NYJC J2 H2 Chemistry Prelim Answers

Paper 4 Answers

1 (a) (i) Dilution of FA 1

Final burette reading for **FA 1** / cm^3	
Initial burette reading for **FA 1** / cm^3	
Volume of **FA 1** used / cm^3	24.20
	2 d.p.

Titration Results

Final burette reading for **FA 4** / cm^3		
Initial burette reading for **FA 4** / cm^3		
Volume of **FA 4** used / cm^3	19.15	19.05
	2 d.p.	2 d.p.

Correct header and units for both tables [1]
Correct evaluation and precision for both tables [1]

$$\text{scaled mean titre} = \frac{24.25}{\text{volume diluted}} \times \text{mean titre}$$

Scaled mean titre diff $\leq \pm 0.5$ cm^3 [2] or $\leq \pm 1.0$ cm^3 [1]

(ii) Ave titre value = 19.10 (2 d.p.)
Choice of titre values are consistent and average correctly evaluated [1]

(b) (i) $n_{S_2O_3^{2-}}$ reacted = $0.200 \times \frac{19.10}{1000}$ = 3.82×10^{-3} mol

$n_{H_2O_2}$ in 25.0 cm^3 **FA 5** reacted = n_{I_2} formed

= $\frac{1}{2} \times 3.82 \times 10^{-3}$ mol

= 1.91×10^{-3} mol [1]

$$[H_2O_2]_{FA\,5} = \frac{1.91 \times 10^{-3}}{\frac{25.0}{1000}} = 7.64 \times 10^{-2} \text{ mol dm}^{-3} \text{ [1]}$$

(ii) $7.64 \times 10^{-2} \times \frac{250}{1000} = [H_2O_2]_{FA\,1} \times \frac{24.20}{1000}$ ecf for $[H_2O_2]_{FA\,5}$ from 2(b)(i)

$$[H_2O_2]_{FA\,1} = \frac{7.64 \times 10^{-2} \times \frac{250}{1000}}{\frac{24.20}{1000}} = 0.789 \text{ mol dm}^{-3} \text{ [1]}$$

(c) $2H_2O_2 \rightarrow 2H_2O + O_2$

n_{O_2} evolved from 1 dm^3 of **FA 1** $= \frac{1}{2} \times 0.789$

$= 0.394$ mol

n_{O_2} evolved from 1 cm^3 of **FA 1** at r.t.p. $= \frac{0.3946}{1000} = 3.946 \times 10^{-4}$ mol [1]

V_{O_2} evolved from 1 cm^3 of **FA 1** at r.t.p. $= 3.946 \times 10^{-4}$ mol $\times$ 24000 cm^3 mol^{-1}

$= 9.47$ cm^3 [1]

Hence 1 cm^3 of **FA 1** gives 9.47 cm^3 of O_2 $\Rightarrow$ Vol strength of **FA 1** is 9.47

(d) H_2O_2 in **FA 5** is the limiting reagent while **FA 3** is measured in excess. Using a measuring cylinder to measure **FA 3** will not affect the amount of I_2 produced hence will not affect the volume / accuracy of **FA 4** used.

(e) Impact: The oxidation of I^- by H_2O_2 is incomplete when titration is carried out as the oxidation reaction is slow hence the concentration of H_2O_2 calculated is smaller than expected.[1]

