

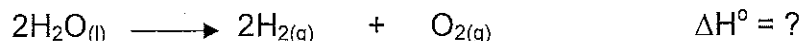
KEY

Unit A: Thermochemical Changes - Review

1.) The correct units for specific heat capacity are

- a. J/(g·°C)
- b. k J/(g·°C)
- c. kg/(J·°C)
- d. °C/(g·J)

2.) Water can be decomposed by electrical energy as shown in the following equation:



Solution at back.

The value for ΔH° for the equation as written is + 572 kJ (to three digits).

3.) Assuming coal is pure carbon, the temperature increase of 1.00 tonnes of water when it absorbs all the energy produced by the complete combustion of 1.00kg of coal is calculated to be 7.82 °C.

Solution at back.

Use the following statements to answer question 4 and 5.

- i. Chemical bonds are broken.
- ii. Chemical bonds are formed.
- iii. Energy is lost to the surroundings.
- iv. Energy is gained from the surroundings.
- v. The chemical potential energy of the products is less than the enthalpy of the reactants.
- vi. The chemical potential energy of the products is greater than the enthalpy of the reactants.

4.) Which statements are true for exothermic reactions?

I, II, III, and V

5.) Which statements are true for endothermic reactions?

I, II, IV, and VI

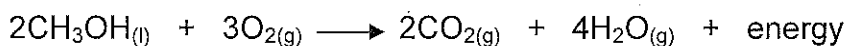
6.) The energy barrier that must be overcome in order for a reaction to occur is referred to as the

- a. activation energy
- b. bond energy
- c. enthalpy change
- d. molar enthalpy

- 7.) Which of the following statements about catalysts are **false**?
- Catalysts are not consumed in the net reaction.
 - (b)** Catalysts increase the enthalpy of the products.
 - Catalysts increase the overall rate of a reaction.
 - Catalysts provide an alternative pathway for the reaction.

- 8.) A chemistry student determines in an experiment that the molar enthalpy of combustion of methanol is -721 kJ/mol.

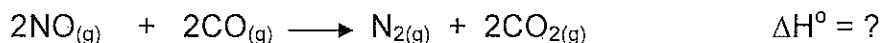
Solution at back



When the energy term is written as a term in the reaction equation under the same conditions, the value is 1442 kJ.

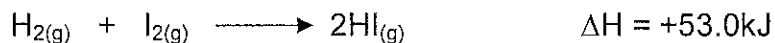
- 9.) An automobile catalytic converter is a technological invention for the conversion of noxious exhaust gases. For example:

Solution at back



The standard enthalpy change for this reaction is (to three digits) - 749 kJ.

- 10.) A much studied chemical reaction is communicated by the reaction equation



This equation is representative of an

- (a)** endothermic formation reaction
 - endothermic decomposition reaction
 - exothermic formation reaction
 - exothermic decomposition reaction
- 11.) What is the standard molar enthalpy of formation for all elements in their most stable state under standard conditions?

0.0 kJ/mol

- 12.) What is the difference between enthalpy change and molar enthalpy change?

→ Enthalpy change (ΔH) is measured in joules and is the potential energy of a chemical reaction

→ Molar enthalpy ($\Delta_r H$) is measured in joules/mol and is the potential energy of a reaction per mole of a specific chemical

13.) You are given a 70g sample of each of the metals shown in the table below, all at 25°C. You heat each metal under identical conditions. Which metal will be first to reach 30°C? Why?

METAL	SPECIFIC HEAT CAPACITY (J/g·°C)
platinum	0.133
titanium	0.523
zinc	0.388

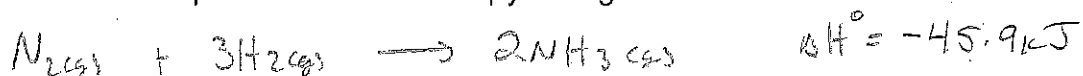
Platinum will reach 30°C first because it has the lowest specific heat capacity. This means that platinum requires the least amount of energy to change the temperature of a specific mass of that metal.

14.) When ammonia is formed from its elements in standard state, 45.9kJ of energy is released.

