

12.E: Thermodynamic Processes (Exercises)

These are homework exercises to accompany the Textmap created for "Principles of Modern Chemistry" by Oxtoby et al. Complementary General Chemistry question banks can be found for other Textmaps and can be accessed [here](#).

Q1A

A sample of O_2 gas is under an external pressure of 17 atm and contained in a cylinder with a volume of 50 L. The gas is cooled and the resulting volume is 25 L. Calculate the amount of work done **on** the O_2 gas.

Solution

$$\begin{aligned} w &= -P\Delta V \\ &= -P_{\text{ext}}(V_2 - V_1) \\ &= -17 \text{ atm} (25 \text{ L} - 50 \text{ L}) = 425 \text{ L atm} \\ &= 425 \text{ atm L} (101.325 \frac{\text{J}}{\text{atm L}}) \\ &= 43,063 \text{ J} \\ &= 43.06 \text{ kJ} \end{aligned}$$

Q1B

A system containing oxygen gas is heated at a constant pressure of 40.0 atm so that its volume increases 177 L to 458 L. Express the amount of work that the system did in kilo-joules.

Solution

The formula for work from the expansion of a gas at constant pressure

$$\begin{aligned} w &= -P_{\text{ext}} \Delta V \\ w &= -40 \text{ atm} \times (458 \text{ L} - 177 \text{ L}) = -11240 \text{ L atm} \end{aligned}$$

Convert from L atm to joules

$$-11240 \text{ L atm} \times 101.325 \frac{\text{J}}{\text{L atm}} = -1138893 \text{ J}$$

Convert from J to kJ and round to get the final answer

$$\begin{aligned} -1138893 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} &= -1140 \text{ kJ} \\ w &= -1140 \text{ kJ} \end{aligned}$$

Q2

The gas mixture inside one of the cylinders of an airplane expands against a constant external pressure of 5.00 atm because of the growing altitude, from an initial volume of 500 mL (at the end of the compression stroke) to the final volume of 1200 mL. Calculate the work done on the gas mixture during this process and express it in Joules.

Solution

$$w = -(5.00 \text{ atm})(1200 \text{ mL} - 500 \text{ mL}) \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) (101.3 \frac{\text{J}}{\text{L} \times \text{atm}}) = -354.55 \text{ J}$$

Q3

The following events are based on a true story. A butcher for the local Chinese restaurant needs to defrost a large chunk of beef, which weights 50 lb and is currently at 0°C . He wants to accomplish this by is going out onto the street and repeatedly drop it onto the ground. Suppose the potential energy of the meat completely transforms into heat each time it hits the ground, and that energy can be calculated from

$$V = mg\Delta h$$

where m is mass of the object, g is the acceleration of gravity, and Δh is change in height. If the man is 5 ft tall and he wants to get the meat to room temperature (25°C), how many times does he have to drop the piece of meat? Assume the environment has no effect on the meat, and that it does not lose any heat. (Specific Heat of the Meat: $0.25 \frac{\text{J}}{\text{g}^\circ\text{C}}$).

Solution

There are two steps to solve this equation. First, to find the amount of heat absorbed by the meat from one drop. Second, to find the total amount of heat needed to warm the meat up from $0^{\circ}C$ to $25^{\circ}C$.

To find the total amount of heat required to bring the meat from $0^{\circ}C$ to $25^{\circ}C$, the equation for calculating heat has to be used:

$$Q = mC_p\Delta T$$

The mass of the meat can be calculated via conversion from lb to g , and is as follows:

$$50lb \times 453.592 \frac{g}{lb} = 22679.6g$$

The change in temperature is just the final temperature minus the initial, so it is just $25^{\circ}C$.

Lastly, C_p is given. Plugging in all of these values gives:

$$Q = 22679.6g \times 0.25 \frac{J}{g^{\circ}C} \times 25^{\circ}C = 141747.5J$$

Thus, $141747.5J$ is the total amount of heat needed to heat up the meat to $25^{\circ}C$.

Now to find the amount of heat transferred into the meat from one drop. This utilizes the potential energy equation. The question mentions how to calculate the potential energy, and all potential energy is translated into heat. Therefore.

$$Q = V = mg\Delta h$$

The problem gives the mass, the acceleration of gravity is $9.8 \frac{m}{s^2}$, and height. Therefore, since the potential energy for one drop is equal to the heat for one drop, the heat can be calculated as such:

$$Q = 22.6796kg \times 9.8 \frac{m}{s^2} \times 1.524m = 338.724 \frac{kg \cdot m^2}{s^2} = 338.724J$$

From there, since no heat is lost, it carries over for each drop. Therefore, the number of times he needs to drop the meat is found by simple division:

$$141747.5J \div 338.724J = 418.474drops$$

Thus, the drops needed is around 419.

Abstract: Calculate heat from $Q = mC_p\Delta T$ and $Q = V = mg\Delta h$. Divide the two.

