

Version A

Name _____

Section _____

The last page contains potentially relevant equations that may be of use to you. You are welcome to tear that page off and keep it. There are 7 pages, be sure you have all pages **before** you begin.

Useful Information

Gas Constant, $R = 8.3 \text{ J}/(\text{mol K}) = 0.082 \text{ liter-atm}/(\text{mol K})$

$$101 \frac{\text{J}}{(\ell - \text{Atm})}$$

$$R \cdot T = 2.48 \frac{\text{kJ}}{\text{mole}} \quad @ T = 298\text{K}$$

$$T(\text{K}) = T(\text{C}) + 273$$

Gravitational Constant, $g = 9.8 \text{ m}/\text{sec}^2$ Avagadro's Number, $N_A = 6.0 \cdot 10^{23} \text{ molecules}/\text{mole}$ Boltzmann's Constant, $k_B = 1.4 \cdot 10^{-23} \text{ J}/\text{K}$ $R = N_A \cdot k_B$

Problems	1	2	3	4	5
Points	18	24	22	24	22

Total Points: xxx

Suggestions:

- 1) Make sure you set up all calculation problems for the entire test, and then go back and finish them. You will get more points covering the full test demonstrating how the problem may be solved. (Hold off using your calculator.)
- 2) Some additional algebra, using variables rather than numbers, will save you time calculating unneeded intermediate quantities.

Q1) (Z9.20) a) (18 pts) Consider a mixture of air and gasoline vapor in a cylinder with a piston. The original volume is 50cm^3 . If the combustion of the mixture releases 1000 J of energy, to what volume will the gasses expand against a constant external pressure of 1 Atm if all the energy of the combustion is converted into work to push back the piston?

Heat is converted to work:

$$\frac{1 \cdot 10^3 J}{101 \frac{J}{\ell} - \text{Atm}} = 10 = w = P\Delta V = 1 \cdot \Delta V$$

$$\Delta V = 10\ell = V - 0.05$$

$$V = 10.05\ell$$

b) (Z9.87) The sun supplies energy at a rate of about 1.0 kilowatt per square meter of surface areal (1Watt = 1 J/second). Solar panels are about 15% efficient in converting sunlight to electricity. A typical home will use about 40 kiloWatt-hours of electricity per day or $144 \cdot 10^6 J$ of energy in one day. Assume 8 hours of useful sunlight per day. Calculate the minimum solar panel surface area necessary to provide the entire home's electricity. A typical house has about 100 square meters of roof area, is there enough area to place solar panels on the roof to meet a home's need?

Two ways to do this: Compute the total amount of energy per meter squared per day and compare with 144 MJ, or just compare kilowatt hours.

$$E / m^2 / \text{day} = 0.15 \cdot 1 \cdot 10^3 \cdot 8 \frac{\text{hr}}{\text{day}} \cdot (60)^2 \frac{\text{sec}}{\text{hr}} = 3 \cdot 1.44 \cdot 10^6 J / m^2 / \text{day}$$

$$\text{Area} = \frac{144 \cdot 10^6}{3 \cdot 1.44 \cdot 10^6} = \frac{100}{3} = 33m^2$$

$$\text{The easier way is: } \text{Area} = \frac{40}{0.15 \cdot 8} = \frac{1}{0.03} = 33m^2$$

In any event the panels will easily fit on he roof; Yes there is enough area.

Q2) (Z9.27) (24 pts) Consider 10 moles of ethane in a container. The heat capacity of ethane $C_v = 40 \text{ J/mole-K}$, and is an ideal gas.

a) Calculate $\Delta E, \Delta H, w, q$ for the change in the gas heated from 30°C to 80°C at constant volume.

$$\Delta E = nC_v\Delta T = 10 \cdot 40 \cdot 50 \text{ J} = 20 \text{ kJ}$$

$$\Delta H = nC_p\Delta T = 10 \cdot 48 \cdot 50 \text{ J} = 24 \text{ kJ}$$

$$w = 0 = -P\Delta V$$

$$q = q_v = \Delta E = 20 \text{ kJ}$$

b) Calculate the same quantities $\Delta E, \Delta H, w, q$ for the change as the gas is heated from 30°C to 80°C but now at constant pressure.

