Name ANSWERS Date _____ Class _____

4. Calculate the amount of heat needed to convert 96 g of ice at −24°C to water at 28°C. The specific heat capacity of $H_2O(s)$ is 2.1 J/g•°C.

## SECTION 17.4 CALCULATING HEATS OF REACTION

1. What is the standard heat of reaction for the combustion of hydrogen sulfide? Refer to Table 17.4 in your textbook.

$$2H_2S(g) + 3O_2(g) \rightarrow 2H_2O(g) + 2SO_2(g)$$

2. Calculate the enthalpy change (in kJ) for the following reaction. State whether the reaction is exothermic or endothermic. Refer to Table 17.4 in your textbook.

$$CaO(s) + CO_2(g) \rightarrow CaCO_3(s)$$

3. What is the enthalpy change for the formation of hydrazine, $N_2H_4(l)$, from its elements?

$$N_2(g) + 2H_2(g) \rightarrow N_2H_4(l)$$

Use the following reactions and enthalpy changes:

$$N_2H_4(l) + O_2(g) \rightarrow N_2(g) + 2H_2O(l) \quad \Delta H = -622.2 \text{ kJ}$$

$$H_2(g) + \tfrac{1}{2}O_2(g) \rightarrow H_2O(l) \quad \Delta H = -285.8 \text{ kJ}$$

1. $\Delta H_f^\circ$ (products)

$$\left(2 \text{ mol } SO_2(g) \times \frac{-296.8 \text{ kJ}}{1 \text{ mol } SO_2(g)}\right) + \left(2 \text{ mol } H_2O(g) \times \frac{-241.8 \text{ kJ}}{1 \text{ mol } H_2O(g)}\right)$$

$$= -593.6 \text{ kJ} + (-483.6 \text{ kJ})$$

$$= -1077 \text{ kJ}$$

$\Delta H_f^\circ$ (reactants)

$$2 \text{ mol } H_2S(g) \times \frac{-20.1 \text{ kJ}}{1 \text{ mol } H_2S(g)} + 0 \text{ kJ}$$

$$= -40.2 \text{ kJ}$$

$$\Delta H_f^\circ = \Delta H_f^\circ(\text{products}) - \Delta H_f^\circ(\text{reactant})$$

$$= -1077 \text{ kJ} - (-40.2 \text{ kJ})$$

$$= 1036.8 \text{ kJ} \quad 1.04 \times 10^3 \text{ kJ}$$

2. $CaO(s) + CO_2(g) \rightarrow CaCO_3(g)$

$-635.5\ KJ/mol \quad -393.5\ KJ/mol \quad -1207.0\ KJ/mol$

$\Delta H_f^\circ = \Delta H_f^\circ(\text{products}) - \Delta H_f^\circ(\text{reactants})$

$= -1207.0\ KJ - (-635.5\ KJ + -393.5\ KJ)$

$= -1207.0\ KJ - (-1029\ KJ)$

$= -178\ KJ$ exothermic

3. $N_2 + 2H_2(g) \rightarrow N_2H_4\,l$

$N_2(g) + \cancel{2H_2O(l)} \rightarrow N_2H_4(l) + \cancel{O_2(g)} \quad \Delta H = 622.2\ KJ$

$2H_2(g) + \cancel{1/2\,O_2(g)} \rightarrow \cancel{2H_2O(l)} \quad \Delta H = -285.8\ KJ \times 2$

$\Delta H = 622.2\ KJ + (-571.6\ KJ)$

$\Delta = 50.6\ KJ$

$N_2(g) + 2H_2(g) \rightarrow N_2H_4(l) \quad \Delta H = 50.6\ KJ$