Q4

Suppose you have a ball ($C_p = 0.85 \frac{J}{g^{\circ}C}$) at $25^{\circ}C$, what will its final temperature be if the amount of work equal to dropping it down from a height of $86.6m$ is done to it? ($g = 9.81 \frac{m}{s^2}$)

Solution

$$U = mg\Delta h \quad (12.E.1)$$

$$= m_{ball} \cdot 86.6m \cdot 9.81 \frac{m}{s^2} \quad (12.E.2)$$

$$= 850m_{ball} \quad (12.E.3)$$

If we assuming all the kinetic energy at the time of collision converts to heat, then

$$\Delta T = \frac{q}{C_p \cdot m_{ball}} \quad (12.E.4)$$

$$= \frac{U}{C_p \cdot m_{ball}} \quad (12.E.5)$$

$$= \frac{850m_{ball}}{85 \frac{J}{g^{\circ}C} \cdot m_{ball}} \cdot \frac{1 \text{ kg}}{1,000 g} \quad (12.E.6)$$

$$= 1.0^{\circ}C \quad (12.E.7)$$

$$T_{final} = 25^{\circ}C + 1^{\circ}C \quad (12.E.8)$$

$$= 26^{\circ}C \quad (12.E.9)$$

Q7A

The [rule of Dulong and Petit](#) shows that the molar heat capacities of most metallic elements group around a certain value "X" $\frac{J}{K \cdot mol}$ at $25^{\circ}C$. By first calculating and showing the molar heat capacities of metals rhenium, silver, lead, tungsten, copper, molybdenum and hafnium (given that the respective

specific heat capacities of these metals are $0.14 \frac{J}{K g}$, $0.23 \frac{J}{K g}$, $0.13 \frac{J}{K g}$, $0.13 \frac{J}{K g}$, $0.39 \frac{J}{K g}$, $0.25 \frac{J}{K g}$ and $0.14 \frac{J}{K g}$), find the value of "X". (hint: take the average value of the calculated molar heat capacities and round off value to nearest whole number).

Solution

To first find the molar heat capacities of each metal, multiply the molar mass of each metal by their specific heat capacities:

Molar heat capacity of:

$$Re = 0.14 \frac{J}{K g} \times 186.207 \frac{g}{mol} = 26.068 \frac{J}{K mol} \quad Ag = 0.23 \frac{J}{K g} \times 107.868 \frac{g}{mol} = 24.809 \frac{J}{K mol}$$

$$Pb = 0.13 \frac{J}{K g} \times 207.2 \frac{g}{mol} = 26.936 \frac{J}{K mol}$$

calculate the average molar heat capacity:

$$\text{Average} \quad \frac{1}{7} (26.068 \frac{J}{K mol} + 24.809 \frac{J}{K mol} + 26.936 \frac{J}{K mol} + 23.899 \frac{J}{K mol} + 24.7829 \frac{J}{K mol} + 23.99 \frac{J}{K mol} + 24.988 \frac{J}{K mol}) = 25.068 \frac{J}{K mol}$$

$$W = 0.13 \frac{J}{K g} \times 183.84 \frac{g}{mol} = 23.899 \frac{J}{K mol}$$

$$Cu = 0.39 \frac{J}{K g} \times 63.546 \frac{g}{mol} = 24.7829 \frac{J}{K mol}$$

$$Mo = 0.25 \frac{J}{K g} \times 95.96 \frac{g}{mol} = 23.99 \frac{J}{K mol}$$

$$Hf = 0.14 \frac{J}{K g} \times 178.49 \frac{g}{mol} = 24.988 \frac{J}{K mol}$$

Q7B

The specific heat capacities of metals aluminum, bismuth, copper, lead, and silver at 25°C are 0.900, 0.123, 0.386, 0.128, and $0.233 \frac{J}{g K}$. Calculate the molar heat capacities of these metals. According to the rule of Dulong and Petit, the molar heat capacities of metallic elements, like these, are approximately $25 \frac{J}{K mol}$.

Solution

Multiply each of the specific heat capacities by its corresponding molar mass to obtain their molar heat capacities.

Aluminum:

$$0.900 \frac{J}{g K} \times 26.98 \frac{g}{mol} = 24.3 \frac{J}{mol K}$$

Bismuth:

$$0.123 \frac{J}{g K} \times 208.98 \frac{g}{mol} = 25.7 \frac{J}{mol K}$$

Copper:

$$0.386 \frac{J}{g K} \times 63.55 \frac{g}{mol} = 24.5 \frac{J}{mol K}$$

Lead:

$$0.128 \frac{J}{g K} \times 207.2 \frac{g}{mol} = 26.5 \frac{J}{mol K}$$

Silver:

$$0.223 \frac{J}{g K} \times 107.87 \frac{g}{mol} = 25.1 \frac{J}{mol K}$$

Q9

An undisclosed volume of water is tightly sealed in a microwave-safe container at room temperature before it is placed in an ice bath where it is cooled by a student.

- During the cooling process within the ice bath, state whether (ΔU), Q , and W of the system are negative, zero, or positive. Explain your reasoning.

- b. After being cooled, the student decides that the water seems too cold so it is placed in the microwave where the container of water is heated back to room temperature. What are the new signs of (ΔU) , Q , and W during the heating process? Explain your reasoning.
- c. Now determine the signs of $(\Delta U_1 + \Delta U_2)$, $(Q_1 + Q_2)$, and $(W_1 + W_2)$, where possible, assuming that the cooling process was step 1 and the heating process was step 2.

