2. (a) (i) What is meant by the *half-life* of a reaction, $t_{½}$?

Time for half a reactant to react ✓

[1]

(ii) $t_{½}$ = 460 ± 10 s ✓

constant half life ✓

evidence on graph to support constant half life (at least two half-lifes shown) ✓

[3]

(iii) no change ✓

[1]

(b) $k = 0.693 / t_{½} = 0.693/460 = 1.51 \times 10^{-3}$ ✓ s^{-1} ✓

for consequential marking: answer should be: 0.693/ans to (a)(ii)

[2]

(c) Rate = $k[C_6H_5N_2Cl(aq)]$ ✓

[1]

(d) (i) After 800s, $[C_6H_5N_2Cl(aq)] = 1.8 \times 10^{-4}$ mol dm^{-3} ✓

(allow any value from 1.7×10^{-4} to 1.8×10^{-4})

[1]

(ii) *Rate* = $k[C_6H_5N_2Cl(aq)] = (1.51 \times 10^{-3}) \times (1.8 \times 10^{-4})$

$= 2.7 \times 10^{-7}$ ✓ mol dm^{-3} s^{-1} ✓

[2]

(iii) measure gradient at t = 800 s ✓

[1]

[Total: 12]

2. **(a)** $k = \dfrac{\text{rate}}{[H_2(g)][NO(g)]^2}$ ✓

$k = 8.3 \times 10^4$ ✓ $dm^6\ mol^{-2}\ s^{-1}$ ✓ calculator value: $8.33333.... \times 10^4$

If [NO] is not squared: $\dfrac{\text{rate}}{[H_2(g)][NO(g)]}$ ×, ans = 250 ✓ units: $dm^3\ mol^{-1}\ s^{-1}$ ✓

If the expression is upside down: $\dfrac{[H_2(g)][NO(g)]^2}{\text{rate}}$ ×, ans = 1.2×10^{-5} ✓ units: $mol^2\ s\ dm^{-6}$ ✓

upside down and not squared: $\dfrac{[H_2(g)][NO(g)]}{\text{rate}}$ ××, ans = $0.004\ mol\ s\ dm^{-3}$ ✓ [3]

(b) **(i)** *effect on rate* x 2 ✓

reason 1st order wrt $H_2(g)$ ✓

[2]

(ii) *effect on rate* x 1/4 ✓

reason 2nd order wrt NO(g) ✓

[2]

(iii) *effect on rate* x 27 ✓

[1]

(c) **(i)** slowest step ✓

[1]

(ii) step 1 (RDS) $H_2(g) + 2NO(g)$ ✓ $\longrightarrow N_2O(g) + H_2O(l)$

step 2 $H_2(g) + N_2O(g) \longrightarrow N_2(g) + H_2O(l)$ rest of equations ✓

[2]

(d) **(i)** NH_3, −3 ✓

NO, +2 ✓

HNO_3 +5 ✓

[3]

(ii) $4NH_3(g) + 5O_2(g) \longrightarrow 4NO(g) + 6H_2O(l)$

products + reactants ⟶ 1 mark; balancing ⟶ 1 mark ✓ ✓

[2]

(iii) molar masses $NH_3 = 17$; $HNO_3 = 63$ ✓

mass = 700 000 x 17/63 = 1.89×10^5 tonnes ✓ calc value $1.888888.... \times 10^5$

ans: mark could be consequential on incorrect molar masses.

[2]

[Total: 18]

2. (a) (i) $m(NH_4NO_3) = 80$ ✓

moles N_2O = moles NH_4NO_3 = 100/80 = 1.25 mol ✓

mass N_2O = 1.25 x (28 + 16) = 55 g ✓

[3]

(ii) nitrogen in NH_4^+: $-3 \longrightarrow +1$ / increases by 4 ✓

nitrogen in NO_3^-: $+5 \longrightarrow +1$ / decreases by 4 ✓

[2]

(b) (i) 1st order has a constant half life ✓

Evidence from graph, either drawn or stated below with 2 half lives ✓

half life approx 52 s✓

[3]

(ii) rate = $k[N_2O(g)]$ ✓

[1]

(iii) evidence of tangent on graph ✓

rate = 0.00524 ✓ mol dm^{-3} s^{-1}

(allow ± 0.005 : i.e. values in range 0.00475 – 0.00575 mol dm^{-3} s^{-1})

[2]

(iv) 0.00524 (ans to (ii)) = k x 0.400

k = 0.0131 ✓ s^{-1} ✓

[2]

(v) rate determining step involves 1 molecule of N_2O ✓

equation shows 2 mol N_2O reacting ✓

[2]

(c) Increases the pressure/rate increases ✓

Gives out heat ✓

Forms oxygen $\longrightarrow$ more efficient combustion ✓

moles of products > moles of reactants ✓

[2 max]

[Total: 17]

1. (a) (i) $Br^-(aq)$ 1st order ✓

$[Br^-(aq)]$ triples rate triples ✓

[2]

$H^+(aq)$ 2nd order ✓

$[H^+(aq)]$ doubles rate quadruples ✓

[2]

$BrO_3^-(aq)$ 1st order ✓

$[BrO_3^-(aq)]$ doubles rate doubles ✓

[2]

(ii) rate = $k[Br^-(aq)]\,[H^+(aq)]^2\,[BrO_3^-(aq)]$ ✓ (state symbols **not** needed)

[1]

(iii)

$$k = \frac{\text{rate}}{[Br^-(aq)]\,[H^+(aq)]^2\,[BrO_3^-(aq)]} = \frac{1.2 \times 10^{-3}}{0.1 \times 0.1^2 \times 0.1} ✓ =$$

rate constant, k:= 12 ✓ units: $dm^9\ mol^{-3}\ s^{-1}$ ✓

(0.0833 would score 1 mark)

[3]

(b) (i) slowest step✓

[1]

(ii) rate equation shows reaction is 1st order wrt HBr and 1st order wrt O_2 ✓

which corresponds to molecules in step 1 ✓

[2]

(iii) $4HBr + O_2 \longrightarrow 2Br_2 + 2H_2O$ ✓

[1]

[Total: 14]