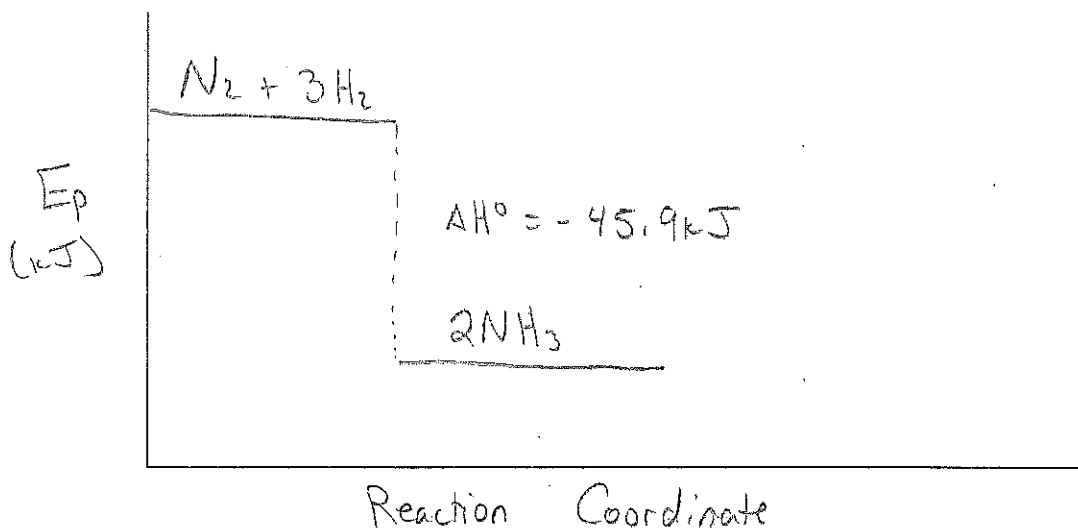
a. Write a thermochemical equation representing this reaction.



b. Write this equation with an enthalpy change.



c. Draw a potential energy diagram for this reaction.



15.) Write balanced thermochemical equations for the formation of the following substances at standard conditions:

a. $\text{CH}_3\text{COOH}(\text{l})$

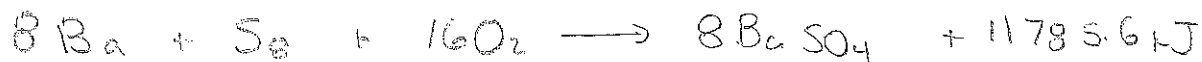
$$\Delta_f H^\circ = -484.3 \text{ kJ}$$

$$\therefore \Delta H^\circ = n \Delta_f H^\circ = (1 \text{ mol})(-484.3 \text{ kJ})$$



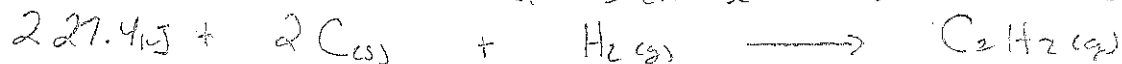
b. $\text{BaSO}_4(\text{s})$ $\Delta_f H^\circ = -1473.2 \text{ kJ}$

$$\therefore \Delta H^\circ = n \Delta_f H^\circ = (8 \text{ mol})(1473.2 \text{ kJ}) = 11785.6 \text{ kJ}$$



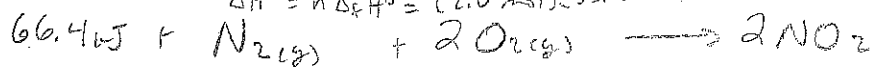
c. $\text{C}_2\text{H}_2(\text{g})$ $\Delta_f H^\circ = +227.4 \text{ kJ}$

$$\Delta H^\circ = n \Delta_f H^\circ = (1 \text{ mol})(227.4 \text{ kJ}) = 227.4 \text{ kJ}$$

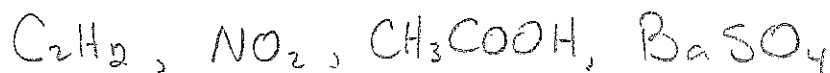


d. $\text{NO}_2(\text{g})$ $\Delta_f H^\circ = +33.2 \text{ kJ}$

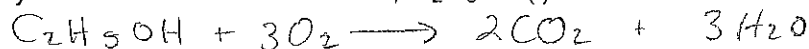
$$\Delta H^\circ = n \Delta_f H^\circ = (2.0 \text{ mol})(33.2 \text{ kJ}) = 66.4 \text{ kJ}$$



16.) Rank the substance in question 15 from least thermally stable to most thermally stable.



