

Energy

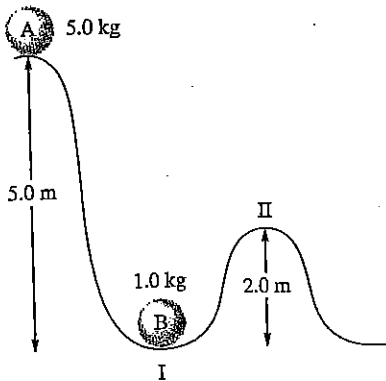
Key

9. Explain why  $\Delta H$  is obtained directly from a coffee-cup calorimeter, whereas  $\Delta E$  is obtained directly from a bomb calorimeter.

Coffee cup -  $\Delta H = Q$  because it is under const. P conditions

Bomb -  $\Delta H = \Delta E$  because it is under const V conditions

20. Consider the accompanying diagram. Ball A is allowed to fall and strike ball B. Assume that all of ball A's energy is transferred to ball B, at point I, and that there is no loss of energy to other sources. What is the kinetic energy and the potential energy of ball B at point II? The potential energy is given by  $PE = mgz$ , where  $m$  is the mass in kilograms,  $g$  is the gravitational constant ( $9.8 \text{ m/s}^2$ ), and  $z$  is the distance in meters.



$$PE_A = mgh = (5)(9.8)(5 \text{ m}) = 245 \text{ J}$$

$$PE_B + KE_B @ II = 245 \text{ J}$$

$$PE_B = mgh = (1)(9.8)(2) = 19.6 \text{ J}$$

$$KE_B = 245 - 19.6 = 225.4 \text{ J}$$

$$\text{Ball B} = KE = \frac{1}{2}mv^2$$

$$v = \sqrt{\frac{2KE}{m}} = \sqrt{\frac{2(225.4)}{1 \text{ kg}}}$$

$$v = 21.2 \text{ m/s} \quad \text{Ball B at Hill 2}$$

23. A gas absorbs 45 kJ of heat and does 29 kJ of work. Calculate  $\Delta E$ .

$$Q = +45 \text{ J}$$

$$W = -29 \text{ J}$$

$$\Delta E = Q + W$$

$$= 45 \text{ J} - 29 \text{ J} = \boxed{16 \text{ J}}$$

24. A system releases 125 kJ of heat while 104 kJ of work is done on it. Calculate  $\Delta E$ .

$$Q = -125 \text{ kJ}$$

$$W = +104 \text{ kJ}$$

$$\Delta E = Q + W = -125 \text{ kJ} + 104 \text{ kJ} = \boxed{-21 \text{ kJ}}$$

25. The volume of an ideal gas is decreased from 5.0 L to 0.05 L at a constant pressure of 2.0 atm. Calculate the work associated with this process.

$$V_i = 5 \text{ L}$$

$$V_f = 0.05 \text{ L}$$

$$P = 2 \text{ atm}$$

$$\text{work} = -(P\Delta V) = -(2 \text{ atm})(0.05 - 5) = +9.99 \text{ Latm}$$

$$9.99 \text{ Latm} \left( \frac{101.3 \text{ J}}{\text{Latm}} \right) = \frac{1011.987 \text{ J}}{1.01 \times 10^3 \text{ J}}$$

27. A balloon filled with 39.1 mol helium has a volume of 876 L at 0.0°C and 1.00 atm pressure. The temperature of the balloon is increased to 38.0°C as it expands to a volume of 998 L, the pressure remaining constant. Calculate  $q$ ,  $w$ , and  $\Delta E$  for the helium in the balloon. (The molar heat capacity for helium gas is 20.8 J/°C · mol.)

$$V_i = 876 \text{ L}$$

$$V_f = 998 \text{ L}$$

$$P = 1 \text{ atm}$$

$$n = 39.1 \text{ mol}$$

$$t_i = 0^\circ \text{C}$$

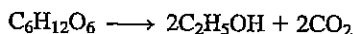
$$t_f = 38^\circ \text{C}$$

$$w = -P(V_f - V_i) = -1 \text{ atm}(998 - 876) = -122 \text{ L atm} \left( \frac{101.3 \text{ J}}{1 \text{ L atm}} \right) = -12,358.6 \text{ J}$$

$$Q = mc\Delta t = (39.1 \text{ mol})(20.8 \text{ J/}^\circ\text{C} \cdot \text{mol})(38^\circ\text{C} - 0^\circ\text{C}) = 30,904.6 \text{ J}$$

$$\Delta E = Q + w = 30,904.6 \text{ J} - 12,358.6 \text{ J} = 18,546.0 \text{ J} = 18.5 \text{ kJ}$$

29. The equation for the fermentation of glucose to alcohol and carbon dioxide is



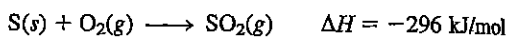
The enthalpy change for the reaction is -67 kJ. Is the reaction exothermic or endothermic? Is energy, in the form of heat, absorbed or evolved as the reaction occurs?