Solution

- a. W is zero, Q is negative, and (ΔU) is negative. Work is zero for this step because neither volume nor pressure of the container is changing. Since the water is cooler than before heat has left the system so it is negative; and because $(\Delta U) = Q + W$ and $W = 0$ then $(\Delta U) = Q$ which is negative so, (ΔU) is negative.
- b. W is zero, Q is positive, and (ΔU) is positive. Work is zero for this step because neither volume nor pressure of the container is changing. Since the water is hotter than before heat has entered the system so it is positive and because $(\Delta U) = Q + W$ and $W = 0$ then $(\Delta U) = Q$ which is positive so, (ΔU) is positive.
- c. Since W was equal to zero for both processes above $(W_1 + W_2)$ is equal to zero. $((\Delta U)_1 + (\Delta U)_2) = (Q_1 + Q_2)$ where they are both equal to zero.

If the water is cooled from room temperature to a lower temperature and then heated directly back to room temperature, wouldn't $((\Delta U)_1 + (\Delta U)_2)$ be 0? Q is the only factor actually changing here and $q(\text{heating to cooling}) = -q(\text{cooling to heating})$ provided the change in t is the same. So both $((\Delta U)_1 + (\Delta U)_2)$ and q should be 0.

Q11A

Your lab partner slipped a sample of unknown hot metal that is 40.0g, which is initially at 130.0°C into a 100.0 g water that is initially at 50.0°C. A temperature probe indicates that equilibrium is reached at 60.15°C. Using the specific heat capacity of water $4.18 \frac{J}{K}$, calculate the specific heat capacity of the unknown metal.

Solution

Imagine two sub-systems: the metal and the water. If mixing the hot metal and cool water inside a well-insulated container (that prevent leaks of heat), then the heat absorbed by the system will equal zero. Since the system is the sum of the two sub-systems:

$$q_{\text{sys}} = 0 = q_{\text{metal}} + q_{\text{water}}$$

For both sub-systems, the amount of heat gained is equal to the specific heat capacity times the mass times the temperature change:

$$q_{\text{metal}} + q_{\text{water}} = m_{\text{water}} C_{s,\text{water}} \Delta T_{\text{metal}} = 0$$

Solving for the specific heat capacity of the metal:

$$c_{s,\text{metal}} = \frac{-m_{\text{water}} C_{s,\text{water}} \Delta T_{\text{water}}}{m_{\text{metal}} \Delta T_{\text{metal}}} = \frac{-100.0 \text{ g} \times 4.18 \frac{J}{K \text{ g}} \times 10.15^\circ C}{40.0 \text{ g} \times -69.85^\circ C} = 1.52 \frac{J}{K \text{ g}}$$

Tip: Do not convert Celsius to Kelvin. The Kelvin and Celsius scales differ only in their location of their zero points. Temperature change in Celsius is the same temperature change of Kelvin.

Q11B

Let's say you 34.5 grams of some hot metal that is initially at 75°C and you put that metal into 64.0 grams of water that is initially at 25°C. If the two objects reach thermal equilibrium at 39°C, what is the specific heat capacity of the metal when the specific heat capacity of water is $4.18 \frac{J}{K \text{ g}}$.

Solution

The formula used to solve this question is

$$m_1 c_1 \Delta T_1 = -m_2 c_2 \Delta T_2$$

Plug in your known values and solve for c_1

$$(34.15 \text{ g})(c_1)(36^\circ C) = -(64 \text{ g})(4.18 \frac{J}{K \text{ g}})(-14^\circ C)$$

$$c_1 = 3.02 \frac{J}{K \text{ g}}$$

Q12

A 10.00 g sample of Aluminum at 60.0 °C and a 30.0 g sample of copper at a temperature of -20.0 °C were thrown simultaneously into a 50.0 g of water at a temperature of 25.0 °C. What will be the final temperature of the system consisting of the two metal samples and the water? Assuming that this system is completely isolated from the surroundings. Use the information below for your calculations.

- $c_{s(\text{Al})} = 0.900 J / (K \cdot g)$
- $c_{s(\text{Cu})} = 0.385 J / (K \cdot g)$

- $c_{s(\text{H}_2\text{O})} = 4.184 \text{ J}/(\text{K} \cdot \text{g})$

Solution

Since our system is isolated, the thermal energy lost by one component is transferred to the other components.

$$q_1 = -q_2$$

which is equivalent to

$$C_1 \Delta T = -C_2 \Delta T$$

Noting that

$$C = mC_s$$

To avoid the complexity of handling three components, we can utilize the fact that temperature is a state function; we will simplify our calculations by choosing a different path to arrive to our final state.

We can do this in two steps:

Ignore the Aluminum sample and treat the copper and water as the only components of our system. After finding the equilibrium temperature, we add the Aluminum sample to the water and copper system, thereby reaching the same final state.

Which is the same answer we got before. This should make intuitive sense because regardless of the path we take, we end up with the exact same amount of thermal energy in our system.