17.) Use the following data to determine an experimental value for the molar enthalpy of combustion for ethanol, $C_2H_5OH(l)$:



initial mass of alcohol burner = 47.39g

final mass of alcohol burner = 47.19g

mass of aluminum can = 84.37g

mass of aluminum can and water = 239.36g

initial temperature = 15.7°C

final temperature of water = 18.3°C

difference is 0.20g

difference is 154.99g

$$\Delta H = n \Delta_c H \quad (3)$$

↓

$$Q = -\Delta H \quad (2)$$

↓

$$Q = mc\Delta T \quad (1)$$

$$(1) \quad Q = mc\Delta T$$

$$Q = (154.99g)(4.19J/g \cdot ^\circ C)(18.3 - 15.7^\circ C)$$

$$Q = +1688.46J$$

$$(2) \quad \Delta H = -Q$$

$$\Delta H = -1688.46J$$

$$(3) \quad \Delta_c H = \frac{\Delta H}{n}$$

$$\Delta_c H = \frac{-1688.46J}{0.00434mol}$$

$$n = \frac{m}{M} = \frac{0.20g}{46.08g/mol} = 0.00434mol$$

$$\Delta_c H = 3.89 \times 10^5 J/mol \quad \text{or} \quad 389J$$

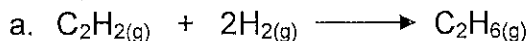
b.) The accepted value for the molar enthalpy of combustion is -507kJ/mol. Calculate the percent error.

$$\% \text{ error} = \frac{H_{\text{reor.}} - \text{exp.}}{H_{\text{reor.}}} \times 100\%$$

$$\% \text{ error} = \frac{(-507) - (-389J)}{-507} \times 100\%$$

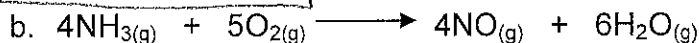
$$\% \text{ error} = 23.3\%$$

18.) Using standard enthalpies of formation, determine the enthalpy of reaction for the following:



$$\Delta H^\circ = (1mol)(-84.0kJ/mol) - (1mol)(+227.4kJ/mol)$$

$$\Delta H^\circ = -311.4kJ$$

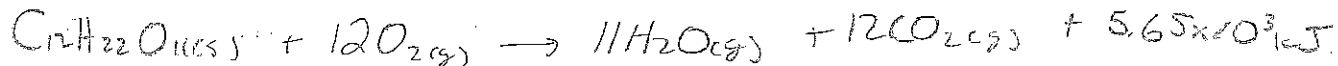


$$\Delta H = [(4mol)(91.3kJ/mol) + (6mol)(-241.8kJ/mol)] - [(4mol)(-45.9kJ/mol)]$$

$$\Delta H = -902.0kJ$$

19.) The complete combustion of 1.00 mol of sucrose, $C_{12}H_{22}O_{11}(s)$, yields $5.65 \times 10^3 \text{ kJ}$ of energy.

a. Write a balanced thermochemical equation for the combustion of sucrose in an open system.



b. Calculate the amount of energy that is released when 5.00 g of sucrose is burned.

