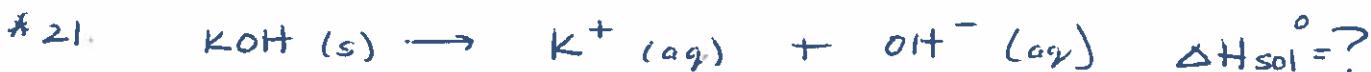


Q #21 - 30 Pg. 305



system

$$m = 0.648g$$

$$M = 56.105 \frac{g}{mol}$$

$$n = 0.648g \times \frac{1mol}{56.105g}$$
$$= 0.011550 mol$$

surroundings

$$Q = mc\Delta T$$

$$= (40g)(4.19 \text{ J/g}^\circ\text{C})(27.8 - 22.6^\circ\text{C})$$

$$= (40g)(4.19 \text{ J/g}^\circ\text{C})(5.2^\circ\text{C})$$

$$= 871.52 \text{ J}$$

$$= 0.87152 \text{ kJ}$$

$$\Delta H_{syst} = -\Delta H_{surr}$$

$$\Delta H_{syst} = -0.87152 \text{ kJ}$$

$$n \Delta H_{sol} = -0.87152 \text{ kJ}$$

$$\Delta H_{sol} = \frac{-0.87152 \text{ kJ}}{0.011550 \text{ mol}}$$
$$= -75.5 \text{ kJ/mol}$$

\therefore the molar enthalpy of solution is -75.5 kJ/mol



$$V = 80.00 \text{ mL} \quad m = 5.022 \text{ g}$$

$$\Delta T = 28.4 - 18.6 \quad M = 84.006 \frac{g}{mol}$$
$$= 9.8^\circ\text{C}$$

$$Q = mc\Delta T$$
$$= (80g)(4.19 \text{ J/g}^\circ\text{C})(9.8^\circ\text{C})$$
$$= 3284.96 \text{ J}$$
$$= 3.28496 \text{ kJ}$$

$$n = 5.022g \times \frac{1mol}{84.006g}$$
$$= 0.05978 \text{ mol}$$

$$n \Delta H_r = -\Delta H_{surr}$$

$$\Delta H_r = \frac{-3.28496 \text{ kJ}}{0.05978 \text{ mol}}$$
$$= -55.0 \text{ kJ/mol}$$

\therefore the molar enthalpy of reaction is -55.0 kJ/mol



$$m = 0.37\text{ g}$$

$$M = 22.990 \frac{\text{g}}{\text{mol}}$$

$$n = 0.37\text{ g} \times \frac{1\text{ mol}}{22.990\text{ g}} \\ = 0.0160940\text{ mol}$$

$$Q = mc\Delta T \\ = (175\text{ g})(4.19\text{ J/g}^\circ\text{C})(6.4^\circ\text{C}) \\ = 4692.8\text{ J} \\ = 4.6928\text{ kJ}$$

$$n\Delta H_r = -\Delta H_{\text{surr}}$$

$$\Delta H_r = \frac{-4.6928\text{ kJ}}{0.0160940\text{ mol}}$$

$$= -290\text{ kJ/mol or } -2.9 \times 10^2\text{ kJ/mol}$$

\therefore the molar enthalpy of reaction is $-2.9 \times 10^2\text{ kJ/mol}$.



