Gravimetric Analysis Methods

Question 1

The calcium in a 200.0 mL sample of a natural water was determined by precipitating the cation as calcium oxalate (CaC_2O_4). The precipitate was filtered, washed, and ignited in a crucible with an empty mass of 26.6002 g. The mass of the crucible plus calcium oxide (CaO) (56.077 g/mol) was 26.7134 g. Calculate the concentration of Ca (40.078 g/mol) in water in units of grams per 100 mL of the water.

$$\underset{\textbf{calcium oxalate}}{CaC_2O_4} \rightarrow \underset{\textbf{calcium carbonate}}{CaCO_3(s)} + CO(g)$$

$$CaCO_3 \rightarrow \underset{\textbf{calcium oxide}}{CaO(s)} + CO_2(g)$$

Solution:

Given

$V_{sample} = 200$ ml

$m_{empty\ crucible} = 26.002$ g

$m_{crucible + CaO} = 26.7134$ g

unknown: C_{Ca} = g/100 ml H_2O

M_r CaO = 56.077 g/mol

A_r Ca = 40.078

① Calculate mass of CaO

$m_{CaO} = m_{crucible + CaO} - m_{empty}$

$= 26.7134\ g - 26.6002\ g = 0.1132\ g$

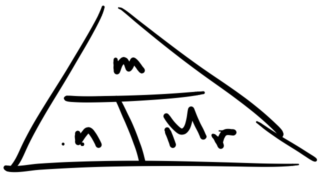
② Calculate no. of moles of CaO

$n = \frac{m}{M_r} = \frac{0.1132}{56.077} = 0.002018$ mol.

③ Calculat no. of moles of Ca in CaC_2O_4

* $n_{CaO} = n_{CaCO_3} = n_{CaC_2O_4} = n_{Ca} = 0.002018$ mol

④ Calculat mass of Ca



$$n = \frac{m}{Mr}$$

$$m = n \times Mr$$

$$= 0.002018 \times 40.078 = 0.0809 \text{ g}$$

→ in 200 ml H_2O

⑤ find Conc. of Ca in 100 ml H_2O

*

H_2O	Ca
	0.0809 g
200 ml	
100 ml	?

$$C_{Ca} (g/100\,ml) = \frac{0.0809 \times 100}{200} = 0.04045 \text{ g/100 ml } H_2O$$

Question 2

An iron ore was analyzed by dissolving a 1.1324 g sample in concentrated HCl. The resul6ng solution was diluted with water, and the iron(III) was precipitated as the hydrous oxide $Fe_2O_3 \cdot xH_2O$ by the addition of NH_3. After filtration and washing, the residue was ignited at a high temperature to give 0.5394 g of pure Fe_2O_3 (159.69 g/mol).

Calculate **(a)** the % Fe (55.847 g/mol) and **(b)** the % Fe_3O_4 (231.54 g/mol) in the sample.

$$3Fe_2O_3(s) \rightarrow 2Fe_3O_4(s) + 1/2O_2(g)$$

Solution:

Given

$m_{Fe\text{-}ore} = 1.324$ g

$m_{Fe_2O_3} = 0.5394$ g

$M_{r\,Fe} = 55.847$ g/mol

$M_{r\,Fe_2O_3} = 159.69$ g/mol

$M_{r\,Fe_3O_4} = 231.54$ g/mol

$m_{Fe} = ?$

(a)

① Calculate no. of moles of Fe_2O_3

$$n = \frac{m}{M_r} = \frac{0.5394}{159.69} = 0.003378 \text{ mol}$$

② Calculate no. of moles of Fe from Fe_2O_3

$$Fe_2O_3 \rightarrow 2Fe$$

1 — 2

0.003378 — ?

$$n_{Fe} = 2 \times 0.003378 = 0.006756 \text{ mol}$$

③ Calculate mass of Fe

$$m_{Fe} = n_{Fe} \times M_r$$

$$= 0.006756 \times 55.847 = 0.3773 \text{ g Fe}$$

④ Calculate % Fe in sample

$$\% Fe = \frac{m_{Fe}}{m_{sample}} \times 100 = \frac{0.3773}{1.1324} \times 100 = \underline{33.32\,\%}$$