$$\Delta E = nC_v\Delta T = 10 \cdot 40 \cdot 50 \text{ J} = 20 \text{ kJ}$$

$$\Delta H = nC_p\Delta T = 10 \cdot 48 \cdot 50 \text{ J} = 24 \text{ kJ}$$

$$w = \Delta E - q = 20 - 24 = -4 \text{ kJ}$$

$$q = q_p = \Delta H = 24 \text{ kJ}$$

Q3) (Z9.4,10.1,10.51) (22 pts)

a) Ice turns to liquid water (As a general reaction $A(s) \rightleftharpoons A(l)$.) Is this process exothermic or endothermic and what is the sign of ΔH for this change?

The process is endothermic; heat goes in; the sign of the enthalpy is positive.

b) The enthalpy of fusion for ice is: $\Delta H = 6 \text{ kJ/mole}$, for the process in part a, and at 1 Atm it occurs at 0°C , what is the entropy change for this transformation?

$$\Delta G = \Delta H - T\Delta S = 0$$

At phase transition:

$$\Delta S = \frac{\Delta H}{T} = \frac{6 \cdot 10^3}{273} = 22 \text{ J/K - mole}$$

c) Explain what is happening in terms of system and surrounding (include the entropy changes):

The system need heat to break the bonds, this causes the entropy of the surrounds to drop. However the entropy of the system increases because liquid water can have more places to go to. So the reaction is favored by system entropy and disfavored by surrounding entropy (or system enthalpy).

d) Explain how the heat is transferred during the transformation.

The surroundings must supply the heat energy, which goes into the ice (system) to transform it to water.

e) If you wanted to favor the process as written, would you raise or lower the temperature. Explain your choice.

Raise the temperature to increase system entropy contribution; From $\Delta G = \Delta H - T\Delta S$, get a larger negative Gibbs energy by increasing temperature. Besides everybody knows water melts at higher temperatures. This is just recasting the usual observation in the general terms of thermodynamics.

Q4) (Z10.26) (24 pts) a) One mole of an ideal gas is in a 1 liter volume container at a pressure of 5 Atm. It is allowed to expand isothermally into an evacuated bulb to give a total volume of 2 liters. Calculate w and q for this change; also calculate the reversible heat transfer (q_{rev}) and the entropy change for this change.

$$q = -w = 0$$

$$w = -P_{ext} \Delta V = 0 \cdot 1 = 0$$

$$q_{rev} = -w_{rev} = nRT \ln \frac{V_2}{V_1} = 0.5 \ln 2 = 0.35 \text{ kJ}$$

$$nRT = PV = 5 \cdot 1 \text{ l} \cdot \text{atm} \cdot 101 \frac{\text{J}}{\text{l} \cdot \text{atm}} = 0.5 \text{ kJ}$$

$$\Delta S = \frac{q_{rev}}{T} = nR \ln 2 = 1 \cdot 8.3 \cdot 0.7 = 5.8 \frac{\text{J}}{\text{K}}$$

b) (Z10.27) The molar heat capacity of CO_2 is $C_v = 30 \frac{\text{J}}{(\text{mole} \cdot \text{K})}$, you may assume $C_p = C_v + R$. The molar entropy under standard conditions (1 Atm, 25C) is $S^\circ = 200 \frac{\text{J}}{\text{mole} \cdot \text{K}}$. What is the molar entropy at 400K, and 1 Atm?