$$V = 0.250\text{ L}$$

$$V = 0.150\text{ L}$$

$$c = 0.120\text{ mol/L}$$

$$c = 0.200 \frac{\text{mol}}{\text{L}}$$

$$n = V \times c$$

$$= 0.250\text{ L} \times 0.120 \frac{\text{mol}}{\text{L}}$$

$$n = \boxed{0.03\text{ mol}}$$

1

$$= 0.03$$

$$n = V \times c$$

$$= 0.150\text{ L} \times 0.200 \frac{\text{mol}}{\text{L}}$$

$$= \boxed{0.03\text{ mol}}$$

1

$$= 0.03$$

no LR

$$\Delta T = 20.49^\circ\text{C} - 20.00^\circ\text{C} = 0.49^\circ\text{C}$$

$$m = \text{total } V$$

$$= 250\text{ mL} + 150\text{ mL} = 400\text{ mL} \\ = 400\text{ g}$$

$$Q = mc\Delta T$$

$$= (400\text{ g})(4.19\text{ J/g}^\circ\text{C})(0.49^\circ\text{C})$$

$$= 821.24\text{ J}$$

$$= 0.82124\text{ kJ}$$

$$\Delta H_r = \frac{-\Delta H_{\text{surr}}}{n}$$

$$= \frac{-0.82124\text{ kJ}}{-0.03\text{ mol}} = -27.4\text{ kJ/mol}$$

\therefore the molar enthalpy of rxn is -27.4 kJ/mol



$$V = 0.1 \text{ L}$$

$$V = 0.2 \text{ L}$$

$$c = 0.2 \frac{\text{mol}}{\text{L}}$$

$$c = 0.2 \frac{\text{mol}}{\text{L}}$$

$$n = 0.1 \text{ L} \times 0.2 \frac{\text{mol}}{\text{L}}$$

$$\boxed{= 0.02 \text{ mol}}$$

1

0.02

↑ smaller #
LR

$$n = 0.2 \text{ L} \times 0.2 \frac{\text{mol}}{\text{L}}$$

$$= \underline{\underline{0.04 \text{ mol}}}$$

1

0.04



$$\Delta H_r = -53.6 \frac{\text{KJ}}{\text{mol}}$$

$$= -53600 \frac{\text{J}}{\text{mol}}$$

$$n \Delta H_r = -\Delta H_{\text{surr}}$$

$$\Delta H_{\text{surr}} = -(0.02 \text{ mol})(-53600 \text{ J/mol}) \\ = 1072 \text{ J}$$

$$= mc\Delta T$$

$$1072 \text{ J} = mc\Delta T$$

$$\Delta T = \underline{\underline{1072 \text{ J}}}$$

$$= \frac{mc}{(300 \text{ g})(4.19 \text{ J/g°C})} \\ = 0.853 \text{ °C}$$

$$m = 100 \text{ mL} + 200 \text{ mL} \\ = 300 \text{ mL} \\ = 300 \text{ g}$$

∴ the temperature change is 0.853 °C



$$V = 0.15 \text{ L}$$

$$V = 0.15$$

$$c = 1 \text{ mol/L}$$

$$c = 1 \text{ mol/L}$$

$$n = 0.15 \text{ mol}$$

$$n = 0.15 \text{ mol}$$

$$\Delta T = 30 - 25 = 5 \text{ °C}$$

$$m = 150 + 150 = 300 \text{ mL} \\ = 300 \text{ g}$$

$$\Delta H_r = \frac{-\Delta H_{\text{surr}}}{n} \\ = \frac{-6285 \text{ KJ}}{0.15 \text{ mol}} \\ = -41.9 \text{ KJ/mol}$$

$$Q = mc\Delta T \\ = (300 \text{ g})(4.19 \text{ J/g°C})(5 \text{ °C}) \\ = 6285 \text{ J} \\ = 6.285 \text{ KJ}$$

∴ the molar enthalpy of reaction



$$V = 0.25 \text{ L}$$

$$c = 0.1 \text{ mol/L}$$

$$Q = mc\Delta T$$

$$n = 0.25 \text{ L} \times 0.1 \frac{\text{mol}}{\text{L}}$$

$$= 0.025 \text{ m.s}^{-1}$$

$$\Delta H_{\text{system}} = -\Delta H_{\text{surr}} (q_{\text{surr}})$$

$$n \Delta H_{sol} = -q_{surr}$$

$$(0.025 \text{ mol})(-55.0 \text{ kJ/mol}) = - m c \Delta T$$

$$\Delta T = \frac{(0.025 \text{ mol})(55.0 \text{ kJ/mol})}{(250 \text{ g})(4.19 \text{ J/g}\cdot\text{C})} \times 1000$$

