Raffles Institution
Year 6 H2 Chemistry 2025

Tutorial 23: The Periodic Table (II) (Suggested Answers)

Q8a

$ZnCO_3(s) \longrightarrow ZnO(s) + CO_2(g)$

Q8b

The ionic radius of Mg^{2+} (0.065 nm) is smaller than that of Zn^{2+} (0.074 nm). Hence with the same charge, Zn^{2+} has a lower charge density and hence weaker polarising power than Mg^{2+}. As a result, Zn^{2+} distorts the electron cloud of CO_3^{2-} ion in $ZnCO_3$ to a smaller extent. The covalent bonds within the CO_3^{2-} ion are weakened to a smaller extent as compared to that in $MgCO_3$. Therefore, more heat energy is needed to decompose $ZnCO_3$. Hence $ZnCO_3$ will decompose at a higher temperature than $MgCO_3$.

Q9a

$2Mg(IO_3)_2(s) \longrightarrow 2MgO(s) + 2I_2(g) + 5O_2(g)$

Q9b

Determine the given three iodates(V):

Assume one iodate(V) given is $Mg(IO_3)_2$.

Molar mass of $Mg(IO_3)_2$ = 374.1 g mol^{-1}

Amount of $Mg(IO_3)_2 = \frac{2.00}{374.1} = 5.346 \times 10^{-3}$ mol

Amount of MgO formed = 5.346×10^{-3} mol

Molar mass of MgO = 40.3 g mol^{-1}

Mass of MgO formed = $(5.346 \times 10^{-3})(40.3)$

= 0.215 g

Since the mass of the oxide calculated does not correspond to that of either **X**O, **Y**O or **Z**O, $Mg(IO_3)_2$ is not among the three given iodates(V).

Hence the given iodates(V) are $Ca(IO_3)_2$, $Sr(IO_3)_2$ and $Ba(IO_3)_2$.

[Note: Calculation needs to be made for at least one of the iodates(V) and then reasoning can be made to deduce the identities of the given iodates(V).]

Comparison of thermal stability of $Ca(IO_3)_2$, $Sr(IO_3)_2$ and $Ba(IO_3)_2$:

- The cationic radius increases from Ca^{2+} to Sr^{2+} to Ba^{2+}.
- This causes the charge density and polarising power of the cations to decrease from Ca^{2+} to Sr^{2+} to Ba^{2+}.
- Hence the extent of distortion of the electron cloud of the IO_3^- anion decreases from $Ca(IO_3)_2$ to $Sr(IO_3)_2$ to $Ba(IO_3)_2$.
- The extent of weakening of the covalent bonds in IO_3^- ion also decreases from $Ca(IO_3)_2$ to $Sr(IO_3)_2$ to $Ba(IO_3)_2$.
- Hence the amount of heat energy and temperature required for decomposition increases from $Ca(IO_3)_2$ to $Sr(IO_3)_2$ to $Ba(IO_3)_2$.
- Therefore, thermal stability increases in the order: $Ca(IO_3)_2 < Sr(IO_3)_2 < Ba(IO_3)$

From the graphs, thermal stability of the given iodates(V) increases in the following order:
Z$(IO_3)_2$ < **X**$(IO_3)_2$ < **Y**$(IO_3)_2$

Hence,
Z$(IO_3)_2$ is $Ca(IO_3)_2$ $\Rightarrow$ its rate of thermal decomposition is the fastest.

X$(IO_3)_2$ is $Sr(IO_3)_2$ $\Rightarrow$ its rate of thermal decomposition is slower than $Ca(IO_3)_2$.

Y$(IO_3)_2$ is $Ba(IO_3)_2$ $\Rightarrow$ it does not undergo thermal decomposition.

Q10a

$Ca(CH_3CO_2)_2 \rightarrow CaCO_3 + CH_3COCH_3$
solid **M**

Comments:
• The question has clearly stated that propanone is formed. When writing the balanced equation, the structural formula of propanone should be shown, not just the molecular formula (C_3H_6O).

Q10b

The $CaCO_3$ sample (i.e. **M**) undergoes partial decomposition to give a white residue.
This white residue contains both $CaCO_3$ and CaO.

Method 1

Amount of $CaCO_3$ formed at 450 ^oC = Amount of $Ca(CH_3CO_2)_2$ = $\frac{2.0}{158.1}$ = 0.01265 mol

Let the amount of CO_2 evolved = y mol.

