REVIEW QUESTIONS
Chapter 6

1. What is the change in internal energy (in J) of a system that releases 575 J of heat to its surroundings and has 425 cal of work done on it?

$$\Delta E = q + w$$

$$q = \text{-575 J} \qquad\qquad w = \text{+ 425 cal x } \frac{\textbf{4.18 J}}{\textbf{1 cal}} = \textbf{+17}\underline{\textbf{7}}\textbf{7 J}$$

$$\Delta E = \text{-575 J} + \textbf{17}\underline{\textbf{7}}\textbf{7 J} = \textbf{+1200 J}$$

2. If an athlete expends 1700 kJ/h playing tennis, how long would she have to play to work off 1.0 lb of body fat? (One pound of body fat is equivalent to 4100 nutritional Calories)

$$\textbf{1.0 } \cancel{\textbf{lb fat}} \textbf{ x } \frac{\textbf{4100 }\cancel{\textbf{Cal}}}{\textbf{1 }\cancel{\textbf{lb fat}}} \textbf{ x } \frac{\textbf{1 }\cancel{\textbf{kcal}}}{\textbf{1 }\cancel{\textbf{Cal}}} \textbf{ x } \frac{\textbf{4.18 }\cancel{\textbf{kJ}}}{\textbf{1 }\cancel{\textbf{kcal}}} \textbf{ x } \frac{\textbf{1 hr}}{\textbf{1700 }\cancel{\textbf{kJ}}} = \textbf{10. hrs}$$

3. A 27.7 g sample of ethylene glycol, a car radiator coolant, absorbs 588 J of heat from the car engine. What is the initial temperature of the engine, if its final temperature is 30.5°C? (C_s of ethylene glycol = 2.42 J/g°C)

$$\Delta T = \frac{\textbf{q}}{\textbf{m x } C_s} = \frac{\textbf{588 J}}{\textbf{27.7 g x 2.42 J/g°C}} = \textbf{8.77 °C}$$

$$T_i = T_f + \Delta T = \textbf{30.5} + \textbf{8.77} = \textbf{39.3 °C}$$

4. When 155 mL of water at 20.0°C is mixed with 75.0 mL of water at 80.0°C, what is the final temperature? (Assume specific heat of hot and cold water are the same, and density of water = 1.00 g/mL)

$$q_h = \text{-}q_c$$

$$(\textbf{m x } \cancel{C_s} \textbf{ x } \Delta T)_h = \textbf{- (m x } \cancel{C_s} \textbf{ x } \Delta T)_c \qquad \textbf{(since } C_s \textbf{ of hot and cold water are same)}$$

$$\frac{\Delta T_h}{\Delta T_c} = \textbf{-}\frac{m_c}{m_h} = \textbf{-}\frac{\textbf{155 g}}{\textbf{75 g}} = \textbf{-2.07}$$

$$\Delta T_h = \textbf{-2.07 } \Delta T_c$$

$$(T_f \textbf{ - 80.0}) = \textbf{-2.07}(T_f \textbf{ -20.0})$$

$$\textbf{3.07 } T_f = \textbf{121.4} \qquad\qquad T_f = \textbf{39.5 °C}$$

5. A 1.500-g sample of a hydrocarbon is placed in a bomb calorimeter and undergoes combustion. The temperature of the calorimeter rises from 20.00°C to 23.55°C. If the calorimeter has a heat capacity of 40.3 J/°C, what is the heat of combustion (q_v) for this hydrocarbon in kJ/g?

$$q_v = C_{cal} \times \Delta T = (40.3 \text{ J/}^o\text{C})(3.55 \text{ }^o\text{C}) = -143 \text{ J} \quad \text{(- sign: exothermic)}$$

$$q_v = -\frac{143 \text{ Jx}\dfrac{1 \text{ kJ}}{10^3 \text{ J}}}{1.500 \text{ g}} = -0.0953 \text{ kJ/g}$$

6. The heat of sublimation for iodine is 62.4 kJ/mol, at 25°C and 1.00 atm. What is ΔE for the sublimation of iodine (shown below) under these conditions?

$$I_2 \text{ (s)} \rightleftharpoons I_2 \text{ (g)}$$
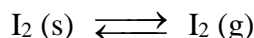
$\Delta E = q + w$

$q = +62.4 \text{ kJ}$ **(sublimation is endothermic)**
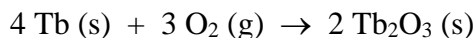
$w = -P\Delta V$ **(work is done by system during expansion)**

