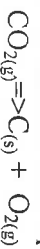
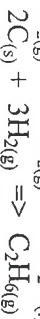


1. Use the thermochemical equations shown below to determine the enthalpy for the following reaction:



$$\Delta H = -214.3 \text{ kJ}$$



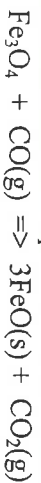
$$\Delta H = -63.5 \text{ kJ}$$



$$\Delta H = 1170.4 \text{ kJ}$$

$$\Delta H_{\text{rxn}} = 295.5 \text{ kJ}$$

2. Use the thermochemical equations shown below to determine the enthalpy for the reaction:



$$\Delta H = 51.7 \text{ kJ}$$



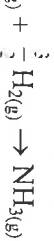
$$\Delta H = 87.7 \text{ kJ}$$



$$\Delta H = -2.3 \text{ kJ}$$

$$\Delta H_{\text{rxn}} = -40.6 \text{ kJ}$$

3. Given equations (1), (2), and (3), calculate the heat of reaction for equation.



$$\Delta H_f^\circ = -46.15 \text{ kJ}$$



$$\Delta H_f^\circ = 33.81 \text{ kJ}$$



$$\Delta H_f^\circ = -241.6 \text{ kJ}$$



$$\Delta H_{\text{rxn}} = -282.44 \text{ kJ}$$

4. Calculate  $\Delta H$  for the following reaction using the thermochemical equations given.



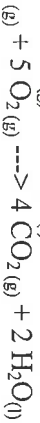
$$\Delta H = ?$$



$$\Delta H = -393.5 \text{ kJ}$$



$$\Delta H = -285.8 \text{ kJ}$$



$$\Delta H = -2598.8 \text{ kJ}$$

$$\Delta H = 226.6 \text{ kJ}$$

5. A small piece of aluminium of mass 5 kg is used on a circuit that has a normal temperature of 20 °C. If the piece of aluminium absorbs 585 000 J, what is its final temperature? Specific heat capacity of Aluminium is 0.900 J/g°C.

$$585000 = 5000(0.9) \Delta T$$

$$T_f = 150^\circ\text{C}$$

6. A metal sample with a mass of 60 g at 100°C is placed into 0.4 kg of water at 20°C. If the final temperature is 22.1°C, what is the specific heat of the metal sample? (Calculate the heat gained by the water and then assume that the heat gained by the water is equal to the heat lost by the metal.)  $c_w = 4.184 \text{ J/g}^\circ\text{C}$  [5]

$$q = m \cdot c \cdot \Delta T = 400(4.184)(22.1) = 3514$$

$$c_{\text{metal}} = 0.1519 \text{ J/g}^\circ\text{C}$$

7.  $\text{HNO}_3(\text{aq})$

$$\Delta H_f^\circ = -207 \text{ kJ/mol}$$

$$\text{H}_2\text{O}(\text{l}) \quad \Delta H_f^\circ = -285.8 \text{ kJ/mol}$$

$$\text{Mg}(\text{NO}_3)_2(\text{aq}) \quad \Delta H_f^\circ = -875 \text{ kJ/mol}$$

- (a) Given the  $\Delta H_f^\circ$  values above, calculate  $\Delta H_{\text{rxn}}^\circ$  for the equation below, using the direct method:



- (b) Is the reaction above exothermic or endothermic? How can you tell? *negative*

- (c) Write the thermochemical equation with the heat value included.

8. Consider the following thermochemical equation.



- Calculate  $\Delta H_f^\circ$  for  $\text{C}_6\text{H}_{12}\text{O}_6(\text{s})$ , given the following information:

$$\Delta H_f^\circ \text{ CO}_2(\text{g}) = -393.5 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{ H}_2\text{O}(\text{l}) = -285.8 \text{ kJ/mol}$$

$$\Delta H = -1273.8 \text{ kJ/mol}$$

9. 1.435 g of naphthalene ( $\text{C}_{10}\text{H}_8$ ) is burned in a bomb calorimeter. There is exactly 2000.0 g of water surrounding the naphthalene. The temperature of the water rises from 20.17 °C to 25.85 °C. Write the thermochemical equation including the heat value.



10. Calculate the standard enthalpy (heat) of reaction for the reaction of ammonia gas with oxygen to produce nitrogen dioxide gas and water vapour using the direct method and table of enthalpies of formation. ( $\text{NO}_2(\text{g}) \Delta H_f^\circ = +34 \text{ kJ mol}^{-1}$ )

$$-565.6 \text{ kJ} / 2 \text{ mol NH}_3$$

11. Use standard heats of formation (from the table) and the direct method to determine the heat of combustion for the following reaction. (4 marks)



$$\Delta H = -2818.4 \text{ kJ}$$

12. Calculate the heat of reaction,  $\Delta H$ , for the following reaction, using the direct method. Use the table of enthalpy values.



$$\Delta H = -890.7 \text{ kJ}$$

$$\begin{aligned}
 6. \quad Q &= m \cdot c \cdot \Delta t \\
 &= 400 \text{g} (4.184) (22.1) \\
 &= 3514.56 \text{ J}
 \end{aligned}$$

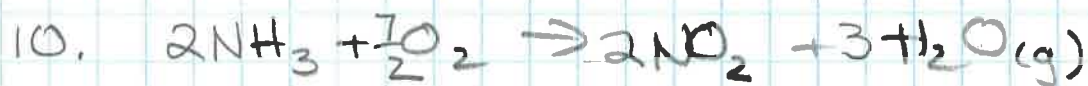
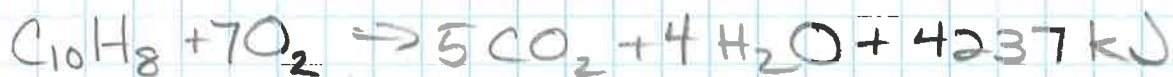
$$\frac{3514.56}{4674} = \frac{60 (c) (100 - 22.1)}{4674}$$

$$c = 0.7519$$

$$\begin{aligned}
 9. \quad Q &= m \cdot c \cdot \Delta t \\
 &= 2000 \times 4.184 (25.85 - 20.17) \\
 &= 47530.24 \text{ J} = 47.5 \text{ kJ}
 \end{aligned}$$

$$1.435 \text{ g } C_{10}H_8 \quad \left. \begin{array}{l} 120 + 8 = 128 \end{array} \right\} 0.01121 \text{ mol}$$

$$\therefore 4237 \text{ kJ}$$



$$\Delta H = [2(34) + 3(-241.8)] - [2(-45.9) + 0]$$

$$= -657.4 + 91.8$$

$$= -565.6 \text{ kJ} / 2 \text{ mol } NH_3$$

$$5. \quad 585000 = 5000 (0.9) \Delta t$$

$$\Delta t = 130$$

$$\therefore t_f = 150^\circ C$$



$$\begin{aligned} \Delta H &= [6(-285.8) + 4(-393.5)] \\ &\quad - [2(-235.2) + 0] \\ &= (-1714.8 - 1574) + 470.4 \\ &= -2818.4 \text{ kJ} \end{aligned}$$



$$\begin{aligned} \Delta H &= [-393.5 + 2(-285.8)] \\ &\quad - [(-74.4) + 0] \\ &= -890.7 \text{ kJ} \end{aligned}$$