gm/gm(solvent)

1- Percentage methods.

$$\text{Weight percent (w/w)} = \frac{\text{mass solute}}{\text{mass solution}} \times 100\%$$

$$\text{Volume percent (v/v)} = \frac{\text{volume solute}}{\text{volume solution}} \times 100\%$$

$$\text{Weight / volume percent (w/v)} = \frac{\text{mass solute,g}}{\text{volume solution,ml}} \times 100\%$$

Ex/Express the concentration with weight percent of the solution weighting 25.0gm and contained in 200.0gm of sodium sulphate Na_2SO_4 .
Solution/

$$\begin{aligned} \%(w/w) &= \frac{\text{weight of solute}}{\text{weight of solution}} \times 100 \\ &= \frac{25.0\text{gm solute}}{200.0\text{gm solution}} \times 100 = 12.5\% \end{aligned}$$

Ex/Use the weight percent, calculate the concentration of the solution when 3gm AgNO_3 dissolve in 1L distilled water ?

Solution/1L H_2O → weight H_2O

$$\text{Density} = \frac{\text{weight}}{\text{volume}}$$

$$\text{Weight (H}_2\text{O)} = \text{density} \times \text{volume}$$

$$= \frac{1\text{gm}}{\text{cm}^3} \times 1\text{L} \times \frac{1000\text{ cm}^3}{\text{L}}$$

$$\text{Weight} = \text{weight} + \text{weight}$$

$$\text{(solution)} \quad \text{(solute)} \quad \text{(solvent)}$$

$$= 3\text{gm} + 1000\text{gm} = 1003\text{gm}$$

$$(w/w)\% = \frac{3\text{gm}}{1003\text{gm}} \times 100 = 0.299\% \rightarrow 0.3\%$$

Ex/Use the volume percent, calculate the concentration of the solution is prepared by addition 50.0ml methanol (CH₃OH) to 200.0ml water?

Solution/

$$\text{Volume(solution)} = 50.0\text{ml} + 200.0\text{ml} = 250.0\text{ml}$$

$$V/V \% = \frac{50 \text{ ml}}{250 \text{ ml}} \times 100 = 20.0 \%$$