ΔV = volume occupied by 1 mol iodine gas at 25°C

$$\Delta V = \frac{nRT}{P} = \frac{(1.00 \text{ mol})(0.0821)(298 \text{ K})}{1.00 \text{ atm}} = 24.5 \text{ L}$$

$w = -(1.00 \text{ atm})(24.5 \text{ L})(0.1013 \text{ kJ/Latm}) = -2.48 \text{ kJ}$

$\Delta E = +62.4 \text{ kJ} -2.48 \text{ kJ} = +59.9 \text{ kJ}$

7. A bomb calorimeter with heat capacity of 1.500×10^3 J/°C was used to study combustion of terbium (Tb) in excess oxygen:

$$4 \text{ Tb (s)} + 3 \text{ O}_2 \text{ (g)} \rightarrow 2 \text{ Tb}_2\text{O}_3 \text{ (s)}$$

When 1.000-g of terbium was burned, a temperature increase of 3.908°C was observed. What is the ΔE for the reaction, in kJ/mol of Tb_2O_3?

$$q_v = C_{cal} \times \Delta T = (1.500 \times 10^3 \text{ J/}^o\text{C})(3.908 \text{ }^o\text{C}) = -5862 \text{ J} \quad \text{(- sign: exothermic reaction)}$$

$$1.000 \text{ g Tb x } \frac{1 \text{ mol}}{158.9 \text{ g}} \text{ x } \frac{2 \text{ mol Tb}_2\text{O}_3}{4 \text{ mol Tb}} = 0.003147 \text{ mol Tb}_2\text{O}_3$$

$$\Delta E = q_v = \frac{-5862 \text{ Jx}\dfrac{1 \text{ kJ}}{10^3 \text{ J}}}{0.003147 \text{ mol}} = -1863 \text{ kJ/mol}$$

8. One mole of nitrogen gas confined in a cylinder with a piston is heated from 0°C to 819°C at 1.00 atm.
   a) Calculate the work of expansion in Joules.

   w = -PΔV   (work is done by system during expansion)

   ΔV = volume change for 1 mol nitrogen gas from 0 °C to 819°C

   $$V_{819} = 22.4 \text{ L x } \frac{1092 \text{ K}}{273 \text{ K}} = 89.6 \text{ L} \qquad ΔV = (89.6 \text{ L} - 22.4 \text{ L}) = 67.2 \text{ L}$$

   w = -(1.00 atm)(67.2 L)(101.3 J/Latm)= -6.81x10³ J

   b) What would have been the temperature change if the gas had been heated with the same amount of energy in a container of fixed volume. (Assume the specific heat of N_2 is 1.00 J/g°C)

   In a bomb calorimeter (fixed volume) the work of expansion calculated above would be converted to heat. Therefore,
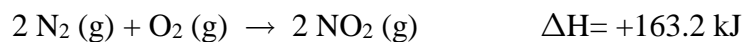
   $$ΔT = \frac{q}{m \text{ x } C_s} = \frac{6.81 \text{x} 10^3 \text{ J}}{28.02 \text{ g x } 1.00 \text{ J/g}°\text{C}} = 243 \text{ }°\text{C}$$

9. How many kJ of heat is produced when 5.00 g of Al reacts with excess chlorine, as shown below:

   $$2 \text{ Al (s)} + 3 \text{ Cl}_2 \text{ (g)} \rightarrow 2 \text{ AlCl}_3 \text{ (s)} \qquad ΔH = -1408 \text{ kJ}$$
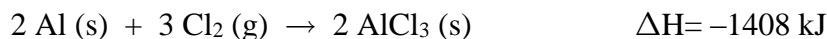
   $$5.00 \text{ g Al x } \frac{1 \text{ mol}}{26.98 \text{ g}} \text{ x} \frac{1408 \text{ kJ}}{2 \text{ mol Al}} = 1.30 \text{x} 10^2 \text{ kJ}$$

10. Consider the following reaction:

   $$2 \text{ N}_2 \text{ (g)} + \text{O}_2 \text{ (g)} \rightarrow 2 \text{ NO}_2 \text{ (g)} \qquad ΔH = +163.2 \text{ kJ}$$

   a) How many grams of N_2 must react to produce an enthalpy change of 5.00 kJ?

   $$5.00 \text{ kJ x} \frac{2 \text{ mol N}_2}{163.2 \text{ kJ}} \text{ x} \frac{28.02 \text{ g}}{1 \text{ mol}} = 1.72 \text{ g N}_2$$

   b) How much heat is transferred when 25.0 g of NO_2 is produced in this reaction?