Step 1

$$m_{\text{Cu}} \times C_{s(\text{Cu})} \times (T_f - T_{i(\text{Cu})}) = -m_{\text{H}_2\text{O}} C_{s(\text{H}_2\text{O})} (T_f - T_{i(\text{H}_2\text{O})})$$

$$30.0 \text{ g} \times 0.385 \text{ J}/(\text{K} \cdot \text{g}) \times (T_f + 20^\circ \text{C}) = -50.0 \text{ g} \times 4.184 \text{ J}/(\text{K} \cdot \text{g}) \times (T_f - 25.0^\circ \text{C})$$

Solving for T_f for the copper and water system gives us

$$T_f = 22.6^\circ \text{C}$$

Step 2

$$C_{(\text{Cu}+\text{H}_2\text{O})} \times (T_f - T_{i(\text{Cu}+\text{H}_2\text{O})}) = -m_{\text{Al}} \times C_{s(\text{Al})} \times (T_f - T_{i(\text{Al})})$$

Noting that

$$C_{(\text{Cu}+\text{H}_2\text{O})} = m_{\text{H}_2\text{O}} \times C_{s(\text{H}_2\text{O})} + m_{\text{Cu}} \times C_{s(\text{Cu})}$$

Combining the two equations and plugging the values gives us

$$[(30.0 \text{ g} \times 0.385 \text{ J}/(\text{K} \cdot \text{g})) + (50.0 \text{ g} \times 4.184 \text{ J}/(\text{K} \cdot \text{g}))] \times (T_f - 22.6^\circ \text{C}) = -10.0 \text{ g} \times 0.900 \text{ J}/(\text{K} \cdot \text{g}) \times (T_f - 60^\circ \text{C})$$

Solving for T_f :

$$T_f = 24.1^\circ \text{C}$$

Which is the final temperature of the whole system.

Alternative Approach

We can also choose a path where we add the Aluminum first, and then we add the copper. Again, temperature is a state function and choosing a different path will not affect the final answer. We also show the calculation for this path for the sake of completion.

Step 1

$$m_{\text{Cu}} \times C_{s(\text{Al})} \times (T_f - T_{i(\text{Al})}) = -m_{(\text{H}_2\text{O})} \times C_{s(\text{H}_2\text{O})} \times (T_f - T_{i(\text{H}_2\text{O})})$$

$$10.0 \text{ g} \times 0.900 \text{ J}/(\text{K} \cdot \text{g}) \times (T_f + 20^\circ \text{C}) = -50.0 \text{ g} \times 4.184 \text{ J}/(\text{K} \cdot \text{g}) \times (T_f - 25.0^\circ \text{C})$$

Solving for T_f for the aluminum and water system gives us

$$T_f = 26.4^\circ \text{C}$$

Step 2

$$[(m_{(\text{H}_2\text{O})} \times C_{s(\text{H}_2\text{O})}) + (m_{\text{Al}} \times C_{s(\text{Al})})] \times (T_f - T_{i(\text{Al}+\text{H}_2\text{O})}) = -m_{\text{Al}} \times C_{s(\text{Cu})} \times (T_f - T_{i(\text{Cu})})$$

Plugging the numbers,

$$[(10.0 \text{ g} \times 0.900 \text{ J}/(\text{K} \cdot \text{g})) + (50.0 \text{ g} \times 4.184 \text{ J}/(\text{K} \cdot \text{g}))] \times (T_f - 26.4^\circ \text{C}) = -30.0 \text{ g} \times 0.385 \text{ J}/(\text{K} \cdot \text{g}) \times (T_f + 20^\circ \text{C})$$

Solving for T_f

$$T_f = 24.1^\circ \text{C}$$

Q15A

Calculate the heat required to melt 3.00 g of ice and the heat required to change the temperature of water from 0°C to 100°C . What is the proportionality of the heat necessary to melt ice compared to the heat required to change the temperature of water from 0°C to 100°C ? Use values from Table S2 and assume $\Delta H_f = 334 \text{ J g}^{-1}$ for calculations? Why is the heat positive instead of negative?

Solution

First let's determine the amount of heat needed to melt ice:

$$q = m\Delta H_f$$

$$q = 1,002 \text{ J}$$

Next let's determine the heat required to raise the temperature of water 100°C :

$$q = mC_s\Delta T$$

$$q = 1,254 \text{ J}$$

So the proportionality was determined to be 4:5

The heat is positive because the heat is required. The heat is needed to make the states go from solid to liquid when melting and then liquid to gas when the temperature is raised from 0 to 100 degrees Celsius.

Q15B

An observation in the 18th century stated that the heat that raised a certain mass of water from its freezing point to boiling point is equal to four-thirds of the heat required to melt the same mass of ice. Using the theory behind the observation, estimate the heat required to melt 10 g of ice, know that the heat capacity of water is $4.18 \text{ J/}^\circ\text{C}$.