The reaction is exothermic ( $\Delta H = -$ )  
Heat is evolved (given off) in reaction.

31. Are the following processes exothermic or endothermic?
- When solid KBr is dissolved in water, the solution gets colder.
  - Natural gas ( $\text{CH}_4$ ) is burned in a furnace.
  - When concentrated  $\text{H}_2\text{SO}_4$  is added to water, the solution gets very hot.
  - Water is boiled in a teakettle.

- a) KBr in  $\text{H}_2\text{O}$  is endothermic. (KBr is absorbing heat from solution)  
b) Burning is exothermic. (Heat is being released)  
c)  $\text{H}_2\text{SO}_4$  to water is exothermic ( $\text{H}_2\text{SO}_4$  is releasing heat to surroundings)  
d)  $\text{H}_2\text{O}$  boiling is endothermic. ( $\text{H}_2\text{O}$  is absorbing heat from stove)

33. For the reaction



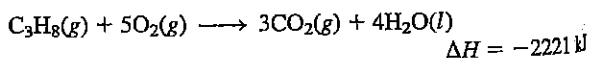
- How much heat is evolved when 275 g sulfur is burned in excess  $\text{O}_2$ ?
- How much heat is evolved when 25 mol sulfur is burned in excess  $\text{O}_2$ ?
- How much heat is evolved when 150. g sulfur dioxide is produced?

$$a) 275 \text{ g S} \left( \frac{1 \text{ mole S}}{32 \text{ g}} \right) \left( \frac{-296 \text{ kJ}}{\text{mole}} \right) = -2540 \text{ kJ}$$

$$b) 25 \text{ mol S} \left( \frac{-296 \text{ kJ}}{\text{mole}} \right) = -7400 \text{ kJ}$$

$$c) 150 \text{ g SO}_2 \left( \frac{1 \text{ mol SO}_2}{64 \text{ g}} \right) \left( \frac{-296 \text{ kJ}}{\text{mol}} \right) = -694 \text{ kJ}$$

35. Consider the combustion of propane:



Assume that all the heat in Sample Exercise 6.3 comes from the combustion of propane. What mass of propane must be burned to furnish this amount of energy?  $Q_{\text{sample}} = 1.3 \times 10^8 \text{ J}$

$$1.3 \times 10^8 \text{ J} \left( \frac{1 \text{ kJ}}{1000 \text{ J}} \right) \left( \frac{\text{mol C}_3\text{H}_8}{-2221 \text{ kJ}} \right) \left( \frac{44 \text{ g}}{\text{mol C}_3\text{H}_8} \right) = 2575 \text{ g C}_3\text{H}_8$$

# Calorimetry

37. The specific heat capacity of aluminum is  $0.900 \text{ J/}^\circ\text{C} \cdot \text{g}$ .

- Calculate the energy needed to raise the temperature of a  $8.50 \times 10^2\text{-g}$  block of aluminum from  $22.8^\circ\text{C}$  to  $94.6^\circ\text{C}$ .
- Calculate the molar heat capacity of aluminum.

$$a) \quad Q = mc\Delta t = (8.50 \times 10^2 \text{ g})(0.900 \text{ J/g}^\circ\text{C})(94.6 - 22.8) = \boxed{54.93 \text{ kJ}}$$

$$b) \quad 0.900 \frac{\text{J}}{\text{g}^\circ\text{C}} \left( \frac{27 \text{ g}}{1 \text{ mol Al}} \right) = \boxed{24.3 \frac{\text{J}}{\text{mol} \cdot ^\circ\text{C}}}$$

39. It takes  $78.2 \text{ J}$  to raise the temperature of  $45.6 \text{ g}$  lead by  $13.3^\circ\text{C}$ . Calculate the specific heat capacity and molar heat capacity of lead.

$$Q = mc\Delta t$$

$$c = \frac{Q}{m\Delta t} = \frac{78.2 \text{ J}}{(45.6 \text{ g})(13.3^\circ\text{C})} = \underline{0.129 \text{ J/g}^\circ\text{C}}$$

$$0.129 \frac{\text{J}}{\text{g}^\circ\text{C}} \left( \frac{207.2 \text{ g}}{1 \text{ mol Pb}} \right) = \underline{26.7 \text{ J/mol} \cdot ^\circ\text{C}}$$

43. A  $28.2\text{-g}$  sample of nickel is heated to  $99.8^\circ\text{C}$  and placed in a coffee-cup calorimeter containing  $150.0 \text{ g}$  water at  $23.5^\circ\text{C}$ . After the metal cools, the final temperature of metal and water is  $25.0^\circ\text{C}$ . Calculate the specific heat capacity of nickel, assuming that no heat escapes to the surroundings or is transferred to the calorimeter.

