

**Chem 155 First Hour Exam 2001**

**Version II: Answers provided**

Name: \_\_\_\_\_

Full credit will be given to correct answers only when all necessary steps are shown.

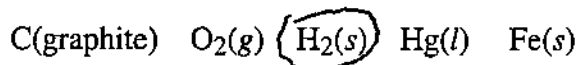
(12) 1. True or false? Just write "T" or "F" after each question.

(a)  $\Delta H$  is a state function. **F**(b) P-V work is usually negligible for solids and liquids. **T**(c)  $\Delta S = 0$  for an adiabatic, reversible expansion of an ideal gas. **T**(d)  $C_v$  is always temperature independent for an ideal gas. **F**(e) In a cyclic process such as  $1 \rightarrow 2$  followed by  $2 \rightarrow 1$ , we have  $\Delta G = 0$ . **T**(f)  $\left(\frac{\partial U}{\partial V}\right)_T = 0$  for an ideal gas. **T**

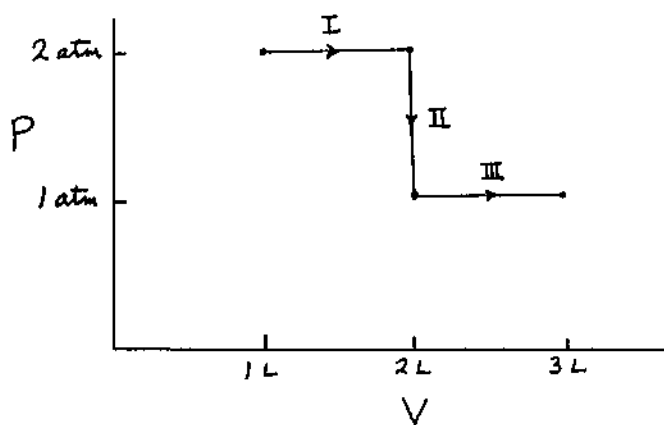
(12) 2. List the condition(s) under which each of the following equations can be applied. If no conditions are required, write "none."

(a)  $\Delta H = \Delta U + \Delta(PV)$ *None*(b)  $P_1 V_1^\gamma = P_2 V_2^\gamma$ *Ideal gas, adiabatic, reversible*(c)  $w = -nRT \ln \frac{V_2}{V_1}$ *Ideal gas, isothermal, reversible*(d)  $\gamma = \frac{\bar{C}_P}{\bar{C}_V} = \frac{5}{3}$ *Ideal gas, monatomic*(e)  $\Delta S = \frac{\Delta H}{T}$ *Constant pressure, at equilibrium*(f)  $\bar{S}^\circ = 0$  at 0 K.*Perfect, pure crystalline*

(4) 3. Circle the only one of the following substances that has a nonzero  $\Delta_f \bar{H}^\circ$  value at 25°C,



(6) 4. A gas undergoes an irreversible expansion along the path shown below. Calculate the work done in joules. (*Hint*: Work done is given by area under curve.)



$$\text{Work done} = -P\Delta V$$

$$\text{Step I: } -(2 \text{ atm})(2-1) \text{ L} = -2 \text{ L atm}$$

$$\text{Step II: } \text{zero because } \Delta V = 0$$

$$\text{Step III: } -(1 \text{ atm})(3-2) \text{ L} = -3 \text{ L atm}$$

$$\begin{aligned} \text{Total work done} &= -3 \text{ L atm} \\ &= -3 \times 10^2 \text{ J} \end{aligned}$$

(8) 5. The internal energy (U) of one mole of a monatomic ideal gas is  $(3/2)RT$ . Calculate  $\Delta H$  of the gas in joules when the temperature of the gas is increased from 300 K to 400 K.

$$\begin{aligned} H &= U + PV \\ &= \frac{3}{2}RT + nRT \\ \Delta H &= \frac{3}{2}R\Delta T + nR\Delta T \\ &= 1 \frac{3}{2} (8.314 \text{ J K}^{-1} \text{ mol}^{-1}) (400-300) \text{ K} + (1 \text{ mol}) (8.314 \text{ J K}^{-1} \text{ mol}^{-1}) \times (400-300) \text{ K} \\ &= 2.08 \times 10^3 \text{ J} \end{aligned}$$

(6) 6. List three processes that can increase the entropy of a system.

Gas expansion  
Mixing of gases (or liquids/solids)  
Heating

(12) 7. In class, we focused mostly on gas expansions. Now consider the reversible, isothermal compression of 0.50 mole of an ideal gas from 2.0 L to 1.0 L at 298 K.

(a) Calculate  $w$ ,  $q$ ,  $\Delta U$ ,  $\Delta H$ , and  $\Delta S$  for the process. (b) Can you still describe the work done in this case as the *maximum* work done? Explain.