Solution

Let the heat required to raise the temperature of water from its freezing point to its boiling point is q_1 :

$$q = mc\Delta T \quad (12.E.10)$$

$$m = \text{mass} \quad (12.E.11)$$

$$c = 4.18 \frac{\text{J}}{\text{g}^\circ\text{C}} \quad (12.E.12)$$

$$\Delta T = \text{freezing point} - \text{boiling point} = 100 - 0 = 100^\circ\text{C} \quad (12.E.13)$$

$$q_1 = (10\text{g})(4.18 \frac{\text{J}}{\text{g}^\circ\text{C}})(100^\circ\text{C}) \quad (12.E.14)$$

$$q_1 = 4180 \text{ J} = 4.18 \text{ kJ} \quad (12.E.15)$$

Let the heat required to melt 10 g of ice is q_2 :

$$q_1 = \frac{4}{3}q_2 \quad (12.E.16)$$

$$q_2 = \frac{3}{4}q_1 \quad (12.E.17)$$

$$q_2 = \frac{3}{4}(4.18 \text{ kJ}) = 3.135 \text{ kJ} \quad (12.E.18)$$

Q17

For his birthday, his John's parents have given him a compressible oven filled with 0.250 mol argon. If he sets this compressible oven at 1.00 atm and 273 K and let it contract from a constant external pressure of 0.100 atm until the gas pressure reaches 10.00 atm and the temperature reaches 400 K, what is the work done on the gas, the internal energy change, and the heat absorbed by the gas?

Solution

Using the ideal gas law $PV = nRT$, we can see that:

$$V_o = \frac{nRT_o}{P_o}$$

$$V_f = \frac{nRT_f}{P_f}$$

$$n = 0.25 \text{ mol}, R = 0.082 \frac{\text{atm L}}{\text{mol K}} \quad P_o = 1 \text{ atm}, P_f = 10 \text{ atm} \quad T_o = 273 \text{ K}, T_f = 400 \text{ K}$$

where V_o and V_f represent the initial and final volume of the chamber. Now, the pressure clearly changes inside the chamber, but outside the chamber, the pressure is held constant, so from the perspective of the surroundings:

$$W = -P(V_f - V_o)$$

where P represents the constant external pressure. Now substituting in V_f and V_o we can see that we've solved for work. Now, for any thermodynamic process:

$$\Delta U = \frac{3}{2} nR(T_f - T_o)$$

So plugging in those values allows us to solve for ΔU . Now, for heat, we simply subtract W from ΔU , by the first law of thermodynamics.

First law of thermodynamics: the total energy of an isolated system is a constant, energy can be transformed from one form to another, but it can not be created or destroyed.

The solution is incomplete.

Q19

Take 4 moles of ideal, monatomic gas going through expansion processes. The gas was initially put at a pressure of 5.00 atm and a temperature of 30°C. The gas first goes through isothermal expansion until the volume doubled. An isochoric process follows as the pressure is halved. $C_v = \frac{3}{2}R$

- Build ΔU , W , and Q table for each process.
- Find the final temperature.

Solution

Isothermal Process:

$$\Delta U = 0 \quad (12.E.19)$$

$$W = Q \quad (12.E.20)$$

$$W = nRT \ln\left(\frac{V_f}{V_i}\right) = (4 \text{ mol})(8.314 \text{ J} \cdot \text{K}^{-1} \text{ mol}^{-1})(303.15 \text{ K}) \ln(2) = 6988 \text{ J} = Q \quad (12.E.21)$$

Isochoric Process

$$W = 0 \quad (12.E.22)$$

$$\Delta U = Q \quad (12.E.23)$$

First find the initial pressure of this process.

Use ideal law to relate pressure to volume. $PV = nRT$, in which P is inversely proportional to V .

Thus, when volume is doubled, pressure is halved so $P_f = 2.50 \text{ atm}$. Now, find the final temperature of this process.

Once again, relate P to T using ideal gas law. T is found to be directly proportional to P . Thus, if the pressure is halved from 2.50 atm, the temperature must also be halved from 303.15K.

$$\frac{1}{2}(303.15 \text{ K}) = 151.58 \text{ K}$$

$$Q = nC_v \Delta T = (4.00 \text{ mol})\left(\frac{3}{2}\right)\left(8.314 \frac{\text{J}}{\text{K mol}}\right)(-151.58 \text{ K}) = -7561 \text{ J} = \Delta U \quad (12.E.24)$$

Processes	ΔU	Q	W
Isothermal	0	6988J	6988J
Isochoric	-7561J	-7561J	0

Q20

An apparatus is set up such that an ideal gas is released into vacuum by opening a stopcock, hence allowing it to freely expand (i.e., no external force is applied). Calculate ΔU of the system, and prove that the free expansion process is adiabatic i.e no heat transfer.

Solution

$$\Delta U = \frac{3}{2} nR \Delta T \quad (12.E.25)$$

As the expansion is isothermal, i.e. it occurs without a change in temperature:

$$\Delta T = 0 \quad (12.E.26)$$

Hence,

$$\Delta U = 0 \quad (12.E.27)$$

and

$$\Delta U = q_p + w \quad (12.E.28)$$

As the external force applied on the gas is zero, the work done by the gas is zero.

Hence,

$$0 = q_p + 0 \quad (12.E.29)$$

$$q_p = 0 \quad (12.E.30)$$

This proves that the process is adiabatic.