$$\Delta H = n \Delta_r H$$

for specific amount (5.00g)

$$\Delta_r H = \frac{\Delta H}{n}$$

from balanced rxn (1.00 mol)

$$\textcircled{1} \Delta_r H = \frac{-5.65 \times 10^3 \text{ kJ}}{1.00 \text{ mol}}$$

$$n = \frac{m}{M} = \frac{5.00 \text{ g}}{342.34 \text{ g/mol}}$$

$$n = 0.145 \text{ mol}$$

$$\Delta_r H = -5.65 \times 10^3 \text{ kJ/mol}$$

$$\textcircled{2} \Delta H = n \Delta_r H$$

$$\Delta H = (0.145 \text{ mol})(-5.65 \times 10^3 \text{ kJ/mol})$$

$$\Delta H = -825 \text{ kJ}$$

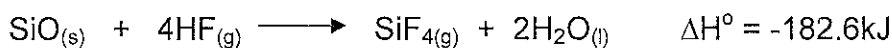
c. How much water (in grams) could be warmed by 5.0°C when 5.00 g of sucrose is completely burned?

$$Q = -\Delta H$$

$$Q = m c \Delta t \Rightarrow m = \frac{Q}{c \Delta t} = \frac{82500 \text{ J}}{(4.19 \text{ J/g}^\circ\text{C})(5.0^\circ\text{C})}$$

$$m = 3.9 \times 10^3 \text{ g}$$

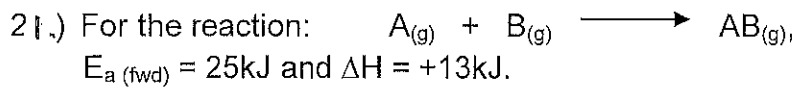
20.) Given the following thermochemical equation, determine the molar enthalpy of formation for hydrogen fluoride, $HF_{(g)}$.