   $$25.0 \text{ g NO}_2 \text{ x } \frac{1 \text{ mol}}{46.01 \text{ g}} \text{ x} \frac{163.2 \text{ kJ}}{2 \text{ mol NO}_2} = 44.3 \text{ kJ}$$

3

11. When 6.50 g of solid NaOH is dissolved in 100.0 g of water in a calorimeter, the temperature rises from 21.6 °C to 37.8 °C. Calculate the change in enthalpy of solution in kJ/mol of NaOH. Assume specific heat of solution is the same as water.

mass of solution = 100.0 g + 6.50 g = 106.5 g

$q = (m \times C_s \times \Delta T) = (106.5 \text{ g})(4.18 \text{ J/g °C})(16.2 \text{ °C}) = -72\underline{1}2 \text{ J}$ (- sign: exothermic)

$6.50 \text{ g NaOH} \times \dfrac{1 \text{ mol}}{40.00 \text{ g}} = 0.16\underline{2}5 \text{ mol}$

$\Delta H = \dfrac{-72\underline{1}2 \text{ J} \times \dfrac{1 \text{ kJ}}{10^3 \text{ J}}}{0.16\underline{2}5 \text{ mol}} = -44.4 \text{ kJ/mol NaOH}$

12. When 1.025 g of naphthalene ($C_{10}H_8$) is burned in a bomb calorimeter, the temperature rises from 24.25°C to 32.33°C. Find ΔE_{rxn} for combustion of naphthalene in kJ/mol. Heat capacity of the calorimeter is 5.11 kJ/°C.

$q_v = C_{cal} \times \Delta T = (5.11 \text{ kJ/°C}) (8.08 \text{ °C}) = -41.3 \text{ kJ}$ (- sign: exothermic reaction)

$1.025 \text{ g } C_{10}H_8 \times \dfrac{1 \text{ mol}}{128.18 \text{ g}} = 8.00 \times 10^{-3} \text{ mol } C_{10}H_8$

$\Delta E = q_v = \dfrac{-41.3 \text{ kJ}}{8.00 \times 10^{-3} \text{ mol}} = -5.16 \times 10^3 \text{ kJ/mol}$

13. When 10.0 g of phosphorus is burned in oxygen gas to form solid P_4O_{10}, enough heat is generated to raise the temperature of 2960 g of water from 18.0°C to 38.0°C. Calculate the enthalpy of formation of solid P_4O_{10} under these conditions.

$$4 P + 5 O_2 \longrightarrow P_4O_{10}$$

$q = (m \times C_s \times \Delta T) = (2960 \text{ g})(4.18 \text{ J/g °C})(20.0 \text{ °C}) = -2.4\underline{7}5 \times 10^5 \text{ J}$ (- sign: exothermic)

$10.0 \text{ g P} \times \dfrac{1 \text{ mol}}{30.97 \text{ g}} \times \dfrac{1 \text{ mol } P_4O_{10}}{4 \text{ mol P}} = 0.0807 \text{ mol } P_4O_{10}$

$\Delta H_f = \dfrac{-2.475 \times 10^5 \text{ J} \times \dfrac{1 \text{ kJ}}{10^3 \text{ J}}}{0.0807 \text{ mol}} = -3070 \text{ kJ/mol } P_4O_{10}$

4

14. Calculate the enthalpy change for the reaction below:
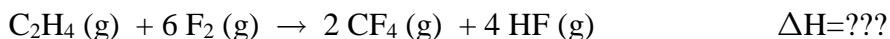
$$C_2H_4 \text{ (g)} + 6 F_2 \text{ (g)} \rightarrow 2 CF_4 \text{ (g)} + 4 HF \text{ (g)} \qquad \Delta H = ???$$

based on the following reactions:

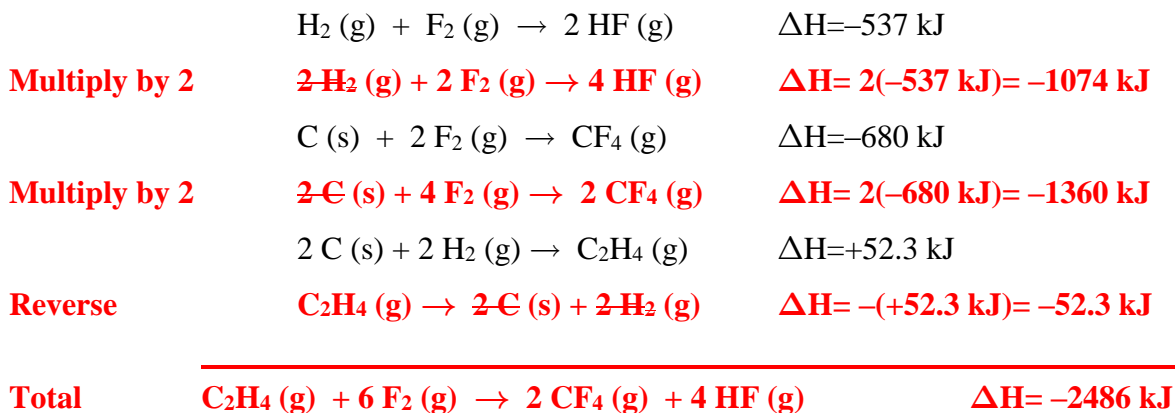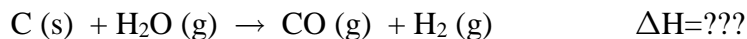|		
	H_2 (g) + F_2 (g) → 2 HF (g)	$\Delta H = -537$ kJ
Multiply by 2	**2 H₂ (g) + 2 F₂ (g) → 4 HF (g)**	**$\Delta H = 2(-537$ kJ$) = -1074$ kJ**
	C (s) + 2 F_2 (g) → CF_4 (g)	$\Delta H = -680$ kJ
Multiply by 2	**2 C (s) + 4 F₂ (g) → 2 CF₄ (g)**	**$\Delta H = 2(-680$ kJ$) = -1360$ kJ**
	2 C (s) + 2 H_2 (g) → C_2H_4 (g)	$\Delta H = +52.3$ kJ
Reverse	**C₂H₄ (g) → 2 C (s) + 2 H₂ (g)**	**$\Delta H = -(+52.3$ kJ$) = -52.3$ kJ**

Total	**C₂H₄ (g) + 6 F₂ (g) → 2 CF₄ (g) + 4 HF (g)**	**$\Delta H = -2486$ kJ**

15. Determine the enthalpy change for the reaction below:

$$C \text{ (s)} + H_2O \text{ (g)} \rightarrow CO \text{ (g)} + H_2 \text{ (g)} \qquad \Delta H = ???$$

based on the following reactions:

	C (s) + O_2 (g) → CO_2 (g)	$\Delta H = -393.5$ kJ
No change	**C (s) + O₂ (g) → CO₂ (g)**	**$\Delta H = -393.5$ kJ**
	2 CO (g) + O_2 (g) → 2 CO_2 (g)	$\Delta H = -566.0$ kJ
Reverse & divide by 2	**CO₂ (g) → CO (g) + ½ O₂ (g)**	**$\Delta H = -½ (-566.0$ kJ$) = +283.0$ kJ**
	2 H_2 (g) + O_2 (g) → 2 H_2O (g)	$\Delta H = -483.6$ kJ
Reverse & divide by 2	**H₂O (g) → H₂ (g) + ½ O₂ (g)**	**$\Delta H = -½(-483.6$ kJ$) = +241.8$ kJ**

Total	**C (s) + H₂O (g) → CO (g) + H₂ (g)**	**$\Delta H = +131.3$ kJ**
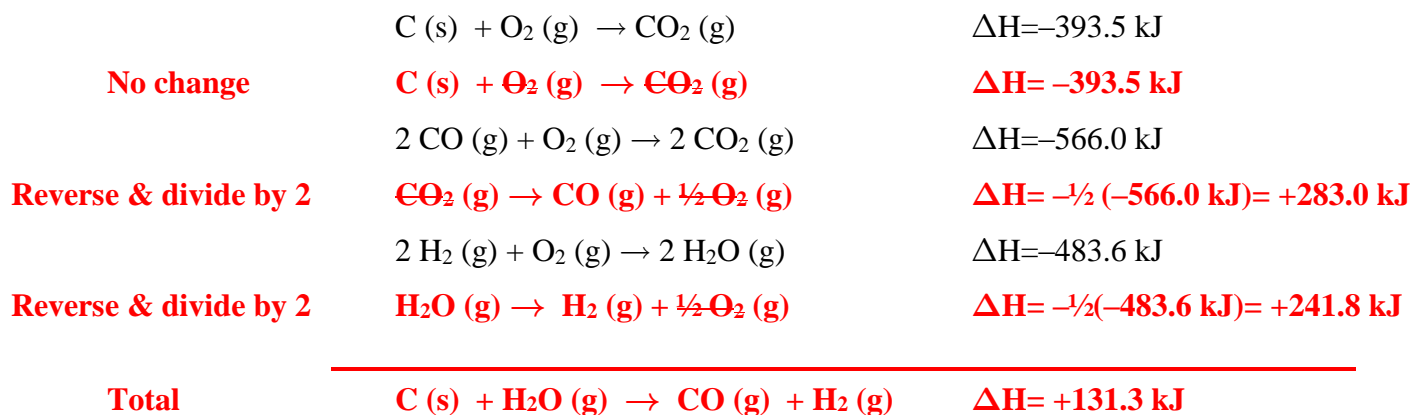
16. Use enthalpies of formation given to determine the standard enthalpy of reaction for the following:

$$2 \text{ Al (s)} \quad + \text{ Fe}_2\text{O}_3 \text{ (s)} \quad \rightarrow \quad \text{Al}_2\text{O}_3 \text{ (s)} \quad + \quad 2 \text{ Fe} \qquad \Delta H° = \text{???}$$
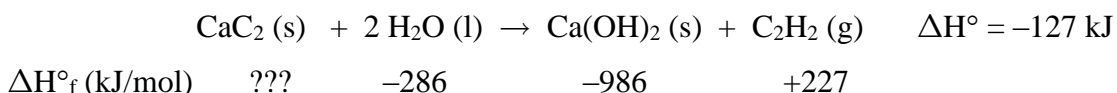
$\Delta H°_f$ (kJ/mol)   0   −825.5   −1676   0
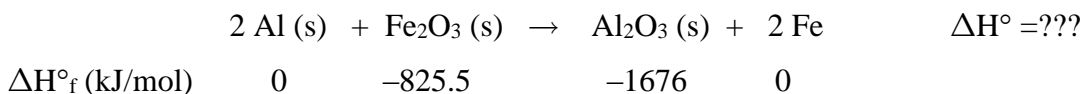
$\Delta H° = \Sigma \Delta H°$ (products) $- \Sigma \Delta H°$ (reactants)

$\Delta H° = [(-1676 \text{ kJ}) + 2(0)] - [2(0) + (-825.5 \text{ kJ})] = -851 \text{ kJ}$

17. Use the enthalpies of formation and the enthalpy of reaction given below to determine the enthalpy of formation for solid CaC_2.

$$\text{CaC}_2 \text{ (s)} \quad + \text{ 2 H}_2\text{O (l)} \quad \rightarrow \quad \text{Ca(OH)}_2 \text{ (s)} \quad + \quad \text{C}_2\text{H}_2 \text{ (g)} \qquad \Delta H° = -127 \text{ kJ}$$

$\Delta H°_f$ (kJ/mol)   ???   −286   −986   +227

$\Delta H° = \Sigma \Delta H°$ (products) $- \Sigma \Delta H°$ (reactants)

$\Delta H° = [(-986 \text{ kJ}) + (+227 \text{ kJ})] - [(\Delta H°_f) + 2(-286 \text{ kJ})] = -127 \text{ kJ}$

$\Delta H° = [-759 \text{ kJ}] - [(\Delta H°_f) + (-572)] = -127 \text{ kJ}$

$\Delta H°_f \text{ (CaC}_2) = -60 \text{ kJ/mol}$

18. The heat of vaporization of a liquid (ΔH_{vap}) is the heat required to vaporize 1.00 g of a liquid. In one experiment, 60.0 g of liquid nitrogen (b.p. −196 °C) are poured into a Styrofoam cup containing 200. g of water at 55.3 °C. Calculate the molar heat of vaporization of liquid nitrogen if the final temperature of the water is 41.0 °C.

