

Exam 2

Review
Solutions

$$1.) \quad q_{H_2O} + q_{Fe} = 0$$

$$m_{H_2O} C_{H_2O} \Delta T_{H_2O} = -m_{Fe} C_{Fe} \Delta T_{Fe}$$

$$1000 (4.184) (88 - 100) = -1800 C_{Fe} (88 - 25)$$

$$C_{Fe} = 0.443 \frac{J}{g^{\circ}C}$$

$$2.) \quad \begin{array}{ll} \text{gains heat} \Rightarrow (+)q & \text{Performs work} \Rightarrow (-)w \\ \text{loses heat} \Rightarrow (-)q & \text{Work on it} \Rightarrow (+)w \end{array}$$

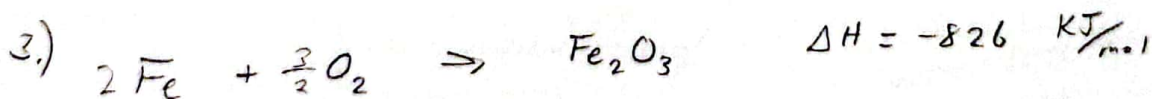
$$\Delta E = q + w$$

if q & w are (+) $\Rightarrow \Delta E$ is (+)

if q & w are (-) $\Rightarrow \Delta E$ is (-)

if one is positive and the other is negative, then their quantities need to be known to determine ΔE

\Rightarrow B is incorrect

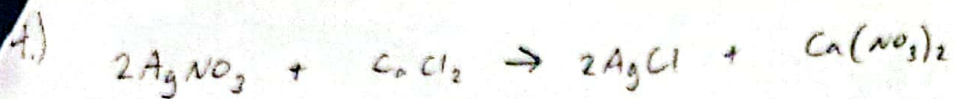


$$Fe: \quad 500g \times \frac{1 \text{ mol}}{55.85g} \times \frac{1}{2} = 4.476 \text{ mol } Fe_2O_3$$

$$O_2: \quad 200g \times \frac{1 \text{ mol}}{32g} \times \frac{1}{3/2} = 4.167 \text{ mol } Fe_2O_3$$

$$q = \Delta H (\text{mol } Fe_2O_3) \\ = -3442 \text{ kJ}$$

"released" implies magnitude $\Rightarrow |q| = 3442 \text{ kJ}$



$$\text{AgNO}_3: (0.05 \text{ L}) (0.2 \frac{\text{mol}}{\text{L}}) \times \frac{2}{2} = 0.01 \text{ mol AgCl}$$

$$\text{CaCl}_2: (0.05 \text{ L}) (0.1 \frac{\text{mol}}{\text{L}}) \times \frac{2}{1} = 0.01 \text{ mol AgCl}$$

$$q_{\text{rxn}} + q_{\text{sol}} = 0$$

$$q_{\text{rxn}} = -m_{\text{sol}} C_{\text{sol}} \Delta T$$

$$= -(50+50)(1.05)(4.20)(100)$$

$$= -441 \text{ J}$$

$$\Delta H = \frac{q_{\text{rxn}}}{\text{mol AgCl}} = -44,100 \text{ J/mol} = \boxed{-44.1 \frac{\text{kJ}}{\text{mol}} = \Delta H_{\text{AgCl}}}$$

$$5.) \quad \times \text{ I: } \text{rate} = k[\text{NO}]^2 \quad \neq \quad k[\text{NO}]^2[\text{O}_2]$$

$$\checkmark \text{ II: } \text{rate} = k[\text{N}_2\text{O}_2][\text{O}_2] \Rightarrow \text{rate} = k[\text{NO}]^2[\text{O}_2] = \checkmark k[\text{NO}]^2[\text{O}_2]$$

NO INTERMEDIATE

$$\text{fast: } [\text{NO}]^2 = [\text{N}_2\text{O}_2]$$

$$\checkmark \text{ III: } \text{rate} = k[\text{NO}]^2[\text{O}_2] = \checkmark k[\text{NO}]^2[\text{O}_2]$$

$$6.) \quad \text{Half life: } \frac{\ln(2)}{k} = t$$

same as double life

$$\Rightarrow \boxed{t = 19.8 \text{ min}}$$

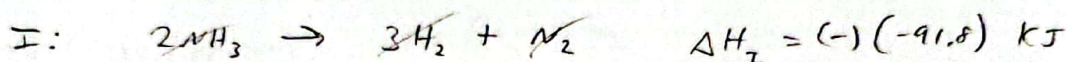
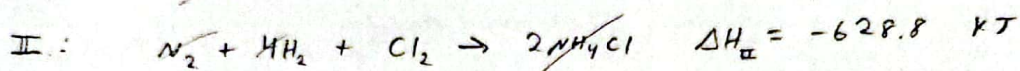
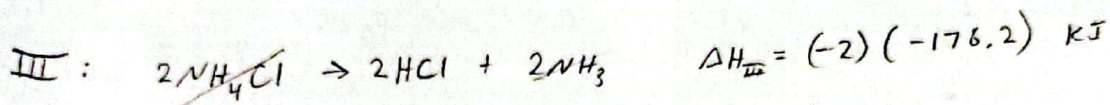
$$7.) \quad q_{\text{H}_2\text{O}} + q_{\text{Fe}} = 0$$

$$q_{\text{H}_2\text{O}} = -q_{\text{Fe}} \Rightarrow (mC\Delta T)_{\text{H}_2\text{O}} = -(mC\Delta T)_{\text{Fe}}$$