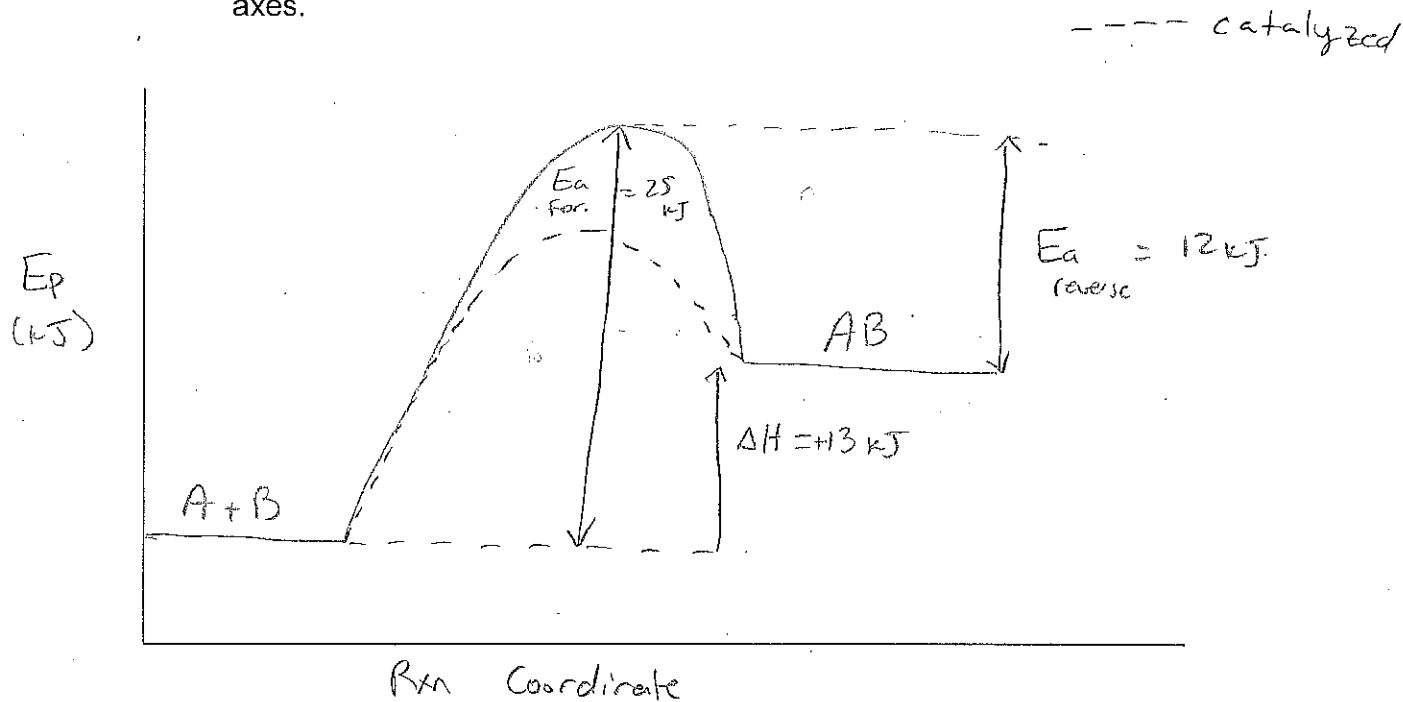
$$\Delta H = n \Delta_r H$$

$$\Delta_r H = \frac{\Delta H}{n} = \frac{-182.6 \text{ kJ}}{4 \text{ mol}}$$

$$\Delta_r H = 44.7 \text{ kJ/mol}$$



- a. Sketch a potential energy diagram for this reaction. Be sure to label the axes.



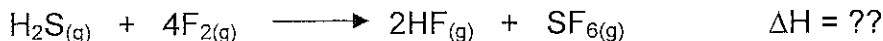
- b. Determine the activation energy for the reverse reaction and add it to the diagram.

$$E_{a \text{ rev.}} = 25 \text{ kJ} - 13 \text{ kJ} = \boxed{12 \text{ kJ}}$$

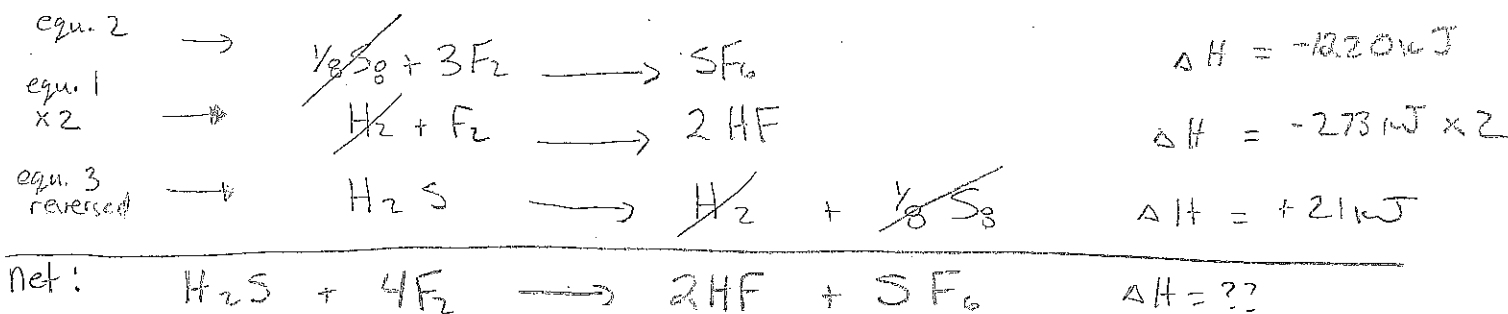
- c. Sketch the effect of a catalyst on the energy pathway.

---- catalyzed

22.) Consider the following reaction:



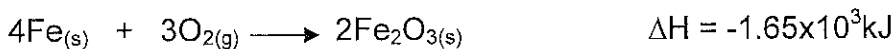
Determine the enthalpy change for the above reaction using the following information.



$$\Delta H = (-1220 \text{ kJ}) + (-273 \text{ kJ})(2) + (21 \text{ kJ})$$

$$\Delta H = 1745 \text{ kJ}$$

23.) Rust, $\text{Fe}_2\text{O}_3(s)$, forms when iron in steel is exposed to water. A simplified reaction is shown below:



a.) How much energy is released when 1.37kg of iron rusts?