$$-Q_{\text{Ni}} = Q_{\text{w}}$$

$$-m_{\text{Ni}}c_{\text{Ni}}\Delta t_{\text{Ni}} = m_{\text{w}}c_{\text{w}}\Delta t_{\text{w}}$$

$$-(28.2 \text{ g})(c_{\text{Ni}})(25.0^\circ\text{C} - 99.8^\circ\text{C}) = (150 \text{ g})(4.184 \text{ J/g}^\circ\text{C})(25 - 23.5)$$

$$\boxed{c_{\text{Ni}} = 0.446 \text{ J/g}^\circ\text{C}}$$

(\*) Mass of Solution = mass water + mass KBr

Calorimetry

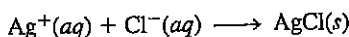
45. A coffee-cup calorimeter initially contains 125 g water at 24.2°C. Potassium bromide (10.5 g), also at 24.2°C, is added to the water, and after the KBr dissolves, the final temperature is 21.1°C. Calculate the enthalpy change for dissolving the salt in J/g and kJ/mol. Assume that the specific heat capacity of the solution is 4.18 J/°C · g and that no heat is transferred to the surroundings or to the calorimeter.

$m_w = 125g$      $m_{KBr} = 10.5g$      $m_{sol'n} = 135.5g$   
 $t_i H_2O = 24.2^\circ C$      $t_i KBr = 24.2^\circ C$   
 $t_f = 21.1^\circ C$

$Q_{w, KBr} = -Q_{sol'n} = -(135.5)(4.18)(21.1 - 24.2)$   
 $Q_{KBr} = 1757.5J$

$\Delta H = \frac{Q}{g} = \frac{1757.5J}{10.5g KBr} = 167.4 J/g$   
 $\Delta H = \frac{Q}{mole} = \frac{1757.5J}{\frac{10.5g}{199g/mole} \cdot \frac{1000J}{1kJ}} = 19.9 kJ/mole$

47. In a coffee-cup calorimeter, 50.0 mL of 0.100 M AgNO<sub>3</sub> and 50.0 mL of 0.100 M HCl are mixed to yield the following reaction:



$AgNO_3 = (0.05L)(0.1M) = .005 mol AgNO_3 = .005 mol Ag^+$   
 $HCl = (0.05L)(0.1M) = .005 mol HCl = .005 mol Cl^-$

The two solutions were initially at 22.60°C, and the final temperature is 23.40°C. Calculate the heat that accompanies this reaction in kJ/mol of AgCl formed. Assume that the combined solution has a mass of 100.0 g and has a specific heat capacity of 4.18 J/°C · g.

$t_i = 22.60^\circ C$      $Q = mc\Delta t = (100g)(4.18 J/g^\circ C)(23.4 - 22.6) =$   
 $t_f = 23.40^\circ C$      $Q = 334.4J$   
 $m = 100g$      $c = 4.18 J/g^\circ C$

$\Delta H = \frac{334.4J}{.005 mol} \cdot \frac{1kJ}{1000J} = 66.9 kJ/mol$

50. A 0.1964-g sample of quinone (C<sub>6</sub>H<sub>4</sub>O<sub>2</sub>) is burned in a bomb calorimeter that has a heat capacity of 1.56 kJ/°C. The temperature of the calorimeter increases by 3.2°C. Calculate the energy of combustion of quinone per gram and per mole.

$0.1964g C_6H_4O_2 \left( \frac{1 mol}{108g} \right) = 1.82 \times 10^{-3} mol C_6H_4O_2$

$Q_{rxn} = -Q_{Bomb} = -4.992 kJ$

$Q_{Bomb} = C_{Bomb} \Delta t = 1.56 \frac{kJ}{^\circ C} (3.2^\circ C) = 4.992 kJ$

$\Delta H = \frac{-4.992 kJ}{0.1964g} = -25.4 kJ/g$

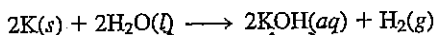
$\Delta H = \frac{-4.992 kJ}{1.82 \times 10^{-3} mol} = -2740 \frac{kJ}{mol}$

76. A swimming pool, 10.0 m by 4.0 m, is filled with water to a depth of 3.0 m at a temperature of 20.2°C. How much energy is required to raise the temperature of the water to 30.0°C?

$V = 10m \times 4m \times 3m = 120 m^3 \left( \frac{1000L}{1m^3} \right) \left( \frac{1000cm^3}{1L} \right) \left( \frac{1g}{cm^3} \right) = 1.2 \times 10^8 g$

$Q = mc\Delta t = (1.2 \times 10^8 g)(4.184 J/g^\circ C)(30 - 20.2) = 4.92 \times 10^9 J$

77. Calculate  $\Delta H^\circ$  for the reaction



A 5.00-g chunk of potassium is dropped into 1.00 kg water at 24.0°C. What is the final temperature of the water after the preceding reaction occurs? Assume that all the heat is used to raise the temperature of the water. (Never run this reaction. It is very dangerous; it bursts into flame!)