(a)  $\Delta U = 0$ ,  $\Delta H = 0$  for an isothermal process involving an ideal gas

$$w = -nRT \ln \frac{V_2}{V_1} = -(0.50 \text{ mol})(8.314 \text{ J K}^{-1} \text{ mol}^{-1})(298 \text{ K}) \ln \frac{1.0}{2.0} = +8.6 \times 10^2 \text{ J}$$

$$q = -w = -8.6 \times 10^2 \text{ J}$$

$$\Delta S = nR \ln \frac{V_2}{V_1} = (0.50 \text{ mol})(8.314 \text{ J K}^{-1} \text{ mol}^{-1}) \ln \frac{1.0}{2.0} = -2.9 \text{ J K}^{-1}$$

(b) No. The work done is the minimum amount because the opposing pressure is always infinitesimally less than the external pressure.

(10) 8. 16.2 moles of Ar and 20.2 moles of Xe are mixed at constant temperature (298 K) and pressure. What is the total entropy of the system after mixing? Assume ideal behavior.

There are two contributions to the total entropy

$$1. \Delta_{\text{mix}} S = -R(n_{\text{Ar}} \ln x_{\text{Ar}} + n_{\text{Xe}} \ln x_{\text{Xe}})$$

$$n_{\text{Ar}} = \frac{16.2}{16.2 + 20.2} = 0.445 \quad n_{\text{Xe}} = \frac{20.2}{16.2 + 20.2} = 0.555$$

$$\Delta_{\text{mix}} S = -(8.314 \text{ J K}^{-1} \text{ mol}^{-1}) [(16.2 \text{ mol}) \ln 0.445 + (20.2 \text{ mol}) \ln 0.555]$$
$$= 208 \text{ J K}^{-1}$$

2.  $\bar{S}^\circ(\text{Ar})$  and  $\bar{S}^\circ(\text{Xe})$

$$S_{\text{total}} = [208 + (16.2)(154.8) + (20.2)(169.6)]$$
$$= 6.14 \times 10^3 \text{ J K}^{-1}$$

(16) 9. For each of the following processes, state whether the following quantities listed are positive (+), negative (-), or zero (0). All the changes refer to the system.

(a) Melting ethanol. (molar volume of liquid ethanol > molar volume of solid ethanol)

(b) Heating a gas at constant volume.

(c) Irreversible, adiabatic expansion of an ideal gas.

(d) Burning hydrogen in air. [  $2\text{H}_2(g) + \text{O}_2(g) \longrightarrow 2\text{H}_2\text{O}(l)$  ]

	q	w	$\Delta U$	$\Delta H$
(a)	+	-	+	+
(b)	+	0	+	+
(c)	0	-	-	-
(d)	-	+	-	-

(6) 10. Ammonium nitrate ( $\text{NH}_4\text{NO}_3$ ) dissolves spontaneously and endothermically in water. Deduce the sign of  $\Delta S$  (of the system) for this process. Does your result make sense physically? Explain.

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

$$\Delta S = \frac{\Delta H - \Delta G}{T}$$

$$\Delta G < 0 \quad (\text{spontaneous process})$$

$$\Delta H > 0 \quad (\text{endothermic})$$

$$\therefore \Delta S > 0$$

- (8) 11. The  $\Delta_f \bar{G}^\circ$  of  $\text{ClF}(\text{g})$  is  $-56.0 \text{ kJ mol}^{-1}$ . (a) Calculate the equilibrium constant for the reaction at 298 K:



- (b) What is  $\Delta G$  for the reaction if the partial pressures of  $\text{Cl}_2$ ,  $\text{F}_2$ , and  $\text{ClF}$  are 1.4 atm, 0.90 atm, and 0.35 atm, respectively. Is the reaction more or less favored under these conditions?

$$\begin{aligned} (a) \quad \Delta G^\circ &= (2) \Delta_f \bar{G}^\circ(\text{ClF}) - (1) \Delta_f \bar{G}^\circ(\text{Cl}_2) - (1) \Delta_f \bar{G}^\circ(\text{F}_2) \\ &= (2) (-56.0 \text{ kJ mol}^{-1}) = -112 \text{ kJ mol}^{-1} \end{aligned}$$

$$\Delta G^\circ = -RT \ln K$$

$$\begin{aligned} K &= e^{-\Delta G^\circ/RT} = e^{-(-112 \times 10^3)/(8.314)(298)} \\ &= 4.3 \times 10^{19} \end{aligned}$$

$$(b) \quad \Delta_r G = \Delta_r G^\circ + RT \ln Q$$

$$\begin{aligned} &= -112 \times 10^3 \text{ J mol}^{-1} + (8.314)(298) \ln \frac{0.35^2}{(1.4)(0.90)} \\ &= -112 \times 10^3 \text{ J mol}^{-1} - 5.77 \times 10^3 \text{ J mol}^{-1} \\ &= -1.18 \times 10^5 \text{ J mol}^{-1} = -1.18 \times 10^2 \text{ kJ mol}^{-1} \end{aligned}$$

The reaction will be more spontaneous.

I have neither given nor received aid on this examination.

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**Useful information**

$$R = 8.314 \text{ J K}^{-1} \text{ mol}^{-1}$$

$$1 \text{ L atm} = 101.3 \text{ J}$$

At 298 K:

$$\bar{S}^\circ(\text{Ar}) = 154.8 \text{ J K}^{-1} \text{ mol}^{-1}$$

$$\bar{S}^\circ(\text{Xe}) = 169.6 \text{ J K}^{-1} \text{ mol}^{-1}$$