∴ the temperature increases by 1.31°C



$$y = 0.12 L$$

$$V = 0.16 \text{ L}$$

$$c = 0.5 \frac{\text{mol}}{\text{L}}$$

$$c = 0.375 \frac{\text{mol}}{\text{L}}$$

$$\Delta H_f = -53.1 \text{ kJ/mol}$$

$$= -53 \text{ } 100 \text{ J/mol}$$

$$n = 0.12 \text{ L} \times 0.5 \frac{\text{mol}}{\text{L}}$$

$$n = 0.06 \text{ mol}$$

$$= \boxed{0.06 \text{ mol}}$$

200

0.06

$$\Delta H_{\text{system}} = (0.06 \text{ mol}) (53100 \text{ J/mol}) \\ = -3186 \text{ J}$$

$$\begin{aligned} m &= 120 \text{ mL} + 160 \text{ mL} \\ &= 280 \text{ mL} \\ &= 280 \text{ g} \end{aligned}$$

$$= 3186 \text{ J} = - mc\Delta T$$

$$\Delta T = \frac{3186 \text{ J}}{m_C}$$

$$T_F - T_I = \frac{3186 \text{ J}}{mc}$$

$$T_i = T_f - \frac{3186 \text{ J}}{(280 \text{ g})(4.19 \text{ J/g}^\circ\text{C})}$$

$$= 24.5^\circ\text{C} - 27.156^\circ\text{C}$$

$$= 24.5^{\circ}C = 27156^{\circ}C$$

∴ the initial temperature of the solution is 21.8°C



$$m = 7.800 \text{ g}$$

$$\Delta H^\circ = -285.0 \text{ kJ}$$

$$M = 77.978 \frac{\text{g}}{\text{mol}}$$

$$\begin{aligned} \Delta H_{\text{sys}} &= 7.800 \text{ g} \times \frac{1 \text{ mol}}{77.978 \text{ g}} \times \frac{-285.0 \text{ kJ}}{2 \text{ mol}} \xrightarrow{\text{from balanced chem eqn.}} \\ &= -14.2540 \text{ kJ} \\ &= -14254 \text{ J} \end{aligned}$$

$$\Delta H_{\text{sys}} = -\Delta H_{\text{surr}}$$

$$-14254 \text{ J} = -\Delta H_{\text{surr}} = -Q$$

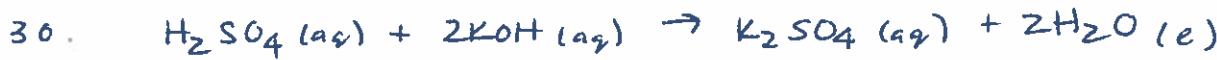
$$-14254 \text{ J} = -mc\Delta T$$

\therefore the temperature change is 30.9°C

$$\Delta T = \frac{14254 \text{ J}}{mc}$$

$$= \frac{14254 \text{ J}}{(110 \text{ g})(4.19 \text{ J/g}^\circ\text{C})}$$

$$= 30.9^\circ\text{C}$$



$$V = 0.2 \text{ L}$$

$$V = 0.2 \text{ L}$$

$$c = 1 \text{ mol/L}$$

$$c = 1 \text{ mol/L}$$

$$n = \frac{0.2 \text{ mol}}{1}$$

$$n = \frac{0.2 \text{ mol}}{2}$$

$$0.1$$

$$\uparrow \text{LR}$$

$$Q = mc\Delta T$$

$$= (400 \text{ g})(4.19 \text{ J/g}^\circ\text{C})(6.50^\circ\text{C})$$

$$= 10894 \text{ J}$$

$$= 10.894 \text{ kJ}$$

$$\Delta H = \frac{-10.894 \text{ kJ}}{0.2 \text{ mol}}$$

$$= -54.5 \text{ kJ/mol}$$

\therefore the molar enthalpy is

$$-54.5 \text{ kJ/mol}$$