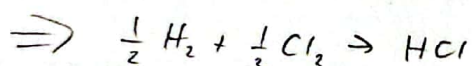
$$150(4.184)(T_f - 20) = -150(0.45)(T_f - 100)$$

$$627.6T_f - 12552 = -67.5T_f + 6750 \Rightarrow \boxed{T_f = 27.8^\circ\text{C}}$$

8.)



$$= -184.6 \text{ kJ}$$



$$\Delta H = \frac{-184.6}{2} \text{ kJ}$$

$$9.) \quad q_{\text{Ag}} + q_{\text{Au}} + q_{\text{H}_2\text{O}} = 0 \quad G + S = 15.3$$

$$S = 15.3 - G$$

$$q_{\text{Ag}} = -q_{\text{Au}} - q_{\text{H}_2\text{O}}$$

$$(m C \Delta T)_{\text{Ag}} = - (m C \Delta T)_{\text{Au}} - (m C \Delta T)_{\text{H}_2\text{O}}$$

$$(15.3 - m_{\text{Au}})(0.235)(22.9 - 62.1) = -m_{\text{Au}}(0.128)(22.9 - 62.1) - 13.1(4.184)(22.9 - 20.9)$$

$$-140.94 + 9.212 m_{\text{Au}} = 5.018 m_{\text{Au}} - 109.621$$

$$4.194 m_{\text{Au}} = 31.32$$

$$m_{\text{Au}} = 7.5 \text{ g}$$

$$10.) \quad q_{\text{rxn}} + q_{\text{cal}} = 0$$

$$10 \text{ g sucrose} \times \frac{1 \text{ mole}}{342.3 \text{ g}} = 0.02921 \text{ mol}$$

$$q_{\text{rxn}} = -q_{\text{cal}}$$

$$= -C_{\text{cal}} \Delta T$$

$$= -165 \text{ kJ}$$

$$\Delta H_{\text{C}_{12}\text{H}_{22}\text{O}_{11}} = \frac{q_{\text{rxn}}}{0.02921} = \boxed{-5648 \frac{\text{kJ}}{\text{mol}}}$$

11.)

Substitute

$$\text{Rate} = k [\text{NO}^+] [\text{NH}_3]$$

Both intermediates

$$\text{II: } [\text{NH}_4^+] = [\text{NH}_3] [\text{H}^+]$$

$$[\text{NH}_3] = [\text{NH}_4^+] [\text{H}^+]^{-1}$$

$$\text{I: } [\text{HNO}_2] [\text{H}^+] = [\text{H}_2\text{O}] [\text{NO}^+]$$

$$[\text{NO}^+] = [\text{HNO}_2] [\text{H}^+] [\text{H}_2\text{O}]^{-1}$$

$$\text{Rate} = k \frac{[\text{NH}_4^+] [\text{HNO}_2] [\text{H}^+]}{[\text{H}_2\text{O}] [\text{H}^+]} = k \frac{[\text{NH}_4^+] [\text{HNO}_2]}{[\text{H}_2\text{O}]} = k [\text{NH}_4^+] [\text{HNO}_2] [\text{H}_2\text{O}]^{-1}$$

$$12) E = \frac{hc}{\lambda} = 2.93 \times 10^{-19} \frac{\text{J}}{\text{photon}}$$

$$h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$$

$$c = 3 \times 10^8 \text{ m/s}$$

$$\lambda = 679 \times 10^{-9} \text{ m}$$

$$\frac{\text{Photons}}{\text{pulse}} = \frac{0.528 \text{ J}}{1 \text{ pulse}} \times \frac{1 \text{ photon}}{2.93 \times 10^{-19} \text{ J}}$$

$$= 1.8 \times 10^{18} \frac{\text{photon}}{\text{pulse}}$$

$$13) \text{Rate} = k [\text{BrO}_3^-]^x [\text{Br}^-]^y [\text{H}^+]^z$$

$$\Rightarrow \text{Find } x: \frac{\text{Rate}_1}{\text{Rate}_2} = \frac{k [\text{BrO}_3^-]^x [\text{Br}^-]^y [\text{H}^+]^z}{k [\text{BrO}_3^-]^x [\text{Br}^-]^y [\text{H}^+]^z}$$

$$\frac{1.126 \times 10^{-2}}{2.251 \times 10^{-2}} = \left(\frac{0.175}{0.35} \right)^x \Rightarrow 0.5 = 0.5^x \Rightarrow x=1$$

$$\Rightarrow \text{Find } y: \frac{\text{Rate}_1}{\text{Rate}_3} = \frac{k [\text{BrO}_3^-]^x [\text{Br}^-]^y [\text{H}^+]^z}{k [\text{BrO}_3^-]^x [\text{Br}^-]^y [\text{H}^+]^z}$$

$$\frac{1.126 \times 10^{-2}}{3.376 \times 10^{-2}} = \left(\frac{0.175}{0.525} \right)^y \Rightarrow 0.333 = 0.333^y \Rightarrow y=1$$

⇒ Find z:

$$\frac{\text{Rate}_2}{\text{Rate}_4} = \frac{k [\text{BrO}_3^-]_2^x [\text{Br}^-]_2^y [\text{H}^+]_2^z}{k [\text{BrO}_3^-]_4^x [\text{Br}^-]_4^y [\text{H}^+]_4^z}$$

$$\frac{2.256 \times 10^{-2}}{5.084 \times 10^{-2}} = \left(\frac{0.175}{0.263} \right)^z \Rightarrow$$

$$0.4437 = 0.6654^z$$

$$\ln(0.4437) = \ln(0.6654^z)$$

$$\ln(0.4437) = z \ln(0.6654)$$

$$\Rightarrow z = 2$$

$$\boxed{\text{Rate} = k [\text{BrO}_3^-] [\text{Br}^-] [\text{H}^+]^2}$$