Q21

A 150 L vessel contains 8.00 moles of neon at 270 K is compressed adiabatically, so that there is no gain nor loss of any heat, and irreversibly until the final temperature is 470 K. Calculate the change in internal energy, the heat added to the gas, and the work done on the gas.

Solution

Since Ne is a monatomic ideal gas and the volume of the vessel remains constant, the heat capacity of Ne can be expressed as:

$$C_p = \frac{3}{2} \times R = \frac{3}{2} \times 8.314 \frac{\text{J}}{\text{K} \cdot \text{mol}} = 12.47 \frac{\text{J}}{\text{K} \cdot \text{mol}} \quad (12.E.31)$$

$$\Delta U = nC_p \Delta T \quad (12.E.32)$$

$$\Delta U = 8.00 \times 12.47 \frac{\text{J}}{\text{K} \cdot \text{mol}} \times (470 - 270) \text{ K} = 19952 \text{ J} = 20.0 \text{ kJ} \quad (12.E.33)$$

Since the vessel is adiabatically compressed, no heat is added, therefore $q = 0$

$$\Delta U = q + w \quad (12.E.34)$$

$$20.0 \text{ kJ} = 0 + w \quad (12.E.35)$$

$$w = 20.0 \text{ kJ} \quad (12.E.36)$$

Q22

A gas expands at constant external pressure of 3.00 atm until its volume has increased by 9.00 to 15.00 L. During this process, it absorbs 800J of heat from the surroundings.

a. Calculate the energy change of the gas, ΔU

b. Calculate the work, w , done on the gas in an irreversible adiabatic ($q = 0$) process connecting the same initial and final state.

Solution

a)

$$\Delta U = q + w \quad (12.E.37)$$

$$\Delta U = 800 \text{ J} + -((3.00 \text{ atm})(15.00 \text{ L} - 9.00 \text{ L})(101.3 \frac{\text{J}}{\text{L atm}})) = -1023.4 \text{ J} \quad (12.E.38)$$

b)

$$q = 0 \quad (12.E.39)$$

$$\Delta U = w \quad (12.E.40)$$

$$\Delta U = -((3.00 \text{ atm})(15.00 \text{ L} - 9.00 \text{ L})(101.3 \frac{\text{J}}{\text{L atm}})) = -1823.4 \text{ J} \quad (12.E.41)$$

Q23

Using the theorem of equipartition of energy, calculate the specific heat capacity at constant pressure C_p at 25°C and 1 atm for O_2 and CO . Compare the calculated values to the experimental data ($\text{O}_2 = 29.36$, $\text{CO} = 29.1$) and thus calculate the percent of the experimental value that results from vibrational motions. Answer: The percent of C_p due to vibrational motions is 0.897% for O_2 and 0.141% for CO .

Solution

The equipartition theorem states that each degree of freedom in a molecule contributes to $\frac{1}{2} RT$ to the molar internal energy of a gas. To solve this problem, the number of degrees of freedom DOF in each molecule must be identified.

O_2 is a linear diatomic particle, thus it has 3 translational degrees of freedom and 2 rotational degrees of freedom. Since CO is a linear molecule it also has 3 translational and 2 rotational degrees of freedom. The molecules are at room temperature. Thus, it is assumed that there is no vibration in the bond (other than the zero point energy).

The low percentage of C_p due to vibrational modes in both molecules indicates that the vibrational motion is extremely small and can be neglected.

$$DOF(O_2) = f_t + f_r = 3 + 2 = 5$$

$$DOF(CO) = f_t + f_r = 3 + 2 = 5$$

Using the equation

$$U = (DOF) \cdot \left(\frac{1}{2} RT \right)$$

find the internal energy of each gaseous molecule.

$$U(O_2) = (5) \cdot \left(\frac{1}{2} RT \right)$$

Since O_2 and CO have the same degrees of freedom,

$$U(O_2) = U(CO) = \frac{5}{2} RT$$

The internal energy is related to C_v by

$$C_v = \left(\frac{\partial U}{\partial T} \right) = \left(\frac{\partial \frac{5}{2} RT}{\partial T} \right) = \frac{5}{2} R$$

C_v is related to C_p by

$$C_p = C_v + R = \frac{5}{2} R + R = \frac{7}{2} R = \frac{7}{2} \times 8.314 \frac{J}{mol K} = 29.099 \frac{J}{mol K}$$

The percent of the experimental value that results from vibration motion is the difference between calculated and experimental value as a percentage of the experimental value.

$$\% \text{vibrational motion}(O_2) = \frac{\text{experimental value} - \text{calculated value}}{\text{experimental value}} \times 100\% = \frac{29.36 - 29.099}{29.099} \times 100\% = 0.897\%$$

$$\% \text{vibrational motion}(CO) = \frac{\text{experimental value} - \text{calculated value}}{\text{experimental value}} \times 100\% = \frac{29.14 - 29.099}{29.099} \times 100\% = 0.141\%$$

Q23B

Using the classical equipartition theorem, calculate the value of C_p at 298 K and 1 atm for $HF(g)$ and $F_2(g)$, assuming that their pressure is constant. Then compare your calculation with the experiment values of $29.13 \frac{J}{K mol}$ and $31.30 \frac{J}{K mol}$ respectively. What is the per cent of the measured value that arises from vibrational motions?