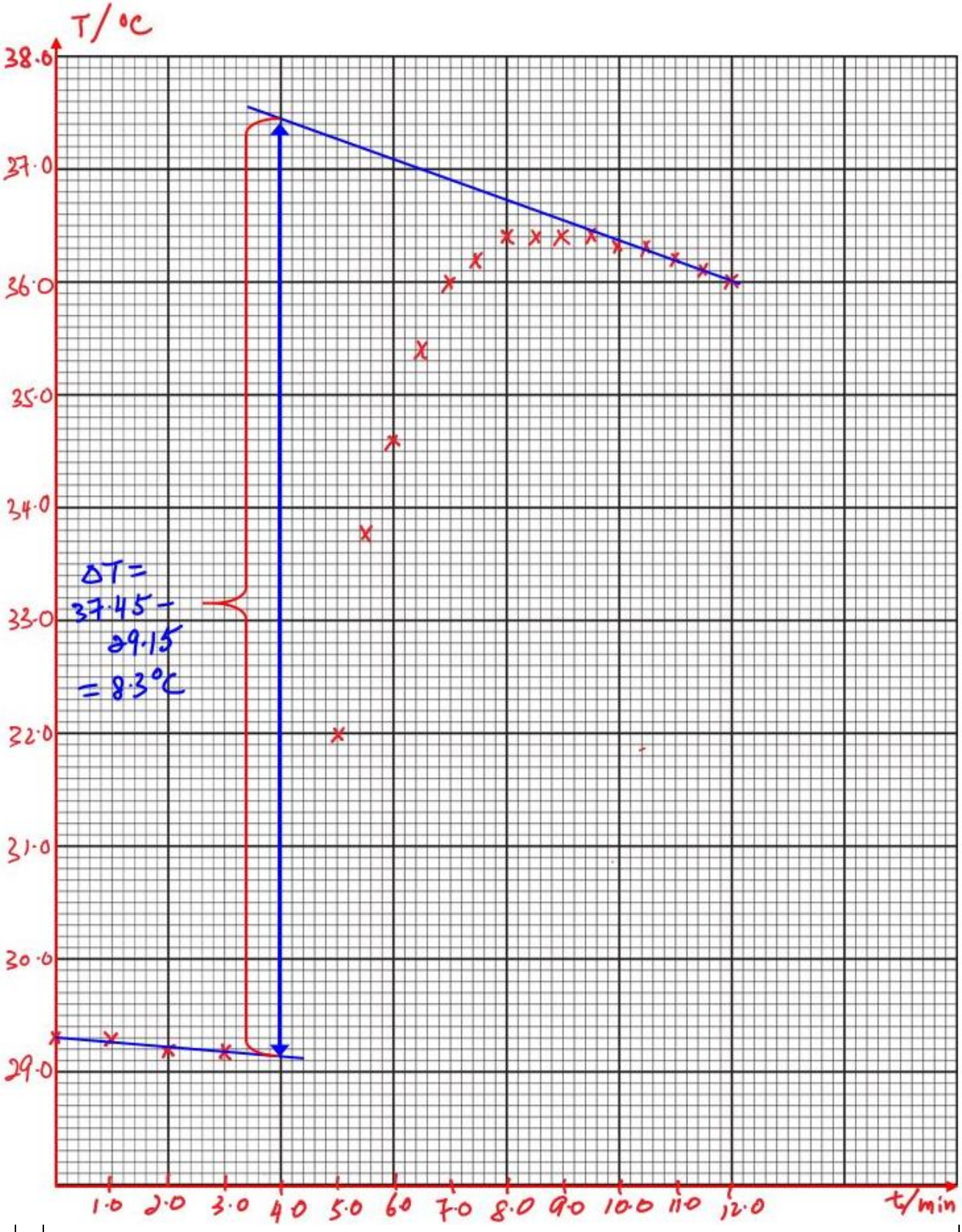
Evidence: The reaction mixture needs to stand for at least 7 minutes before titration. **OR** slow return of the blue black colour at end point suggest oxidation of I^- still continues.[1]

2 (a) Temperature readings

Time / min	Temperature / °C		Time / min	Temperature / °C
0	29.3		7 ½	36.2
1	29.3		8	36.4
2	29.2		8 ½	36.4
3	29.2		9	36.4
4	—		9 ½	36.4
5	32.0		10	36.3
5 ½	33.8		10 ½	36.3
6	34.6		11	36.2
6 ½	35.4		11½	36.1
7	36.0		12	36.0

Correct heading and units [1]

Record time to nearest half minute and temp to 1 d.p except at 4th min.[1]


T/°C
38.0
37.0
36.0
35.0
34.0
33.0
32.0
31.0
30.0
29.0
ΔT =
37.45 −
29.15
= 8.3°C
1.0
2.0
3.0
4.0
5.0
6.0
7.0
8.0
9.0
10.0
11.0
12.0
t/min

Temperature change, ΔT = ... °C [4]

Correctly labelled axes and units, no odd scale used [1]
Correctly plotted all points within ½ sq except at 4th min [1]
Drew and extrapolated 2 best–fit lines to the 4th min, allowing 1 anomalous pt after mixing [1]
Correctly determined and labelled ΔT clearly [1]

(c) q = mcΔT
= 100 x 4.18 x Ans fr **2(b)** J [1] including units

$$\Delta H_{decomp} = -\frac{mc\Delta T}{n_{H_2O_2}}$$

$$= -\frac{100 \times 4.18 \times \text{Ans fr } \mathbf{2(b)}}{\frac{50}{1000} \times 0.950}\ [1]$$

