

2018 Fall Semester Quiz 6
For General Chemistry I (CH101)

Date: December 3 (Mon), Time: 19:00 ~ 19:45

Professor Name	Class	Student I.D. Number	Name
			Solution

Gas constant, R:
8.315 J K⁻¹ mol⁻¹
0.08206 L atm mol⁻¹ K⁻¹

T/K = T/°C + 273.15

1. (Total 6 pts, 1 pt each) Fill the table with the correct answer.

Molecule	H ₂ O	CO ₂	C ₂ H ₄
Number of total degrees of freedom	9	9	18
Number of vibrational modes	3	4	12

2. (Total 5 pts) To vaporize 100.0 g carbon tetrachloride (CCl₄) at its normal boiling point, 349.9 K, and P = 1 atm, 19.5 kJ of heat is required. (The molecular weight of CCl₄ is 153.8 g/mol.)

(a) (2 pts) What is the molar enthalpy of vaporization (ΔH_{vap}) for CCl₄?

The molar mass of CCl₄ is 153.8 g mol⁻¹; thus, the number of moles in 100.0 g is

$$\frac{100.0 \text{ g CCl}_4}{153.8 \text{ g mol}^{-1}} = 0.6502 \text{ mol CCl}_4$$

The enthalpy change for 1 mol CCl₄ is then

$$\left(\frac{19.5 \text{ kJ}}{0.6502 \text{ mol CCl}_4} \right) \times 1.00 \text{ mol CCl}_4 = 30.0 \text{ kJ} = \Delta H_{\text{vap}} \quad (+2 \text{ pts})$$

(b) (3 pts) Calculate ΔU for the same process.

The internal energy change is then

$$\Delta U = \Delta H_{\text{vap}} - \Delta(PV) = \Delta H_{\text{vap}} - RT \Delta n_g \quad (+1 \text{ pt})$$

Inserting $T = 349.9 \text{ K}$ and $\Delta n_g = 1$ (because there is an increase of 1 mol of gaseous products for each mole of liquid that is vaporized) gives

$$\begin{aligned} \Delta U &= 30.0 \text{ kJ} - (8.315 \text{ J K}^{-1} \text{ mol}^{-1})(349.9 \text{ K})(1.00 \text{ mol})(10^{-3} \text{ kJ J}^{-1}) \\ &= (30.0 - 2.9) \text{ kJ} = +27.1 \text{ kJ mol}^{-1} \quad (+2 \text{ pts}) \end{aligned}$$

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3. (Total 10 pts) Consider an **ideal gas** enclosed in a piston-cylinder arrangement that is maintained at **constant temperature** ($\Delta T = 0$) in a heat bath. The gas can be expanded **reversibly** by changing the position of the piston to accomplish a specified change in volume. Considering U for an ideal gas was shown to depend only on the temperature, **show that** $\Delta S_{\text{sys}} = nR \ln \frac{V_f}{V_i} > 0$ ($V_f > V_i$).

$\Delta U = 0$; for any **isothermal** ideal gas process

$$\therefore w = -q \quad (+2 \text{ pt})$$

$$dw = -P_{\text{ext}} dV \quad (+1 \text{ pt})$$

$P_{\text{ext}} = P$; for **reversible** isothermal expansions of an ideal gas (+1 pts)

$$w = - \int_{V_i}^{V_f} P_{\text{ext}} dV = - \int_{V_i}^{V_f} P dV$$

$$w = - \int_{V_i}^{V_f} \frac{nRT}{V} dV \quad (\text{ideal gas law; } +3 \text{ pts})$$

$$w = -nRT \int_{V_i}^{V_f} \frac{1}{V} dV = -nRT \ln \frac{V_f}{V_i}$$

$$q = nRT \ln \frac{V_f}{V_i}$$

$$\Delta S = \int_i^f \frac{dq_{\text{rev}}}{T} = \frac{1}{T} \int_i^f dq_{\text{rev}} = \frac{q_{\text{rev}}}{T} \quad (\text{definition of entropy; } +3 \text{ pts})$$

$$\therefore \Delta S = nR \ln \frac{V_f}{V_i}$$

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4. (Total 8 pts, 2 pts each) Ethanol's enthalpy of vaporation is 38.7 kJ mol^{-1} at its normal boiling point, 78°C . Calculate q , ΔU , ΔS_{sys} , and ΔG when **1.00 mol ethanol is vaporized reversibly at 78°C and 1 atm** . Assume that the vapor is an ideal gas and neglect the volume of liquid ethanol relative to that of its vapor.