Solution

Diatomic molecules possess a total of 6 degrees of freedom:

- 3 degrees in **translational** motion
- 2 degrees in **rotational** motion
- 1 degree in their **vibrational** motions

$$6.11 \text{ for } F_2(g) = 0.735 R \quad (12.E.42)$$

The differences are due to the fact that some quantized energy levels are not available at certain temperatures, in this case room temperature. This means that $HF(g)$ contributed only $0.004R$ of its experimental value of $29.13 \frac{J}{K mol}$ to the vibrational degree of freedom, which is 0.014%. $F_2(g)$ contributed $0.265R$ of its experimental value $31.30 \frac{J}{K mol}$, which is 0.847%.

Each translational and rotational degree of freedom contributes $\frac{R}{2}$ to the heat capacity of a gas, while each vibrational degree of freedom contributes $\frac{2R}{2} (=R)$. Therefore the predicted heat capacity (C_v) at a constant volume is:

$$C_v = 3 \times \left(\frac{R}{2}\right) + 2 \times \left(\frac{R}{2}\right) + 1 \times \left(\frac{2R}{2}\right) = 7\left(\frac{R}{2}\right) \quad (12.E.43)$$

Since the question is asking for the heat capacity of these gases that are under a constant pressure, the result is multiplied by an additional $\frac{2R}{2}$ because the volume can change in order to keep the pressure constant. Therefore:

$$C_v = \left(\frac{2R}{2}\right) + \left(\frac{7R}{2}\right) = \frac{9R}{2} \quad (12.E.44)$$

$$R = 8.3145 \frac{\text{J}}{\text{K mol}} \quad (12.E.45)$$

$$C_p = \frac{9(8.3145 \frac{\text{J}}{\text{K mol}})}{2} = 37.41 \frac{\text{J}}{\text{K mol}} \quad (12.E.46)$$

This is the calculated value for both of the two gases. Since the experiment values were $29.13 \frac{\text{J}}{\text{K mol}}$ for HF(g) and $31.30 \frac{\text{J}}{\text{K mol}}$ for F₂(g), the differences between the calculations were:

$$8.28 \text{ for HF(g)} = 0.996R \quad (12.E.47)$$

Q25

- Calculate the change of enthalpy when a 69-grams sample of zinc is heated from 753 K to 927 K at a constant pressure of 1 atm.
- Calculate the change of enthalpy when 3 moles of butane are heated from 204 K to 258 K at a constant pressure of 1 atm.

Solution

a)

$$q = mC_s\Delta T$$

$$m = 69 \text{ g}$$

$$\Delta T = 927 - 753 = 174 \text{ K}$$

$$C_{s(\text{zinc})} = 0.39 \frac{\text{J}}{\text{g}^\circ\text{C}}$$

$$q = 69 \times 0.39 \times 174 = 4682.34 \text{ J} = 4.56 \text{ kJ}$$

b)

$$q = nC_p\Delta T$$

$$n = 3$$

$$\Delta T = 258 - 204 = 54 \text{ K}$$

$$C_p = 132.42$$

$$q = 3 \times 132.42 \times 54 = 21452.04 \text{ J} = 21.4504 \text{ kJ}$$

Q31

What is the change in enthalpy when 8.19 grams of ethane (C₂H₆) vaporizes, assuming a normal boiling point and $\Delta H_{\text{vap}} = 14.72 \frac{\text{kJ}}{\text{mol}}$.

Solution

First, convert grams of ethane to moles:

$$(8.19 \text{ g C}_2\text{H}_6) \left(\frac{1 \text{ mol}}{30.07 \text{ g C}_2\text{H}_6} \right) = 0.272 \text{ mol}$$

This step is to match quantity units of ethane to those in the given value of ΔH_{vap} .

Then, given the ΔH_{vap} ,

$$(0.272 \text{ mol C}_2\text{H}_6) \left(\frac{14.72 \text{ kJ}}{1 \text{ mol C}_2\text{H}_6} \right) = 4.004 \text{ kJ}$$

$$\Delta H = 4.004 \text{ kJ}$$

This value is the change in enthalpy, or thermodynamic energy, when the specified amount of ethane undergoes the above reaction.

Q33

The heat capacity C_p of ice is $38 \frac{\text{J}}{\text{K mol}}$ and C_p of water is $75 \frac{\text{J}}{\text{K mol}}$. A 24.0 g ice cube at -15°C is placed into 120 g of water at room temperature (25°C). What is the temperature of the water when it reaches equilibrium?