(b)

⑤ calculate no. of moles of Fe_3O_4

ratio from chem. eqn. →

Fe_2O_3	Fe_3O_4
3	2
0.003378	?

$$n_{Fe_3O_4} = \frac{2 \times 0.003378}{3} = 0.002252 \text{ mol.}$$

⑥ Calculate mass of Fe_3O_4

$$m_{Fe_3O_4} = n \times M_r = 0.002252 \times 231.54$$

$$= 0.5213 \text{ g}$$

⑦ Finding % Fe_3O_4 in sample

$$\% Fe_3O_4 = \frac{m_{Fe_3O_4}}{m_{sample}} \times 100 = \frac{0.5213}{1.1324} \times 100 = \underline{\underline{46.04\,\%}}$$

<u>Question 3</u>

Treatment of a 0.2500 g sample of impure potassium chloride (KCl) with an excess of silver nitrate ($AgNO_3$) resulted in the forma6on of 0.2912 g of silver chloride (AgCl). Calculate the percentage of KCl in the sample.
(M_W AgCI = 143.42 g/mol; M_W KCI = 74.55 g/mol)

$$KCI + AgNO_3 \rightarrow AgCI\ (s) + KNO_3$$

silver chloride

Solution:

$m_{sample} = 0.2500$ g M_r AgCl = 143.42 g/mol

$m_{AgCl} = 0.2912$ g M_r KCl = 74.55 g/mol

% KCl = ?

① Calculate no. of moles of AgCl

$$n = \frac{m}{M_r} = \frac{0.2912}{143.42} = 0.002030 \text{ mol}$$

② Calculate no. of moles of KCl

$$n_{AgCl} = n_{KCl} = 0.002030 \text{ mol}$$

Same ratio in chem. equ.

③ Calculate mass of KCl

$$m = n \times M_r = 0.002030 \times 74.55 = 0.1513 \text{ g}$$

$$\% KCl = \frac{m_{KCl}}{m_{sample}} \times 100 = \frac{0.1513}{0.2500} \times 100 = \underline{\underline{60.52\%}}$$

Question 4

What mass of copper iodate [$Cu(IO_3)_2$] can be formed from 0.50 g of $CuSO_4 \cdot 5H_2O$?

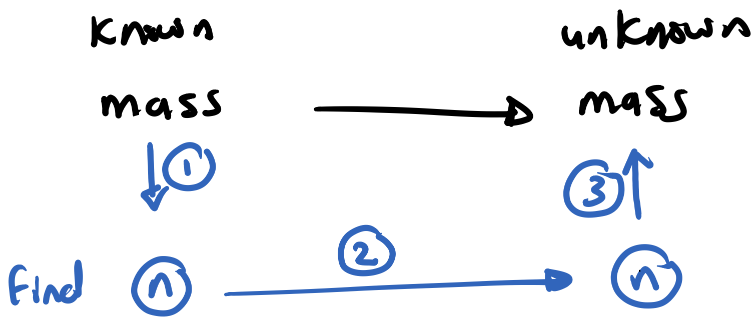
(M_W $CuSO_4.5H_2O$ = 249.67 g/mol; M_W $Cu(IO_3)_2$ = 413.35 g/mol)

$$^{1}CuSO_4 + 2KIO_3 \rightarrow K_2SO_4 + {}^{1}Cu(IO_3)_2$$

copper iodate

Solution:

m = 50 g   m = ?