= - 8800 x Ans fr **2(b)** J mol^{-1} [1] with sign and units

(d) **(i)** $2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$

(ii) Iodide is acting as a catalyst. It is reacted in step 1 but regenerated in step 3 hence is not involved in the overall decomposition reaction.[1]

(iii) The enthalpy changes for both experiments will be the same / very similar.[1]
Energy evolved in forming bonds between H_2O_2 and the ions will require the same amount of energy to eventually break it up hence both catalysts are not involved in the overall reaction.[1]

(e) **(i)** The results would be equally reliable. Enthalpy change of reaction, ΔH_{rxn}, and specific heat capacity, *c*, remain the same, while the mass of reactants, m, and the amount of reactants, n, used are half the original quantity hence ΔT would remain the same.[1]

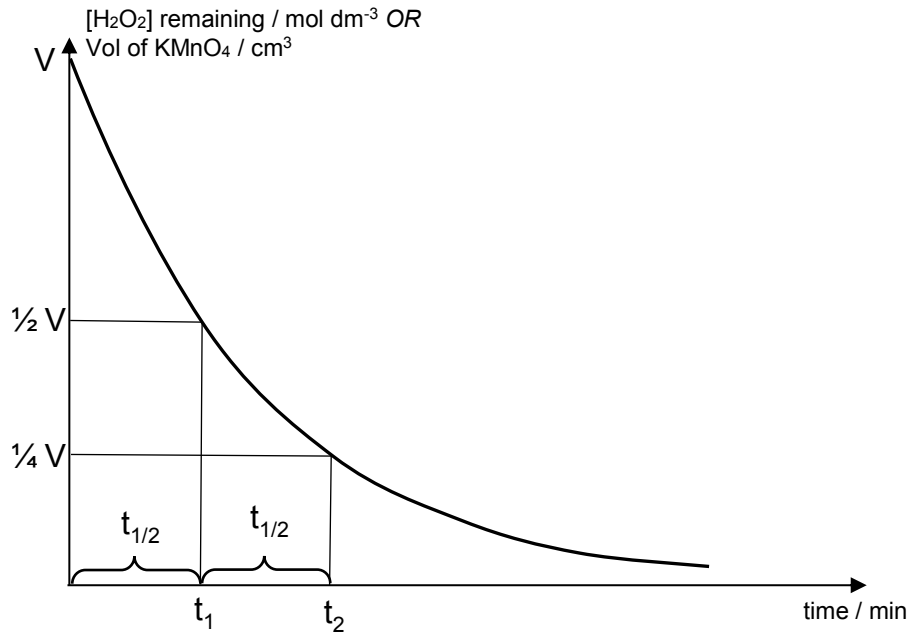
$$\Delta H_{rxn} = \frac{\frac{1}{2}(100) \times 4.18 \times \Delta T}{\frac{1}{2}n}$$

(ii) The ΔT value would still be reliable as cooling happened throughout the experiment at a constant rate hence extrapolating the cooling portion of the graph to the point of mixing would compensate for the heat loss due to cooling.[1]

3 (a) Planning

1. Fill a burette with aq $KMnO_4$ solution.
2. Using a 100 cm^3 measuring cylinder, transfer 100 cm^3 of H_2O_2 into a 250 cm^3 conical flask (allow beaker). [1]
3. Add a small spatula of MnO_2 into the conical flask, swirl the mixture and start the stopwatch at the same time. Allow the mixture to stand. [1]
4. Before 2 minutes, use a 10.0 cm^3 pipette to transfer a sample into a clean 250 cm^3 conical flask. step 1 + step 4 [1]
5. At 2 min, use a clean beaker to transfer 100 (to 150) cm^3 of deionised water into the conical flask (containing the 10 cm^3 sample) and record the timing. [1]
6. Using a 25.0 cm^3 measuring cylinder, add 20.0 (to 25) cm^3 of H_2SO_4 into the conical flask. [1]
7. Titrate the resulting solution against $KMnO_4$ from the burette until 1 drop gives a permanent pink colour. [1]
8. Repeat steps 4 to 7 for another 4 (or 5) times but at time intervals of approximately 4 (to 5) min. [1]
9. Calculate the amount of H_2O_2 remaining at each timing using the volume of $KMnO_4$.