This is a constant pressure change in temperature:

$$\Delta S = C_p \ln \frac{T_f}{T_i} = 38 \cdot \ln \left(\frac{400}{298} \right) = 38 \cdot .3 = 11 \frac{\text{J}}{(\text{mole} \cdot \text{K})}$$

$$S = S^\circ + \Delta S = 211 \frac{\text{J}}{(\text{mole} \cdot \text{K})}$$

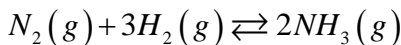
c) (Z10.7d) What should the sign of ΔS be for the case of heating an ideal gas at constant pressure? Explain.

$$\Delta S = \frac{q_{rev}}{T} \quad \text{or} \quad \Delta S = nC_p \ln \frac{T_f}{T_i}$$

The sign should be positive because entropy change is equal to the reversible heat transferred at constant temperature; or when changing the temperature the entropy increases because more states are accessible. When putting heat in (either at constant pressure or temperature) heat goes in so q is positive so

entropy change must be positive.

Q5) (Z10.63) (22 pts) Using the following data, calculate ΔH° , ΔS° , and K @T=298K for the synthesis of ammonia (by the Haber Process) for the reaction:



Species	ΔH_f° kJ/mole	S° J/K-mole
$N_2(g)$	----	200
$H_2(g)$	-----	130
$NH_3(g)$	-50	200

$$\Delta H^\circ = -50 \cdot 2 = -100 \text{ kJ}$$

$$\Delta S^\circ = 2 \cdot 200 - 200 - 3 \cdot 130 = -190 \text{ J/k}$$

$$\ln K = -\frac{\Delta G^\circ}{RT} = -\frac{\Delta H^\circ - T\Delta S^\circ}{RT} = -\frac{-100 + 0.298 \cdot 190}{2.48} = -\frac{-43.4}{2.48} = 17.5$$

$$K = 10^8$$

b) Is the reaction spontaneous under standard conditions?

Both Gibbs energy is negative or K is much larger than 1, indicate the reaction is definitely spontaneous.

c) Calculate ΔG_{rxn} for this reaction when the partial pressure of each of the gasses is: $P_{N_2} = P_{H_2} = P_{NH_3} = 1 \cdot 10^{-4}$ Atm at T=298K.

$$\begin{aligned} \Delta G_{rxn} &= \Delta G_{rxn}^\circ + RT \ln Q = -43.4 + 2.48 \ln \left(\frac{P^2}{P^4} \right) = -43.4 - 2 \cdot 2.48 \ln(P) \\ &= -43.4 - 2 \cdot 2.48 \ln(10^{-4}) = -43.4 + 8 \cdot 2.48 \cdot \ln 10 = 2.3 \text{ kJ}/_{rxn} \end{aligned}$$

d) Is the reaction spontaneous under these conditions? (Explain)

The reaction is not spontaneous anymore, there are too few reactants, so Q is larger than K or the Gibbs Free energy of reaction is positive.

Relevant Equations

$\Delta E = q + w$ $K.E. = \frac{1}{2}mv^2$ $P.E. = mgh$ $PV = nRT$ $\Delta H = \Delta E + \Delta(PV)$ $P = \frac{F}{A}$ $\Delta E = nC_V \Delta T$ $\Delta H = nC_P \Delta T$ $C_P = C_V + R$ $\Delta S = \frac{q_{rev}}{T}$ $w = -P_{ext} \Delta V$ $w_{rev} = -nRT \ln \frac{V_f}{V_i}$ $\Delta S = nR \ln \frac{V_f}{V_i}$	$\Delta S = \int \frac{dq_{rev}}{T}$ $S = k \ln \Omega$ $\Delta H = \Delta H_{rxn} \Delta X$ $\Delta G = \Delta H - T \Delta S$ $\Delta G^\circ = \Delta H^\circ - T \Delta S^\circ$ $\Delta H_{rxn}^\circ = \sum c_i \Delta H_f^\circ(i)$ $\Delta S_{rxn}^\circ = \sum c_i S_f^\circ(i)$ $\Delta G_{rxn}^\circ = \sum c_i \Delta G_f^\circ(i)$ $\Delta G_{rxn} = \Delta G_{rxn}^\circ + RT \ln Q$ $\ln K = -\frac{\Delta G_{rxn}^\circ}{RT}$ $\Delta S = nC_V \ln \frac{T_f}{T_i}$ $\Delta S = nC_P \ln \frac{T_f}{T_i}$
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