$$\textcircled{2} \Delta H = n \Delta_r H \quad \text{for specific amount}$$

$$\textcircled{1} \Delta_r H = \frac{-1.65 \times 10^3 \text{ kJ}}{4 \text{ mol}} = -412.5 \text{ kJ/mol}$$

$$\textcircled{1} \Delta_r H = \frac{\Delta H}{n} \quad \text{from balanced rxn}$$

$$\textcircled{2} \Delta H = n \Delta_r H = (24.52999 \text{ mol})(-412.5 \text{ kJ/mol})$$

$$\Delta H = 1.01 \times 10^4 \text{ kJ}$$

$$n = \frac{m}{M} = \frac{1370 \text{ g}}{55.85 \text{ g/mol}} = 24.52999 \text{ mol}$$

b.) How much rust forms along with the release of 3.75MJ of energy?

$$\textcircled{2} \Delta H = n \Delta_r H \quad \text{for specific amount}$$

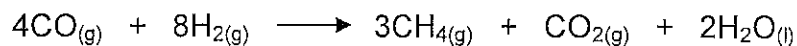
$$\textcircled{1} \Delta_r H = \frac{-1.65 \times 10^3 \text{ kJ}}{2 \text{ mol}} = 825 \text{ kJ/mol}$$

$$\textcircled{1} \Delta_r H = \frac{\Delta H}{n} \quad \text{from balanced rxn}$$

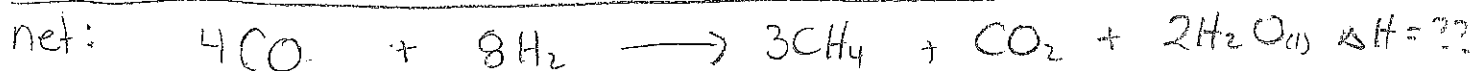
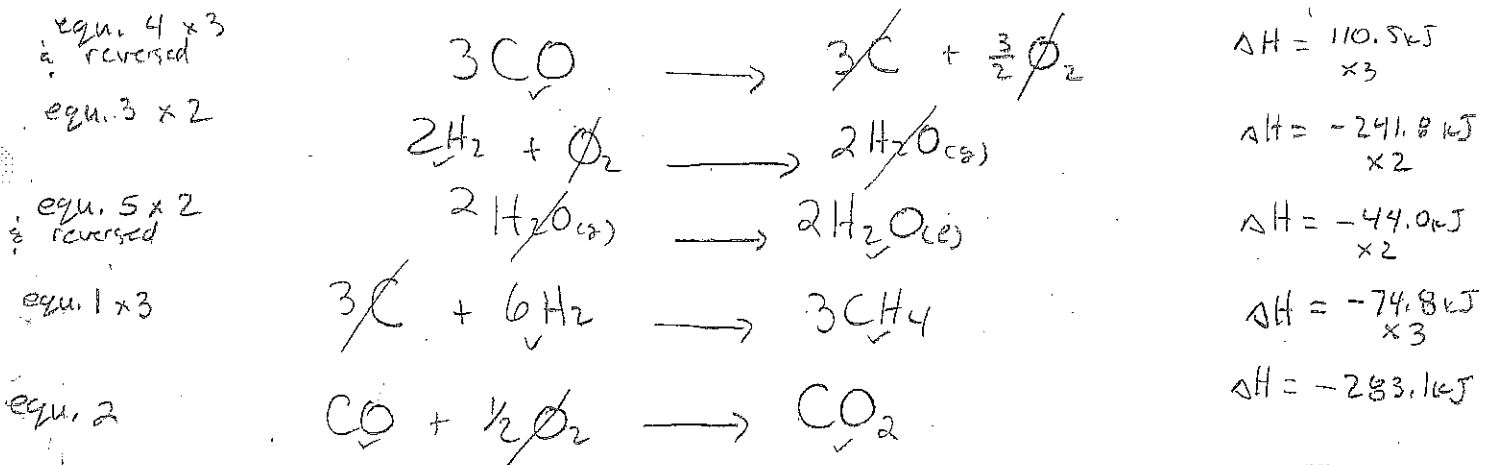
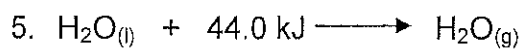
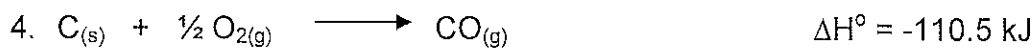
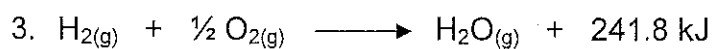
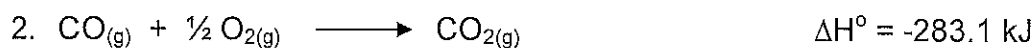
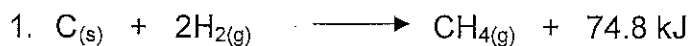
$$\textcircled{2} n = \frac{\Delta H}{\Delta_r H} = \frac{3.75 \times 10^3 \text{ kJ}}{825 \text{ kJ/mol}} = 4.54 \text{ mol}$$