Amount/mol	$CaCO_3$	$\longrightarrow$	CaO	+	CO_2
Initial	0.01265		-		-
Change	-y		+y		+y
Final	0.01265 - y		y		y

Mass of $CaCO_3$ in the residue = (0.01265 – y)(100.1) = (1.266 – 100.1y) g
Mass of CaO in the residue = 56.1y g

$$\frac{\text{mass of } CaCO_3 \text{ in the residue}}{\text{mass of CaO in the residue}} = \frac{70}{30}$$

$$\Rightarrow \quad \frac{1.266-100.1y}{56.1\ y} = \frac{7}{3}$$

$$1.266 - 100.1y = 130.9y$$

$$y = 0.005483$$

Amount of CO_2 evolved = 0.005483 mol
Mass of CO_2 evolved = (0.005483)(44.0) = 0.241 g

Method 2
Mass of $CaCO_3$ formed at 450°C = (2/158.1)(100.1) = 1.266 g

	$CaCO_3$	$\longrightarrow$	CaO	+	CO_2
mass after partial decomposition / g	1.266 – x – y		x		y

Since mass of white residue = (1.266 – y) g and % of CaO in white residue = 30%,

$$\frac{x}{1.266 - y} = 0.3$$

Since amount of CaO formed = amount of CO_2 formed,

$$\frac{x}{56.1} = \frac{y}{44}$$

Solving simultaneous equations, y = mass of CO_2 lost = <u>0.241 g</u>

Method 3
Mass of $CaCO_3$ formed at 450°C = (2/158.1)(100.1) = 1.266 g

Let mass of white residue be x g and mass of CO_2 evolved by y g.
Mass of CaO = (0.3x) g

Since $CaCO_3 \rightarrow CaO + CO_2$, amount of CaO = amount of CO_2

$$\frac{0.3x}{56.1} = \frac{y}{44}$$

By conservation of mass, x + y = 1.266 g
Solving simultaneous equations, y = <u>0.241 g</u>

Comments:
- Candidates should show all essential steps with appropriate units.
- Candidates should keep the answers for intermediate steps to 3 or 4 significant figures.
- The question stated that the white residue contains 70% by mass of solid **M**. Many candidates misinterpreted this information given.

  Wrong interpretation:
  Mass of $CaCO_3$ in the white residue = 70% of mass of $CaCO_3$ formed initially
  = 0.70 x 1.266
  Correct interpretation:
  Mass of $CaCO_3$ in the white residue = 70% of mass of white residue
  = 0.70 x (mass of $CaCO_3$ + mass of CaO)

<u>Q10c</u>

$Ba(CH_3CO_2)_2$ will decompose at a <u>higher temperature</u>. Ba^{2+} ion has a larger ionic radius as compared to Ca^{2+} ion resulting in a <u>lower charge density</u> and <u>weaker polarising power</u>. Hence, Ba^{2+} <u>distorts the electron cloud of the $CH_3CO_2^-$ ions to a smaller extent</u> and thus <u>weakens the covalent bonds within $CH_3CO_2^-$ ions to a smaller extent</u> as compared to Ca^{2+} ion. This results in <u>more heat energy</u> and a <u>higher temperature</u> required for decomposition of $Ba(CH_3CO_2)_2$.

Comments:
- Reference must be made to the ion (Ba^{2+}), not atom (Ba).

Q11(a)(i)

Down Group 17, the value of $E^{\ominus}$ for $X_2(g) + 2e^- \rightleftharpoons 2X^-(aq)$ becomes less positive and the position of equilibrium of the reduction of X_2 to X^- lies increasingly to the left. Hence, down the group, X_2 has less tendency to be reduced and the oxidising power of X_2 decreases.

Q11(a)(ii)

Chlorine, being a stronger oxidising agent, oxidises thiosulfate to sulfate ions, where the average oxidation state of S increases from +2 to +6.

Iodine, being a weaker oxidising agent, oxidises thiosulfate to tetrathionate ions, where the average oxidation state of S only increases from +2 to +2.5.

(equations are not required by question)
$S_2O_3^{2-}(aq) + 4Cl_2(aq) + 5H_2O(l) \longrightarrow 2SO_4^{2-}(aq) + 8Cl^-(aq) + 10H^+(aq)$
$2S_2O_3^{2-}(aq) + I_2(aq) \longrightarrow S_4O_6^{2-}(aq) + 2I^-(aq)$

Alternative answer
Chlorine, being a stronger oxidising agent, oxidises Fe to Fe^{3+} (i.e. from Fe to Fe^{2+}, and then from Fe^{2+} to Fe^{3+}), where the oxidation state of Fe increases from 0 to +3.
Iodine, being a weaker oxidising agent, oxidises Fe to Fe^{2+}, where the oxidation state of Fe only increases from 0 to +2.

