Name:

Chem 3321 test #1 practice questions

I want complete, detailed answers to the questions. Show all your work to get full credit.

DATA:

R = 0.082 L atm mol^{-1} K^{-1} = 8.314 J mol^{-1} K^{-1}

1 atm = 1.01 × 10^5 Pa = 1.01 bar = 760 torr

1 mL = 1 cm^3

1 Pa = 1 N m^{-2}

0 °C = 273 K

Avogadro's constant = $N_A = 6.022 \times 10^{23}$ mol^{-1}

$\pi = 3.14159$

1 eV = 1.6022×10^{-19} J (electron volt to joule conversion)

mass conversion from amu to kg: 1.66×10^{-27} kg/amu

$C_P - C_V = nR$ for any ideal gas

$\Delta U = 0$; $\delta w = -\delta q$ when $\Delta T = 0$

$-q = w = -nRT \ln\left(\frac{V_2}{V_1}\right)$

at $\Delta T = 0$, $\frac{V_2}{V_1} = \frac{P_1}{P_2}$

$\overline{C_V} = \frac{3}{2} R$ for monoatomic

$PV = nRT$

$\overline{C_V}$ $\frac{5}{2} R$ for diatomic

$q = m C_p \Delta T$

$w = -P_{ext} \Delta V$

$\Delta U = w + q$

$$\Delta S_{fusion} = \frac{\Delta H_{fusion}}{T_{fusion}}$$

$$\Delta S = \int C_p \frac{dT}{T}$$

$S = \frac{3}{2} R \ln T + R \ln V$ or $\Delta S = \frac{3}{2} nR \Delta T + nR \ln \Delta V$

$$\frac{\Delta S_{mix}}{R} = -X_{Ne} \ln X_{Ne} - X_{He} \ln X_{He}$$

$\Delta S = \frac{q_{rev}}{T}$ and $q_{rev} \overset{\text{when } \Delta T = 0}{=} -w = nRT \int_{V_1}^{V_2} \frac{dV}{V}$

* Don't forget your calculator.
or notecard.

Problem 1

At atmospheric pressure, liquid water and ice exist in equilibrium at 0 °C. At this temperature, $\Delta H_{\text{fusion}} = 6.00$ kJ/mol. (Fusion is the phase transition from solid to liquid.) Also at this temperature, the heat capacity of ice is 38.0 J K^{-1} mol^{-1} and the heat capacity of water is 76.0 J K^{-1} mol^{-1}.

a) Calculate the molar entropy of fusion of H_2O at 0 °C. 273K

b) At −10 °C (263 K), calculate the molar entropy of freezing (opposite of fusion) of H_2O. In other words, calculate the molar entropy change when supercooled water undergoes a phase transition to ice. Assume that the heat capacities of each phase are constant in the −10 °C to 0 °C temperature range.

a)

$$\Delta S_{\text{fusion}} = \frac{\Delta H_{\text{fusion}}}{T_{\text{fusion}}} = \frac{6.00\,\text{kJ/mol}^{-1}}{273\text{K}} = 22\ \text{J}\cdot\text{K}^{-1}\cdot\text{mol}^{-1}$$

b)

$$\Delta S = \int_{263}^{273} C_{p(l)} \frac{dT}{T} - 22\ \text{J}\cdot\text{K}^{-1}\cdot\text{mol}^{-1} + \int_{273}^{263} C_{p(s)} \frac{dT}{T}$$

$$= C_{p(l)} \int_{263}^{273} \frac{dT}{T} - 22 \frac{\text{J}}{\text{K}\cdot\text{mol}} + C_{p(s)} \int_{273}^{263} \frac{dT}{T}$$

$$\left[= C_{p(l)} \ln\left(\frac{273}{263}\right) - 22 + C_{p(s)} \ln\left[\frac{263}{273}\right) \right]$$