(b)



Sketch labelled graph of V_{KMnO4} against time with units, downward sloping curve [1]

If the decomposition is first order with respect to H_2O_2, the graph would show constant half–life when volume of $KMnO_4$ decreases by half.

Note: V_{KMnO4} reacted α n_{KMnO4} reacted α n_{H2O2} reacted α $[H_2O_2]$

Explains constant half–life and marked out on graph [1]

4 (a) (i)

test 1 To 1 cm^3 of **FA 7** in a test–tube, add about 2–4 cm^3 of H_2SO_4 / HC*l* / HNO_3. If effervescence observed and white ppt is formed when gas evolved is bubbled into limewater, carbonate ions are present.

test 2 To 1 cm^3 of **FA 7** in a test–tube, (add 1 cm^3 of HNO_3 and) add 1 cm^3 of aq $AgNO_3$ followed by aq NH_3 dropwise till in excess. If white ppt is formed and is soluble in excess aq NH_3, Cl^- ions are present. If cream ppt is formed and is partially soluble in excess aq NH_3, Br^- ions are present. If yellow ppt is formed and is insoluble in excess aq NH_3, I^- ions are present.

test 3 To 1 cm^3 of **FA 7** in a test–tube, add 1 cm^3 of aq $BaCl_2$ followed by 2 cm^3 of aq HNO_3 / HC*l*. If white ppt is formed and is insoluble in acid, SO_4^{2-} ions are present.

(ii)

test 1: No effervescence observed

test 2: No ppt observed.

Note: There are no halides present, but **FA 7** contains Fe^{2+}. When aq NH_3 is added, green ppt is formed and is insoluble in excess aq NH_3. (Ppt turns brown upon standing in air.) The presence of $AgNO_3$ and aq NH_3 may also cause a black ppt to form.

test 3: White ppt formed is insoluble in excess acid.

identity of anion: SO_4^{2-}

(iii)

Table 4.1

Test	Observation
Add 2 cm^3 of **FA 7** into a boiling tube followed by aq NaOH dropwise.	Green ppt formed and is insoluble in excess aq NaOH. Ppt turns brown upon standing in air. [1]
Warm the mixture.	Gas evolved turns damp red litmus paper blue.[1]

Cations present in **aqueous FA 7**: ..... NH_4^+ ..... and ..... Fe^{2+} .....[1]

Identity of **solid FA 7**: ..... $(NH_4)_2Fe(SO_4)_2$ .....[1]

(b) (i)

Table 4.2

test	observation	
	FA 8	**FA 9**
1. Add 2 cm^3 of aqueous sulfuric acid and **FA** solution into a test–tube followed by 1 cm^3 of potassium manganate(VII) solution.	Purple $KMnO_4$ decolourised. Effervescence. Gas evolved relights glowing splinter.	No observable change

Warm the test–tube.	No further / observable change	Purple $KMnO_4$ decolourised.
2. Add 1 cm^3 of aqueous sulfuric acid and **FA 8** solution into a test–tube followed by 1 cm^3 of **FA 7**.	Pale green **FA 7** turned (light) yellow / brown.	
Add aqueous ammonia.	Red–brown ppt formed is insoluble in excess $NH_3(aq)$	
3. Add 1 cm^3 of **FA** solution into a test–tube followed by 2 cm^3 of aqueous iodine. Add aqueous sodium hydroxide dropwise until the yellow colour just disappears. Warm the test–tube.	Yellow / colourless solution formed.	(Pale) Yellow ppt formed.
4. Add 1 cm^3 of **FA 9** into a test–tube followed by 1 cm^3 of 2,4–DNPH and warm.		Orange ppt formed
5. To 1 cm^3 of aqueous silver nitrate, add 2 drops of aqueous sodium hydroxide. Add aqueous ammonia dropwise until the precipitate just dissolves. Add 1 cm^3 of **FA 9** to the resulting mixture and warm.		No grey ppt / silver mirror formed.

(ii) **FA 8** is both an oxidising and reducing agent. It reduces $KMnO_4$ in test 1 and oxidises (Fe^{2+} in) **FA 7** in test 2.

(iii)

Structure 1	Structure 2	Structure 3
$CH_3COCH(OH)CH_3$	$CH_3COCH_2CH_2OH$	$CH_3COCH=CH_2$

Any 2 structures with 4 carbon atoms and matches the correct observations [1]