(equations are not required by question)
$2Fe(s) + 3Cl_2 \longrightarrow 2FeCl_3(s)$
$Fe(s) + I_2 \longrightarrow FeI_2(s)$

Alternative answer
Chlorine, being a stronger oxidising agent, oxidises Br^- to Br_2 / I^- to I_2.

(equations are not required by question)
$Cl_2 + 2I^- \longrightarrow 2Cl^- + I_2$

Note: We can also compare Br_2 and I^-. Furthermore, only displacement reactions allow us to compare the relative oxidising power of Cl_2, Br_2 and I_2. On the other hand, reaction with thiosulfate or Fe only allows us to conclude that Cl_2 and Br_2 are stronger oxidising agent than I_2.

Q11(b)(i)
$2HX(g) \rightarrow H_2(g) + X_2(g)$
(**Reversible arrows are acceptable too since the extent of reaction depends on the type of hydrogen halides.)

Q11(b)(ii)
The thermal stability of the hydrogen halides decreases down Group 17.

Thermal stability is related to the H–X bond strength. The more endothermic bond dissociation energy of the H–X bond, the stronger the H–X bond and that HX is more thermally stable. Down Group 17, there is less effective orbital overlap between the 1s orbital of H and the valence p orbital of the halogen. Also, the difference in electronegativity between H and the halogens also decrease down the group. This leads to a decrease in bond polarity. Hence, the H–X bond strength decreases down group.

Bond dissociation energy decreases in the order:
H–F > H–Cl > H–Br > H–I.

Q12(a)

T : I_2 **U**: Br_2

Q12(b)

From Data Booklet,

$Br_2 + 2e^- \rightleftharpoons 2Br^-$ $E^\ominus = +1.07$ V

$I_2 + 2e^- \rightleftharpoons 2I^-$ $E^\ominus = +0.54$ V

As observed from the less positive $E^\ominus$ values from Br_2 to I_2, the oxidising strength of the halogens decreases down the group.

In the reaction between bromine and thiosulfate,
bromine, a stronger oxidising agent than iodine, is able to oxidise sulfur such that the average oxidation state of sulfur increases from +2 to +6.

$$4Br_2 + S_2O_3^{2-} + 5H_2O \longrightarrow 8Br^- + 2SO_4^{2-} + 10H^+$$

In the reaction between iodine and thiosulfate,
iodine, a weaker oxidising agent, is only able to oxidise sulfur such that the average oxidation state of sulfur increases from +2 to +2.5.

$$I_2 + 2S_2O_3^{2-} \longrightarrow 2I^- + S_4O_6^{2-}$$

Q13a

Reaction 1: Cl_2 and Fe^{2+}

Oxidation: $Fe^{2+} \longrightarrow Fe^{3+} + e^-$

Reduction: $Cl_2 + 2e^- \longrightarrow 2Cl^-$

$2Fe^{2+}(aq) + Cl_2(g) \longrightarrow 2Fe^{3+}(aq) + 2Cl^-(aq)$

[$E^\ominus_{cell} = +0.59V > 0$ (reaction is spontaneous)]

Observation: Pale green solution of Fe^{2+} turns to pale yellow due to the formation of Fe^{3+}.

Reaction 2: NaOH and Fe^{3+}

$Fe^{3+}(aq) + 3OH^-(aq) \longrightarrow Fe(OH)_3(s)$

Observation: A red-brown ppt of $Fe(OH)_3$ is formed, which is insoluble in excess NaOH.

Q13(b)

Reaction 1: I_2 is shaken with KI

$I_2(s) \rightleftharpoons I_2(aq)$

$I_2(aq) + I^-(aq) \rightleftharpoons I_3^-(aq)$ ----------(1)

Observation: Black crystals of I_2 dissolves to form brown solution due to the formation of the complex I_3^-.

Reaction 2: $Na_2S_2O_3$ reacts with I_2

$2S_2O_3^{2-}(aq) + I_2(aq) \longrightarrow S_4O_6^{2-}(aq) + 2I^-(aq)$ ----(2)

From (2), $S_2O_3^{2-}$ reacts with I_2, thus decreasing $[I_2(aq)]$. By Le Chatelier's Principle, position of equilibrium (1) shifts to the left, decreasing $[I_3^-]$, until all I_2 is reacted with $S_2O_3^{2-}$.

Observation: Brown solution is decolourised. (Note: $S_4O_6^{2-}(aq)$ is colourless).