Heat lost by water = Heat gained by liquid nitrogen

$q_{H_2O} = m \text{ x } C_s \text{ x } \Delta T = (200. \text{ g})(4.18 \text{ J/g °C})(14.3 °C) = +11\underline{9}55 \text{ J}$ (+ sign: endothermic)

$60.0 \text{ g N}_2 \text{ x } \dfrac{1 \text{ mol}}{28.02 \text{ g}} = 2.14 \text{ mol}$

$\Delta H_{vap} = \dfrac{+11\underline{9}55 \cancel{\text{ J}} \text{ x } \dfrac{1 \text{ kJ}}{10^3 \cancel{\text{ J}}}}{2.14 \text{ mol}} = 5.59 \text{ kJ/mol}$

6

19. Ice at 0°C is placed in a Styrofoam cup containing 361 g of a soft drink at 23.0°C. Some ice remains after the ice and the soft drink reach an equilibrium temperature of 0 °C. Determine the mass of the ice that has melted. Assume the specific heat of the drink to be the same as pure water. (It requires 334 J of heat to melt 1 g of ice at 0°C)

$$\text{Heat lost by drink = Heat gained by ice}$$

$$q_{drink} = m \times C_s \times \Delta T = (361 \text{ g})(4.18 \text{ J/g }^{\circ}\text{C})(23.0^{\circ}\text{C}) = 34\underline{7}00 \text{ J}$$

$$\text{mass of ice melted} = 34\underline{7}00 \text{ J} \times \frac{1 \text{ g ice}}{334 \text{ J}} = 104 \text{ g}$$

20. A quantity of 200. mL of 0.862 M HCl is mixed with 200. mL of 0.431 M Ba(OH)$_2$ in a Styrofoam cup. The initial temperature of both solutions is 20.48 °C. Calculate the heat of neutralization in kJ/mol of H$_2$O, if the final temperature of the solution is 26.30°C. (Assume specific heat and density of solution to be the same as water)

$$2 \text{ HCl} + \text{Ba(OH)}_2 \longrightarrow \text{BaCl}_2 + 2 \text{ H}_2\text{O}$$
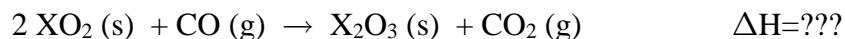
$$\text{mass of solution} = 400. \text{ mL} \times \frac{1.00 \text{ g}}{1 \text{ mL}} = 400. \text{ g}$$

$$q = (m \times C_s \times \Delta T) = (400. \text{ g})(4.18 \text{ J/g }^{\circ}\text{C})(5.82 ^{\circ}\text{C}) = -97\underline{3}0 \text{ J} \quad (\text{- sign: exothermic})$$

$$\text{mol H}_2\text{O} = 200. \text{ mL HCl} \times \frac{0.862 \text{ mol}}{1000 \text{ mL}} \times \frac{2 \text{ mol H}_2\text{O}}{2 \text{ mol HCl}} = 0.17\underline{2}4 \text{ mol}$$

$$\Delta H_{neut} = \frac{-97\underline{3}0 \text{ J} \times \frac{1 \text{ kJ}}{10^3 \text{ J}}}{0.17\underline{2}4 \text{ mol H}_2\text{O}} = \text{-56.4 kJ/mol}$$

21. Determine the enthalpy change for the reaction below:

$$2 XO_2 (s) + CO (g) \rightarrow X_2O_3 (s) + CO_2 (g) \qquad \Delta H=???$$

based on the following reactions:

$XO_2 (s) + CO (g) \rightarrow XO (s) + CO_2 (g)$ $\qquad \Delta H= -26.8$ kJ

Multiply by 2 $\qquad$ **2 XO₂ (s) + 2̶ CO (g) → 2̶ X̶O̶ (s) + 2̶ CO₂ (g)** $\qquad$ **ΔH= 2(–26.8) = –53.6 kJ**

$X_3O_4 (s) + CO (g) \rightarrow 3 XO (s) + CO_2 (g)$ $\qquad \Delta H= + 7.3$ kJ

Rev & multiply by 2/3 $\qquad$ **2̶ X̶O̶ (s) + 2/3 C̶O̶₂̶ (g) → 2/3 X₃O₄ (s) + 2/3 C̶O̶ (g)** $\qquad$ **ΔH= –2/3(7.3) = –4.9 kJ**

$3 X_2O_3 (s) + CO (g) \rightarrow 2 X_3O_4 (s) + CO_2 (g)$ $\qquad \Delta H= -10.6$ kJ

Rev & multiply by 1/3 $\qquad$ **2/3 X̶₃̶O̶₄̶ (s) + 1̶/̶3̶ ̶C̶O̶₂̶ (g) → X₂O₃ (s) + 1̶/̶3̶ ̶C̶O̶ (g)** $\qquad$ **ΔH= –1/3(–10.6) = +3.5 kJ**

Total $\qquad$ **2 XO₂ (s) + CO (g) → X₂O₃ (s) + CO₂ (g)** $\qquad$ **ΔH= –55.0 kJ**