

$$\textcircled{\#1} \quad T_f = 22.2^\circ\text{C}$$

$$-q_{\text{lost}} = q_{\text{gained}}$$

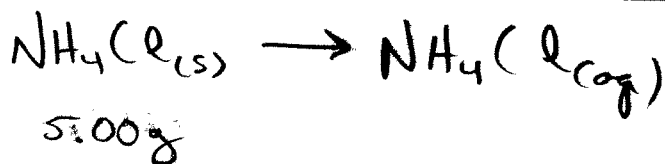
$$-q_{\text{zn}} = q_{\text{H}_2\text{O}} + q_{\text{calorimeter}}$$

$$-(65.45\text{g})(0.388 \text{ J/g}^\circ\text{C})(-52.8^\circ\text{C}) =$$

$$(125\text{g})(4.184 \text{ J/g}^\circ\text{C})(2.3^\circ\text{C}) + H_c(2.3^\circ\text{C})$$

$$13411 \text{ J} = 1203 \text{ J} + H_c(2.3^\circ\text{C})$$

$$\text{A)} \quad H_c = 59.9 \text{ J/}^\circ\text{C} = 60. \text{ J/}^\circ\text{C}$$



$$\frac{5.00\text{g} \mid 1 \text{ mol}}{53.5\text{g}} = 0.09346 \text{ mol}$$

$$-q_{\text{rxn}} = q_{\text{soln}} + q_{\text{calorimeter}}$$

$$-q_{\text{rxn}} = (155\text{g soln})(4.18 \text{ J/g}^\circ\text{C})(-1.8^\circ\text{C}) + (60. \text{ J/}^\circ\text{C})(-1.8^\circ\text{C})$$

$$-q_{\text{rxn}} = -1166 \text{ J} + -108 \text{ J} = -1274 \text{ J}$$

$$\text{B)} \quad q_{\text{rxn}} = 1274 \text{ J}$$

$$\text{C)} \quad \Delta H = \frac{1274 \text{ J}}{0.09346 \text{ mol}} = 13,600 \text{ J/mol} = 13.6 \text{ kJ/mol}$$

2 A 0.712-g sample of magnesium is burned in excess oxygen in a bomb calorimeter with a heat capacity of 722 J/C° containing 350.0 g of water initially at 21.57°C. The temperature of the apparatus rises to 26.32°C

$$\frac{0.712 \text{ g Mg}}{24.31 \text{ g/mol}} = 0.02929 \text{ mol}$$

A) What is the enthalpy of combustion (in kJ/mol) of magnesium?

$$-q_{\text{rxn}} = q_w + q_{\text{calorimeter}}$$

$$-q_{\text{rxn}} = m_w c_w \Delta T_w + H_c \Delta T_w$$

$$-q_{\text{rxn}} = (350.0 \text{ g})(4.184 \text{ J/g}^\circ\text{C})(4.75^\circ\text{C}) + (722 \text{ J/C}^\circ)(4.75^\circ\text{C})$$

$$-q_{\text{rxn}} = 10385 \text{ J} = 10.385 \text{ kJ}$$

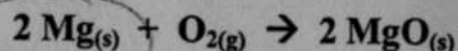
$$q_{\text{rxn}} = -10.385 \text{ kJ}$$

$$\Delta H = \frac{q_{\text{rxn}}}{\text{mol reacting}}$$

$$\Delta H = \frac{-10.385 \text{ kJ}}{0.02929 \text{ mol}}$$

$$\Delta H = -355 \text{ kJ/mol}$$

B) What is the heat of reaction in kJ for the balanced equation:



$$= -355 \text{ kJ/mol Mg} \times 2 \text{ mol Mg reacting}$$

$$\Delta H_{\text{rxn}} = -709 \text{ kJ}$$

#3

$$\frac{0.7000 \text{ L}}{1 \text{ L}} \times \frac{0.500 \text{ mol HCl}}{1 \text{ L}} = 0.350 \text{ mol HCl} \times \frac{1 \text{ mol Ba(OH)}_2}{2 \text{ mol HCl}} = 0.175 \text{ mol Ba(OH)}_2 \text{ needed}$$

$$\frac{0.3000 \text{ L}}{1 \text{ L}} \times \frac{0.500 \text{ mol}}{1 \text{ L}} = 0.150 \text{ mol Ba(OH)}_2$$

L.R.

$$1000 \text{ mL soln} = 1000 \text{ g soln}$$

$$\frac{0.150 \text{ mol Ba(OH)}_2}{1 \text{ mol Ba(OH)}_2} \times \frac{-118 \text{ kJ}}{1 \text{ mol Ba(OH)}_2} = -17.7 \text{ kJ}$$

A) (17,700 J produced) $q_{\text{rxn}} = -17,700 \text{ J}$

$$-q_{\text{rxn}} = q_{\text{soln}} + q_{\text{calorimeter}}$$

$$-q_{\text{rxn}} = m_{\text{soln}} c_{\text{soln}} \Delta T_{\text{soln}} + H_c \Delta T_{\text{soln}}$$

$$-(-17,700 \text{ J}) = (1000 \text{ g})(4.184 \text{ J/g}^\circ\text{C})\Delta T + (180.3 \text{ J/C}^\circ)\Delta T$$

$$17700 \text{ J} = (4364.3 \text{ J/C}^\circ)\Delta T$$

$$\Delta T = 4.06 \text{ C}^\circ = T_f - 25.0^\circ\text{C}$$

B) $T_f = 25.0^\circ\text{C} + 4.06^\circ\text{C} = 29.1^\circ\text{C}$