$\Delta H = 2(-481) - 2(-286) = -390 kJ/mole$

$5g K \left( \frac{1 mol K}{39.1g} \right) \left( \frac{-390 kJ}{mole} \right) = -498.72 kJ$

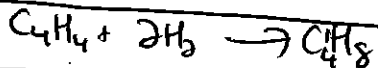
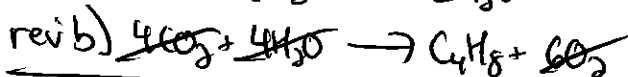
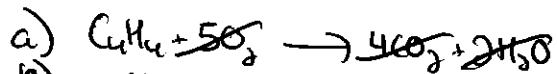
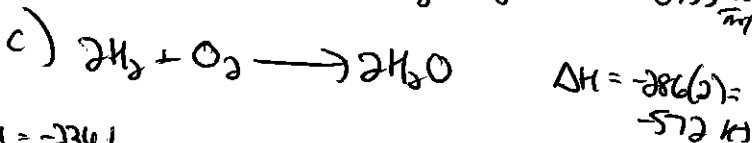
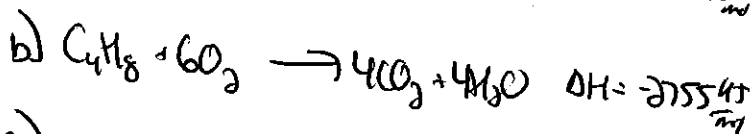
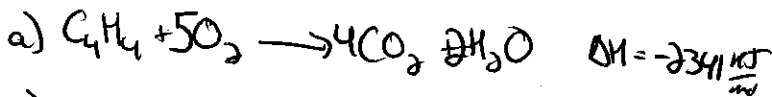
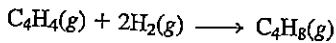
$\Delta t = t_f - t_i$      $-11.9 = t_f - 24$

$Q_w = -Q_r = +498.72J$   
 $t_f = 12.1^\circ C$

$\Delta t = \frac{Q}{mc} = \frac{498.72J}{(1000g)(4.184)} = 11.9^\circ C$

# Hess' Law

52. Combustion reactions involve reacting a substance with oxygen. When compounds containing carbon and hydrogen are combusted, carbon dioxide and water are the products. Using the enthalpies of combustion for  $C_4H_4$  ( $-2341 \text{ kJ/mol}$ ),  $C_4H_8$  ( $-2755 \text{ kJ/mol}$ ), and  $H_2$  ( $-286 \text{ kJ/mol}$ ), calculate  $\Delta H$  for the reaction



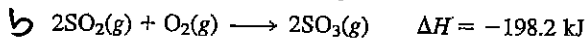
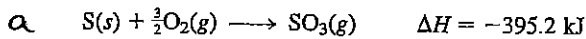
$$\Delta H = -2341$$

$$\Delta H = -572$$

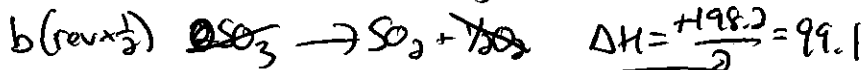
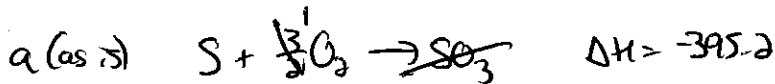
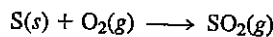
$$\Delta H = +2755$$

$$\Delta H = -158 \text{ kJ/mol}$$

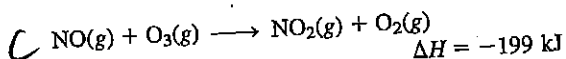
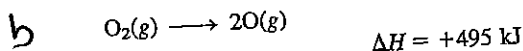
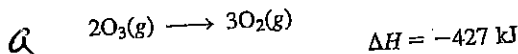
53. Given the following data:



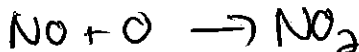
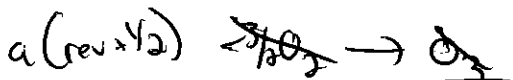
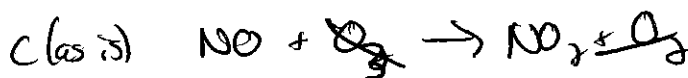
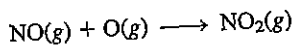
calculate  $\Delta H$  for the reaction



55. Given the following data:



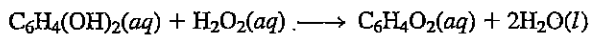
calculate  $\Delta H$  for the reaction



$$\Delta H = -233$$

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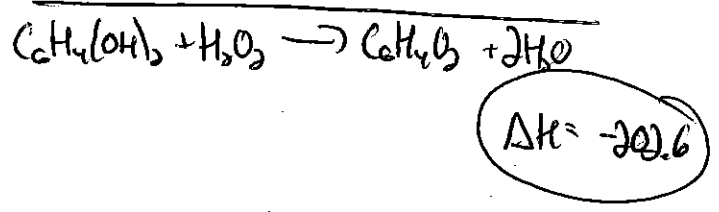
56. The bombardier beetle uses an explosive discharge as a defensive measure. The chemical reaction involved is the oxidation of hydroquinone by hydrogen peroxide to produce quinone and water:



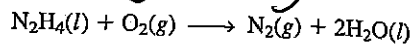
Calculate  $\Delta H$  for this reaction from the following data:

- a  $C_6H_4(OH)_2(aq) \rightarrow C_6H_4O_2(aq) + H_2(g)$   $\Delta H = +177.4 \text{ kJ}$
- b  $H_2(g) + O_2(g) \rightarrow H_2O_2(aq)$   $\Delta H = -191.2 \text{ kJ}$
- c  $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(g)$   $\Delta H = -241.8 \text{ kJ}$
- d  $H_2O(g) \rightarrow H_2O(l)$   $\Delta H = -43.8 \text{ kJ}$

- a) as is  $C_6H_4(OH)_2 \rightarrow C_6H_4O_2 + H_2$   $\Delta H = +177.4$
- b (rev)  $H_2O_2 \rightarrow H_2 + O_2$   $\Delta H = +191.2$
- d (x2)  $2H_2O(g) \rightarrow 2H_2O(l)$   $\Delta H = 2(-43.8)$
- e (x2)  $2H_2O(g) \rightarrow 2H_2 + O_2$   $\Delta H = 2(-241.8)$



58. Calculate the  $\Delta H$  for the reaction



given the following data:

- a  $2NH_3(g) + 3N_2O(g) \rightarrow 4N_2(g) + 3H_2O(l)$   $\Delta H = -1010 \text{ kJ}$
- b  $N_2O(g) + 3H_2(g) \rightarrow N_2H_4(l) + H_2O(l)$   $\Delta H = -317 \text{ kJ}$
- c  $2NH_3(g) + \frac{1}{2}O_2(g) \rightarrow N_2H_4(l) + H_2O(l)$   $\Delta H = -143 \text{ kJ}$
- d  $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$   $\Delta H = -286 \text{ kJ}$

Handwritten solution for problem 58:

- (a)  $\frac{1}{4} \times [2NH_3 + 3N_2O \rightarrow 4N_2 + 3H_2O] \quad \Delta H = \frac{-1010}{4} = -252.5$
- (b)  $\frac{1}{4} \times [N_2O + 3H_2 \rightarrow N_2H_4 + H_2O] \quad \Delta H = \frac{-317}{4} = -79.25$
- (c)  $\frac{1}{4} \times [2NH_3 + \frac{1}{2}O_2 \rightarrow N_2H_4 + H_2O] \quad \Delta H = \frac{-143}{4} = -35.75$
- (d)  $\frac{1}{4} \times [2H_2 + O_2 \rightarrow 2H_2O] \quad \Delta H = \frac{-286}{2} = -143$

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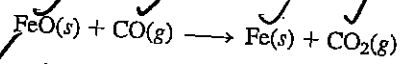

$$N_2H_4 + O_2 \rightarrow N_2 + 2H_2O$$

$\Delta H = -622.5 \text{ kJ}$

82. Given the following data:

- a  $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$   $\Delta H^\circ = -23 \text{ kJ}$
- b  $3Fe_2O_3(s) + CO(g) \rightarrow 2Fe_3O_4(s) + CO_2(g)$   $\Delta H^\circ = -39 \text{ kJ}$
- c  $Fe_3O_4(s) + CO(g) \rightarrow 3FeO(s) + CO_2(g)$   $\Delta H^\circ = +18 \text{ kJ}$

calculate  $\Delta H^\circ$  for the reaction



- (a)  $FeO + \frac{1}{3}CO_2 \rightarrow \frac{1}{3}Fe_2O_3 + \frac{1}{3}CO$
- (b)  $\frac{1}{3}Fe_2O_3 + \frac{1}{3}CO_2 \rightarrow \frac{1}{3}Fe_2O_3 + \frac{1}{3}CO$
- (c)  $\frac{1}{3}Fe_2O_3 + \frac{1}{3}CO \rightarrow Fe + \frac{1}{3}CO_2$

Handwritten solution for problem 82:

- c  $\frac{-18}{3} = -6$
- a  $\frac{-23}{3} = -7.67$

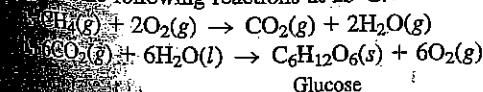
$\Delta H = -11 \text{ kJ}$

# CH 16 $\Delta H, \Delta S + \Delta G$ Problems (p. 814-816)

27. Given the values of  $\Delta H$  and  $\Delta S$ , which of the following changes will be spontaneous at constant  $T$  and  $P$ ?
- $\Delta H = +25 \text{ kJ}$ ,  $\Delta S = +5.0 \text{ J/K}$ ,  $T = 300. \text{ K}$
  - $\Delta H = +25 \text{ kJ}$ ,  $\Delta S = +100. \text{ J/K}$ ,  $T = 300. \text{ K}$
  - $\Delta H = -10. \text{ kJ}$ ,  $\Delta S = +5.0 \text{ J/K}$ ,  $T = 298 \text{ K}$
  - $\Delta H = -10. \text{ kJ}$ ,  $\Delta S = -40. \text{ J/K}$ ,  $T = 200. \text{ K}$
28. At what temperatures will the following processes be spontaneous?
- $\Delta H = -25 \text{ kJ}$  and  $\Delta S = -5.0 \text{ J/K}$
  - $\Delta H = +25 \text{ kJ}$  and  $\Delta S = +5.0 \text{ J/K}$
  - $\Delta H = +25 \text{ kJ}$  and  $\Delta S = -5.0 \text{ J/K}$
  - $\Delta H = -25 \text{ kJ}$  and  $\Delta S = +5.0 \text{ J/K}$

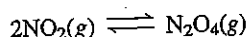
37. Predict the sign of  $\Delta S^\circ$  and then calculate  $\Delta S^\circ$  for each of the following reactions.
- $\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{g})$
  - $3\text{O}_2(\text{g}) \rightarrow 2\text{O}_3(\text{g})$
  - $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}(\text{g})$

Using data in Appendix 4, calculate  $\Delta H^\circ$ ,  $\Delta S^\circ$ , and  $\Delta G^\circ$  for each of the following reactions at  $25^\circ\text{C}$ .



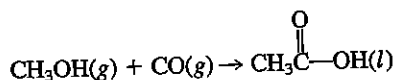
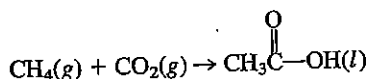
- $\text{P}_4\text{O}_{10}(\text{s}) + 6\text{H}_2\text{O}(\text{l}) \rightarrow 4\text{H}_3\text{PO}_4(\text{s})$
- $\text{HCl}(\text{g}) + \text{NH}_3(\text{g}) \rightarrow \text{NH}_4\text{Cl}(\text{s})$

46. For the reaction at  $298 \text{ K}$ ,



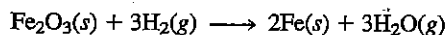
the values of  $\Delta H^\circ$  and  $\Delta S^\circ$  are  $-58.03 \text{ kJ/mol}$  and  $-176.6 \text{ J/K} \cdot \text{mol}$ , respectively. What is the value of  $\Delta G^\circ$  at  $298 \text{ K}$ ? Assume that  $\Delta H^\circ$  and  $\Delta S^\circ$  do not depend on temperature. At what temperature is  $\Delta G^\circ = 0$ ? Is  $\Delta G^\circ$  negative above or below this temperature?

47. Using data from Appendix 4, calculate  $\Delta H^\circ$ ,  $\Delta S^\circ$ , and  $\Delta G^\circ$  for the following reactions that produce acetic acid:



Which reaction would you choose as a commercial method for producing acetic acid ( $\text{CH}_3\text{CO}_2\text{H}$ ) at standard conditions? What temperature conditions would you choose for the reaction? Assume  $\Delta H^\circ$  and  $\Delta S^\circ$  do not depend on temperature.

54. Consider the reaction



- Use  $\Delta G_f^\circ$  values in Appendix 4 to calculate  $\Delta G^\circ$  for this reaction.
- Is this reaction spontaneous under standard conditions at  $298 \text{ K}$ ?
- The value of  $\Delta H^\circ$  for this reaction is  $100. \text{ kJ}$ . At what temperatures is this reaction spontaneous at standard conditions? Assume that  $\Delta H^\circ$  and  $\Delta S^\circ$  do not depend on temperature.

Solutions on following pages.

21. When 2.0 g of  $\text{CaCl}_2$  are dissolved in 100 mL of water, the temperature of the solution went up to 24.5°C.

- 27)
- a)  $\Delta G = \Delta H - T\Delta S = (25 \text{ kJ}) - (300 \text{ K})(.005 \text{ kJ/mol}\cdot\text{K}) = 23.5 \frac{\text{kJ}}{\text{mol}}$  - Non
  - b)  $\Delta G = \Delta H - T\Delta S = (25 \text{ kJ}) - (300 \text{ K})(.100 \text{ kJ/mol}\cdot\text{K}) = -5 \text{ kJ/mol}$  - Spont
  - c)  $\Delta G = \Delta H - T\Delta S = (-10 \text{ kJ}) - (298 \text{ K})(-.005 \text{ kJ/mol}\cdot\text{K}) = -11.5 \frac{\text{kJ}}{\text{mol}}$  - Spont
  - d)  $\Delta G = \Delta H - T\Delta S = (-10 \text{ kJ}) - (200 \text{ K})(-.040 \text{ kJ/mol}\cdot\text{K}) = -2 \frac{\text{kJ}}{\text{mol}}$  - Spont

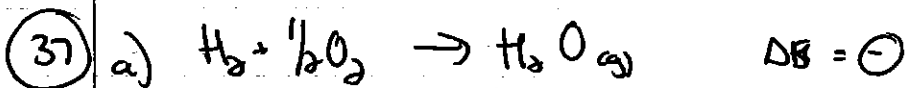
28) a)  $T = \frac{\Delta H}{\Delta S} = \frac{-25 \text{ kJ}}{-.005 \frac{\text{kJ}}{\text{mol}\cdot\text{K}}} = 5000 \text{ K}$  ~~is~~ Lower than 5000 K - Spont.