① Calculate no. of moles of $CuSO_4$

$$n = \frac{m}{M_r} = \frac{0.50}{249.67} = 0.002002 \text{ mol}$$

② Calculate no. of moles of $Cu(IO_3)_2$

$$n_{Cu(IO_3)_2} = n_{CuSO_4} = 0.002002 \text{ mol}$$

③ Finding mass of $Cu(IO_3)_2$

$$m = n \times M_r = 0.002002 \times 413.35$$

$$= 0.8272 \text{ g}$$

Question 5

What mass of silver iodide (AgI) can be produced from a 0.512 g sample that assays 20.1% aluminium iodide (AlI_3) ?

(M_w AlI_3 = 407.770 g/mol; M_w AgI = 234.773 g/mol)

$$AlI_3 + 3AgNO_3 \rightarrow 3AgI(s) + Al(NO_3)_3$$

aluminium iodide   silver iodide

Solution:

20.1% → ① m → ② n → ③ n → ④ m = ?

① Finding mass of AlI_3

$\% AlI_3 = 20.1\%$ / sample.

$$\% AlI_3 = \frac{m_{AlI_3}}{m_{sample}} \times 100$$

$$\frac{20.1}{100} = \frac{m_{AlI_3}}{0.512} \times 100 / 100$$

$$0.201 = \frac{m_{AlI_3}}{0.512}$$

$$m_{AlI_3} = \frac{20.1}{100} \times 0.512 =$$

$$m_{AlI_3} = 0.512 \times 0.201$$

$$= 0.1029 \text{ g}$$

② Calculate no. of moles of AlI_3

$$n = \frac{m}{M_r} = \frac{0.1029}{407.770} = 0.0002524 \text{ mol}$$

$$= 2.524 \times 10^{-4} \text{ mol}$$

③ finding no. of moles of AgI

AlI_3	AgI
1	3
0.0002524	?

$$n_{AgI} = 3 \times 0.0002524 = 0.0007572 \text{ mol}$$

④ finding mass of AgI

$$m = n \times M_r = 0.0007572 \times 234.773$$

$$= 0.1777 \text{ g}$$

Question 6

A 0.2121 g sample of an organic compound was burned in a stream of oxygen, and the CO_2 produced was collected in a solution of barium hydroxide. Calculate the percentage of carbon in the sample if 0.6006 g of $BaCO_3$ was formed.

(M_W $BaCO_3$ = 197.34 g/mol; M_W C = 12.011 g/mol)

$$CO_2 + Ba(OH)_2 \rightarrow \underset{\text{barium carbonate}}{BaCO_3(s)} + H_2O$$

Solution:

Question 7

How many grams of CO_2 is evolved from a 1.500-g sample that is 38.0% $MgCO_3$ and 42.0% K_2CO_3 by mass?

(M_W CO_2 = 44.01 g/mol; M_W $MgCO_3$ = 84.31 g/mol; M_W K_2CO_3 = 138.21 g/mol)

$$MgCO_3 \rightarrow MgO + CO_2(g)$$

$$K_2CO_3 \rightarrow K_2O + CO_2(g)$$

Solution:

Question 8

The mercury in a 1.0451 g sample was precipitated with an excess of paraperiodic acid (H_5IO_6):

$$5Hg^{2+} + 2H_5IO_6 \rightarrow Hg_5(IO_6)_2 + 10H^+$$

paraperiodic acid — mercury periodat

The precipitate was filtered, washed free of precipitating agent, dried, and weighed, and 0.4114 g was recovered. Calculate the percentage of mercury chloride (Hg_2Cl_2) in sample.

(M_W $Hg_5(IO_6)_2$ = 1448.75 g/mol; Hg_2CI_2 = 472.09 g/mol)

Solution:

Question 9

Ammoniacal nitrogen can be determined by treatment of the sample with chloroplatinic acid; the product is slightly soluble ammonium chloroplatinate:

$$H_2PtCI_6 + 2NH_4^+ \rightarrow (NH_4)_2PtCI_6 + 2H^+$$

Cloroplatinic acid — Amonium cloroplatinate

The precipitate decomposes on ignition, yielding metallic platinum and gaseous Products:

$$(NH_4)_2PtCI_6 \xrightarrow{\Delta} Pt(s) + 2CI_2(g) + 2NH_3(g) + 2HCI(g)$$

m = 0.4693 — (1) ↓ n — (2) → n — (3) ↑ m = ? — (4) → %

↑ m_{sample}

Calculate the percentage of ammonia in a sample if 0.2115 g gave rise to 0.4693 g of platinum. (M_W NH_3 = 17.0306 g/mol; Pt = 195.08 g/mol)