$$\textcircled{3} m = M_n = (159.70 \text{ g/mol})(4.54 \text{ mol}) \Rightarrow m = 726 \text{ g}$$

24.) Carbon monoxide reacts with hydrogen gas to produce a mixture of methane, carbon dioxide, and water. The reaction is summarized below:



Calculate the enthalpy change of the reaction.



$$\Delta H = (110.5 \text{ kJ})(3) + (-241.8 \text{ kJ})(2) + (-44.0 \text{ kJ}) + (-74.8 \text{ kJ})(3) + (-283.1 \text{ kJ})$$

$\Delta H = -747.6 \text{ kJ}$

$$2.) \Delta H^\circ = \sum n \Delta_f H^\circ_{\text{product}} - \sum n \Delta_f H^\circ_{\text{reactants}}$$

$$\Delta H^\circ = 0 - [(2 \text{ mol})(-285.8 \text{ kJ/mol})]$$

$$\Delta H^\circ = 571.6 \text{ kJ} \Rightarrow \boxed{\Delta H^\circ = 572 \text{ kJ}}$$



$\Delta H =$ use molar enthalpies of formation

$$\Delta H^\circ = \sum n \Delta_f H^\circ_{\text{products}} - \sum n \Delta_f H^\circ_{\text{reactants}}$$

$$\Delta H^\circ = (1 \text{ mol})(-393.5 \text{ kJ/mol}) - 0$$

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

Now ④ $Q = mc\Delta t$

↓

③ $Q = -\Delta H$

↓

② $\Delta H = n \Delta_r H$ for specific amount

↓

① $\Delta_r H = \frac{\Delta H}{n}$ from balanced rxn

① $\Delta_r H = \frac{-393.5 \text{ kJ}}{1 \text{ mol}} = -393.5 \text{ kJ/mol}$

② $\Delta H = n \Delta_r H = \left(\frac{m}{M}\right) \Delta_r H = \left(\frac{1000 \text{ g}}{12.01 \text{ g/mol}}\right) (-393.5 \text{ kJ/mol}) = -32764 \text{ kJ}$
↑
of carbon

③ $Q = -(-32764 \text{ kJ}) = +32764 \text{ kJ}$

④ $\Delta t = \frac{Q}{mc} = \frac{32764000 \text{ J}}{(1.00 \times 10^6 \text{ g})(4.19 \text{ J/g}^\circ\text{C})} \Rightarrow \boxed{\Delta t = 7.82^\circ\text{C}}$

$$8.) \Delta H = n \Delta_r H$$

$$\Delta H = (2 \text{ mol}) (-721 \text{ kJ/mol})$$

$$\Delta H = 1442 \text{ kJ}$$

9.) Use molar enthalpies of formation

$$\Delta H^\circ = \sum_{\text{prod.}} n \Delta_f H^\circ - \sum_{\text{react.}} n \Delta_f H^\circ$$

$$\Delta H^\circ = \left[(2 \text{ mol}) (-393.5 \text{ kJ/mol}) \right]_{\text{CO}_2} - \left[(2 \text{ mol}) (-110.5 \text{ kJ}) + (2 \text{ mol}) (+91.3 \text{ kJ/mol}) \right]_{\text{CO NO}}$$

$$\Delta H^\circ = -787 \text{ kJ} - (-38.4 \text{ kJ})$$

$$\Delta H^\circ = -748.6 \text{ kJ}$$

$$\therefore \Delta H^\circ = -749 \text{ kJ}$$