$\frac{\Delta H - \Delta G}{\Delta S} > 0$

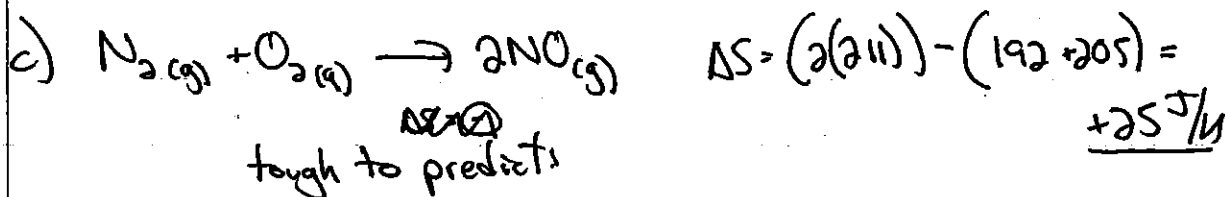
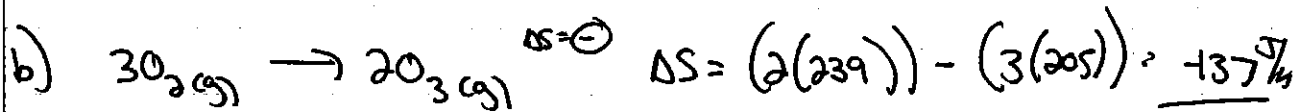
b)  $T = \frac{\Delta H}{\Delta S} = \frac{25 \text{ kJ}}{.005} = 5000 \text{ K}$  Higher than 5000 K - Spont.

c)  $T = \frac{\Delta H}{\Delta S} = \frac{25 \text{ kJ}}{-.005 \frac{\text{kJ}}{\text{mol}\cdot\text{K}}} = -5000 \text{ K}$  - Never Spontaneous (unfavorable  $\Delta H + \Delta S$ )

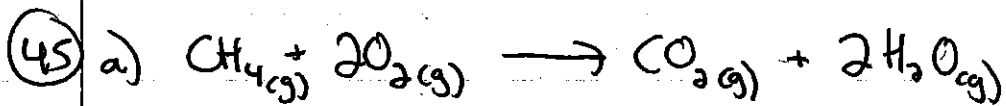
d)  $T = \frac{\Delta H}{\Delta S} = \frac{-25 \text{ kJ}}{.005 \frac{\text{kJ}}{\text{mol}\cdot\text{K}}} = -5000 \text{ K}$  - Always Spontaneous (favorable  $\Delta H + \Delta S$ )



$\Delta S = 189 - \left( \frac{131}{1} + \frac{1}{2}(205) \right) = -44.5 \frac{\text{J}}{\text{mol}\cdot\text{K}}$





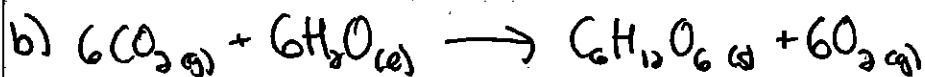


$$\Delta H = (-393.5 + 2(242)) - ((-75) + 2(0)) = -802.5 \text{ kJ/mole}$$

$$\Delta S = (214 + 2(\frac{189}{2})) - ((186) + 2(205)) = -4 \text{ J/mol}\cdot\text{K}$$

$$\Delta G = (-394 + 2(-229)) - ((-51) + 2(0)) = -801 \text{ kJ/mol}$$

$\Delta G = \Delta H - T\Delta S = -801$

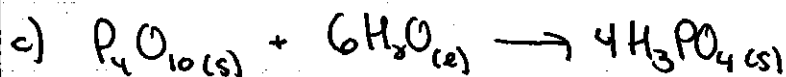


$$\Delta H = (-1275 + 6(0)) - (6(-393.5) + 6(-286)) = +2802 \text{ kJ/mol}$$

$$\Delta S = (212 + 6(205)) - (6(214) + 6(70)) = -262 \text{ J/mol}\cdot\text{K}$$