*Be careful with the unit and sign.

q (kJ)	ΔU (kJ)	ΔS_{sys} (J K ⁻¹)	ΔG (kJ)
38.7	35.8	110	0

$q = \Delta H = 38.7 \text{ kJ}$ when 1.00 mol of ethanol is vaporized at constant pressure.

$$\Delta S = \frac{q_{\text{rev}}}{T} = \frac{38.7 \text{ kJ}}{351.15 \text{ K}} = 110 \text{ J K}^{-1}$$

$$\Delta G = \Delta H - T\Delta S = 38.7 \text{ kJ} - (351.15 \text{ K})(0.110 \text{ kJ K}^{-1}) = 0 \text{ kJ}$$

$$\Delta U = \Delta H - \Delta(PV) = \Delta H - P(V_2 - V_1)$$

$$V_2 = \frac{nRT}{P} = \frac{(1.00 \text{ mol})(0.08206 \text{ L atm mol}^{-1}\text{K}^{-1})(351.15 \text{ K})}{1.00 \text{ atm}} = 28.8 \text{ L}$$

Since V_1 is negligible, $P(V_2 - V_1) \approx PV_2 = 28.8 \text{ atm L} = 2.92 \text{ kJ}$

$$\therefore \Delta U = 38.7 \text{ kJ} - 2.92 \text{ kJ} = 35.8 \text{ kJ}$$

Partial points:

+1 pt for the correct equation for each value

5. **[Problem 13-30] (Total 5 pts)** Solid Tin exists in two forms: white and gray. For the transformation, $\text{Sn}(s, \text{white}) \rightarrow \text{Sn}(s, \text{gray})$, the enthalpy change is -2.1 kJ and the entropy change is -7.4 J K^{-1} .

- (a) **(2 pts)** Calculate the Gibbs free energy change for the conversion of 1.00 mol white tin to gray tin at $-30 \text{ }^\circ\text{C}$.

+ 1 pt for correct equation; +2 pts for correct answer

- (b) **(1 pt)** Will white tin convert spontaneously to gray tin at $-30 \text{ }^\circ\text{C}$?

+1 pt for correct answer

- (c) **(2 pts)** At what temperature are white and gray tin in equilibrium at a pressure of 1 atm?

+ 1 pt for correct equation; +2 pts for correct answer

6. **(Total 4 pts) Answer the following questions.**

- (a) **(2 pts, 1 pt each)** Calculate **the heat and the work** associated with a process in which 5.00 mol of gas expands **reversibly** at constant temperature $T = 298 \text{ K}$ from a pressure of 10.00 to 1.00 atm.

$$\frac{V_f}{V_i} = \frac{P_i}{P_f} = 10.0$$

$$w = -nRT \ln \left(\frac{V_f}{V_i} \right) = -28.5 \text{ kJ}$$

$$q = -w = 28.5 \text{ kJ}$$

+1 pt each for correct answers

- (b) **(2 pts, 1 pt each)** Calculate **the heat absorbed and the work done on a system** of 5.00 mol of an ideal gas as it expands **irreversibly** at constant temperature $T = 298 \text{ K}$ from a pressure of 10.0 to 1.00 atm. The external pressure is held constant at 1.00 atm.

The initial volume V_1 is

$$V_1 = \frac{nRT}{P_1} = \frac{(5.00 \text{ mol})(0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1})(298 \text{ K})}{10.0 \text{ atm}} = 12.2 \text{ L}$$

The final volume is 10 times this, or

$$V_2 = 122 \text{ L}$$

For a constant external pressure,

$$w_{\text{irrev}} = -P_{\text{ext}} \Delta V = -(1.00 \text{ atm})(122 \text{ L} - 12.2 \text{ L}) = -110 \text{ L atm}$$

$$= -11.1 \text{ kJ}$$

At constant T , $\Delta U = 0$, however, so

$$q_{\text{irrev}} = -w_{\text{irrev}} = 11.1 \text{ kJ}$$

+1 pt each for correct answers