$$\ln\left(\frac{263}{273}\right) = -\ln\left(\frac{273}{263}\right)$$

$$= -20.6 \frac{\text{J}}{\text{K}\cdot\text{mol}}$$

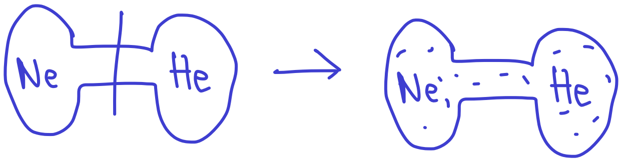
~~$= \left(76.0 \frac{\text{J}}{\text{K}\cdot\text{mol}} - 38.0 \frac{\text{J}}{\text{K}\cdot\text{mol}}\right) \ln\left(\frac{273}{263}\right) - 22 = -20.6 \frac{\text{J}}{\text{K}\cdot\text{mol}}$~~

Problem 2

An insulated container is divided in half with one mole of neon on one side of a partition and one mole of helium on the other side. Both sides of the container are at 1 atm and 273 K. At some point the partition is removed and the gases mix.

a) Find ΔS_{Ne} and ΔS_{He} and the total entropy change for the system and the surroundings.

b) Comment from your result whether the process is spontaneous or not.



a) $\frac{\Delta S_{mix}}{R} = -X_{Ne} \ln X_{Ne} - X_{He} \ln X_{He} = \ln 2 = 0.693$

→ X = mole fraction = 1/2 for each one

Multiply by R and by 2 moles, we get

$\Delta S_{mix} = 2 \Delta S_{Ne} = 2 \Delta S_{He} = 11.52 \text{ J/K}$

Since this is adiabatic, $q = 0$ so $\Delta S_{surr} = 0$

b) The total entropy is positive, so it's Spontaneous

Problem 3

pressure is changing

One mole of a monoatomic ideal gas initially at T= 400 K and V= 30.0 L has its pressure doubled through an isothermal process in which $w = 3.00$ kJ of work is performed on the gas.

a) Calculate $\Delta S = \Delta S_{\text{system}}$.

b) Calculate $\Delta S_{\text{surroundings}}$.

c) Is this process spontaneous or non-spontaneous? Support your answer with an explanation and/or calculations.

$\frac{V_2}{V_1} = \frac{P_1}{P_2}$

P is doubled, V is halved

a) Since isothermal, $\Delta T = 0$

1) $\Delta S = \cancel{\frac{3}{2} nR\Delta T} + R\ln\Delta V \Rightarrow R\ln\left(\frac{15}{30}\right) = -5.76\ \text{J/K}$

$\Delta S = R\ln\Delta V$

or

2) $\Delta S = \frac{q_{rev}}{T}$ and $q_{rev} = -w = nRT\int_{V_1}^{V_2}\frac{dV}{V}$

b) $q + w = 0$ bc isothermal

$\Delta S_{surr} = \frac{-q_{sys}}{T} = \frac{w}{T} = \frac{3\text{ kJ}}{400\text{ K}} = 7.5\ \text{J/K}$

c) $\Delta S_{universe}$ is positive, so spontaneous

universe is isolated. $\Delta S_{sys} + \Delta S_{surr} = \Delta S_{universe} > 0$

Problem 4

1.50 mol of an ideal gas at 450 K is expanded from an an initial pressure of 5.00 bar to a final pressure of 1.00 bar. $\bar{C_P} = 5/2$ R. The expansion is isothermal and reversible. Calculate q and w.

$$\downarrow$$

$$\Delta T = 0$$

$$\Delta U = \delta w + \delta q$$

$$\Delta U = 0 \; ; \; \delta w = -\delta q$$

$$-q = w = -nRT \ln\left(\frac{V_2}{V_1}\right)$$

$$\text{at } \Delta T = 0, \quad \frac{V_2}{V_1} = \frac{P_1}{P_2}$$

$$-q = w = -nRT \ln\left(\frac{P_1}{P_2}\right)$$

$$-q = w = -1.5 \text{ mol} \cdot 8.3145 \frac{\text{J}}{\text{mol} \cdot \text{K}} \cdot 450 \text{ K} \cdot \ln\left(\frac{5 \text{ bar}}{1 \text{ bar}}\right)$$

$$-q = w = -9.03 \text{ kJ}$$