Solution:

(1) Calculate no. of moles of pt

$$n_{pt} = \frac{m}{M_r} = \frac{0.4693}{195.08} = 0.002405 \text{ mol}$$

(2) Calculate no. of moles of NH_3

pt	NH_3
1	2
0.002405	?

$$n_{NH_3} = 2 \times 0.002405 = 0.004810 \text{ mol}$$

(3) Calculate mass of NH_3

$$m = n \times M_r = 0.004810 \times 17.0306 = 0.08192 \text{ g}$$

④ Calculate % NH_3

$$\% NH_3 = \frac{m_{NH_3}}{m_{sample}} \times 100 = \frac{0.08192}{0.2115} \times 100 = 38.73\%$$

Question 10

A 50.0 mL portion of a solution containing 0.200 g of $BaCl_2.2H_2O$ is mixed with 50.0 mL of a solution containing 0.300 g of sodium iodide [$NaIO_3$]. Assume that the solubility of barium iodide [$Ba(IO_3)_2$] in water is negligibly small and calculate

a) the mass of the precipitated barium iodide [$Ba(IO_3)_2$].

b) the mass of the unreacted compound that remains in solution.

(M_w $BaCI_2.2H_2O$ = 244.26 g/mol; $NaIO_3$ = 197.89 g/mol; $Ba(IO_3)_2$ = 487.13 g/mol)

$$BaCI_2 + 2NaIO_3 \rightarrow Ba(IO_3)_2 + 2NaCI$$

Solution:

m = ✓ m = ✓ m = ?

ⓝ ← → ⓝ

limiting reactant

$BaCl_2$ → n $Ba(IO_3)_2$

ⓝ

$NaIO_3$ → n $Ba(IO_3)_2$. excess

V_{BaCl_2} = 50 ml

m_{BaCl_2} = 0.200 g Mr = 244.26 g/mol

V_{NaIO_3} = 50 ml

m_{NaIO_3} = 0.300 g Mr = 197.89 g/mol

① finding ⓝ for $BaCl_2$ and $NaIO_3$

$$n_{BaCl_2} = \frac{m}{M_r} = \frac{0.200}{244.26} = 0.000819 \text{ mol}$$

$$n_{NaIO_3} = \frac{m}{M_r} = \frac{0.300}{197.89} = 0.001516 \text{ mol}$$

② Determine the limiting reactant

= reacted + unreacted

in excess

$$n_{BaCl_2} = 0.000819 \longrightarrow n_{Ba(IO_3)_2} = 0.000819$$

$$n_{NaIO_3} = 0.001516 \longrightarrow n_{Ba(IO_3)_2} = \frac{0.001516}{2}$$

$$= 0.000758 \; *$$

it will be used in calc.

limiting reactant

$NaIO_3$	$Ba(IO_3)_2$
2	1
0.001516	?

③ Calculate mass of $Ba(IO_3)_2$

$$m = n \times M_r = 0.000758 \times 487.13 = 0.3691 \text{ g}$$

ⓑ

④ Calculate mass of unreacted $BaCl_2$

$BaCl_2$	$NaIO_3$
1	2
?	0.001516

$$n_{BaCl_2\text{-reacted}} = \frac{0.001516}{2} = 0.000758 \text{ mol}$$

$n_{reacted + unreacted} = 0.000819$

$n_{unreacted} = 0.000819 - n_{reacted}$

$= 0.000819 - 0.000758 = 0.000061$ mol

$m_{BaCl_2 - unreacted} = n \times M_r$

$= 0.000061 \times 244.26 = 0.0149$ g