$$\Delta G = (-911 + 6(0)) - (6(-394) + 6(-237)) = 2875 \text{ kJ/mol}$$

$$\Delta G = \Delta H - T\Delta S = 2802 - (298)(-262) = 2880 \text{ kJ/mol}$$

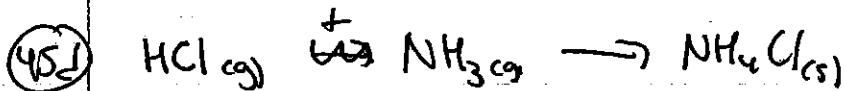


$$\Delta H = (4(-1279)) - (-2984 + 6(-286)) = -416 \text{ kJ/mole}$$

$$\Delta S = (4(110)) - (229 + 6(70)) = -209 \text{ J/mol}\cdot\text{K}$$

$$\Delta G = (4(-1119)) - (-2698 + 6(-237)) = -356 \text{ kJ/mol}$$

$$\Delta G = \Delta H - T\Delta S = -416 - (298)(-209) = -354$$

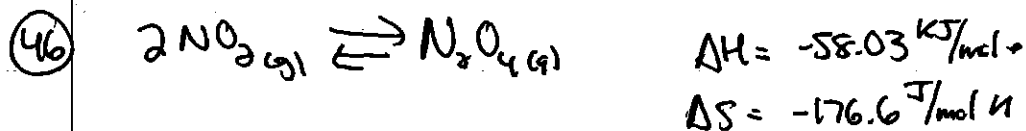


$$\Delta H = (-314) - (-92 + (-46)) = -176 \text{ kJ/mol}$$

$$\Delta S = (96) - (187 + 193) = -284 \text{ J/mol}\cdot\text{K}$$

$$\Delta G = (-203) - (-95 + (-17)) = -91 \text{ kJ/mol}$$

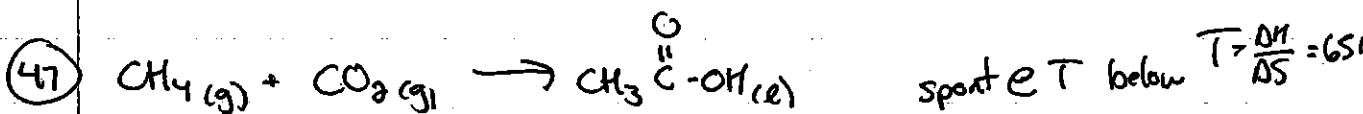
$$\Delta G = \Delta H - T\Delta S = -176 - (298)(-284) = -91.4 \text{ kJ/mol}$$



$$\Delta G = ? @ 298 \quad \Delta G = \Delta H - T\Delta S = -58.03 - (298)(-176.6) = -5.40 \text{ kJ/mol}$$

$$\Delta G = 0 \quad T = ? \quad T = \frac{\Delta H - \Delta G}{\Delta S} = \frac{\Delta H}{\Delta S} = \frac{-58.03 \text{ kJ/mol}}{-176.6 \text{ J/mol}\cdot\text{K}} = 328.5 \text{ K}$$

$\Delta G =$  negative @ temps BELOW 328.5 K



$$\Delta H = -484 - (-75 + (-393.5)) = -15.5 \text{ kJ}$$

$$\Delta S = 160 - (186 + 214) = -240 \text{ J/mol}\cdot\text{K}$$

spontaneous only @ Low Temps

$$\Delta G = \Delta H - T\Delta S$$

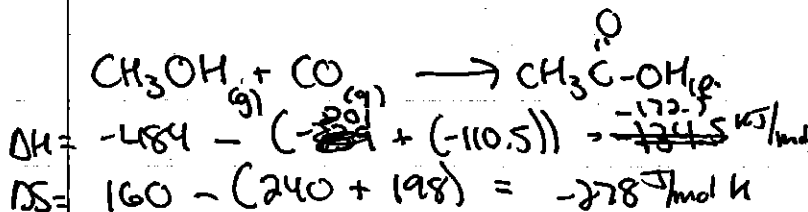
$$\Delta G = -15.5 - (298)(-240) =$$

$$\Delta G = 56.02 \text{ kJ/mol}$$

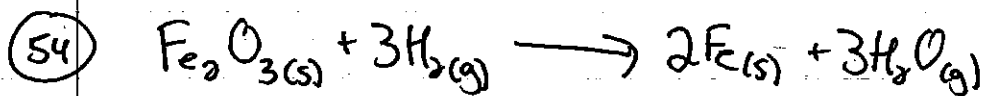
spont @ T below  $T = \frac{\Delta H}{\Delta S} = 620 \text{ K}$

$$\Delta G = -172.5 - (298)(-278)$$

$$\Delta G = -89.6 \text{ kJ/mol}$$



More Feasible



a)  $\Delta G = (2(0) + 3(-229)) - (\overset{-740 + 3(0)}{\cancel{229} + 3\cancel{229}}) = 53 \text{ kJ/mole}$

b) Non-spontaneous  $\Delta G = (+)$

c)  $\Delta G = \Delta H - T\Delta S$        $\Delta S = \frac{\Delta G - \Delta H}{-T} > \frac{\Delta H - \Delta G}{T} = \frac{100 - 53}{298} = .157 \text{ kJ/mole}$

$T = \frac{\Delta H}{\Delta S} = \frac{100 \text{ kJ}}{.157 \text{ kJ/mole}} = 634 \text{ K}$       Spont. at temps above 634 K