Solution

$$q_{\text{ice}} = -q_{\text{water}} \quad (12.E.48)$$

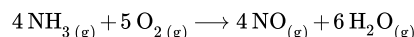
$$m_{\text{ice}} C_{\text{ice}} (T_f - T_{\text{ice}}) = -m_{\text{water}} C_{\text{water}} (T_f - T_{\text{water}}) \quad (12.E.49)$$

$$\left(\frac{24 \text{ g ice}}{18.02 \frac{\text{g}}{\text{mol}}} \right) \times \left(38 \frac{\text{J}}{\text{K mol}} \right) \times (T_f - (-15^\circ\text{C})) = - \left(\frac{120 \text{ g water}}{18.02 \frac{\text{g}}{\text{mol}}} \right) \times \left(75 \frac{\text{J}}{\text{K mol}} \right) \times (T_f - 25^\circ\text{C}) \quad (12.E.50)$$

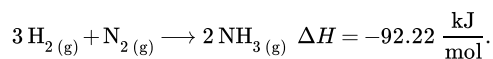
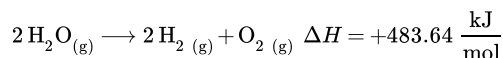
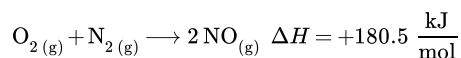
$$T_f = 21.316^\circ\text{C} \quad (12.E.51)$$

Q35

Determine the change in enthalpy (ΔH) for the following chemical reaction

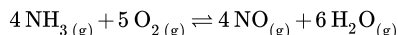


given that:

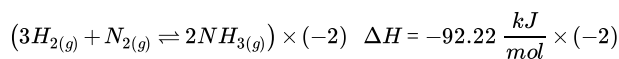
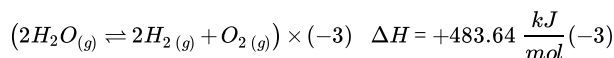
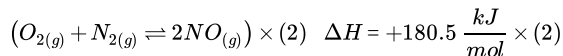


Solution

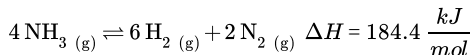
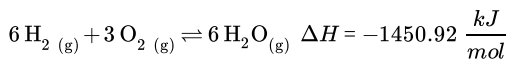
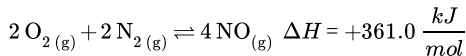
Using [Hess's Law](#), we can obtain the ΔH for



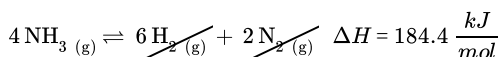
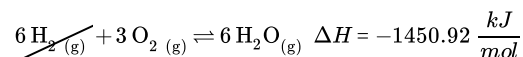
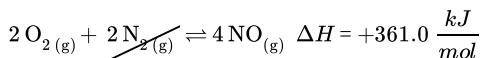
by adding and modifying the other equations provided until they resemble the equation above.



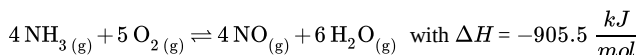
The above modifications simplify out to



Species that are common to both sides can be cancelled out.



Finally, after adding the remaining specimens and the enthalpy change that occurs with them we have:

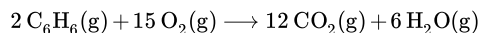


Q37

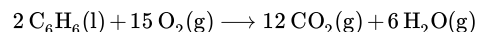
ΔH of $C_6H_6(g)$ from $C_6H_6(l)$ to its gaseous form is +33.9 kJ/mol. Which of these two substances absorbs off more heat when 10 lbs of the substance is burned? Which of these two substances require less energy when condensed to solid form?

Solution

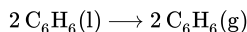
An alternative way to view this is to inspect the reactions under comparison



vs.



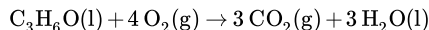
Recognizing that the difference between these equations is



The problem states that the ΔH of the conversion of $C_6H_6(g)$ to $C_6H_6(l)$ is positive, and therefore, is an endothermic process. Consequently, ten pounds of $C_6H_6(g)$ contains more enthalpy than 10 pounds of $C_6H_6(l)$. Since these will both produce carbon dioxide and water when burned, their products are of the same final energy level. As a result, $C_6H_6(g)$ will give off more energy than $C_6H_6(l)$ when burned and $C_6H_6(l)$ will require less energy to form a solid.

Q39

Calculate the enthalpy change under standard conditions for the reaction:



given information below:

- $\Delta H^\circ_f, O_2(g)$: 0 kJ/mol
- $\Delta H^\circ_f, CO_2(g)$: -393.5 kJ/mol
- $\Delta H^\circ_f, H_2O(l)$: -285.83 kJ/mol
- $\Delta H^\circ_f, C_3H_6O(l)$: -248 kJ/mol

Solution

Hess' Law:

$$\Delta H^\circ_{rxn} = \sum \Delta H^\circ_{f, products} - \sum \Delta H^\circ_{f, reactants}$$

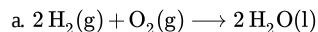
$$\Delta H^\circ_{rxn} = [3 (-393.5 \text{ kJ/mol}) + 3 (-285.83 \text{ kJ/mol})] - [(-248 \text{ kJ/mol}) + 4 (0 \text{ kJ/mol})] = -1790 \text{ kJ/mol}$$

Q47

A 10 gram sample of pure hydrogen is burned completely with excess oxygen to generate liquid water in a constant volume calorimeter at 25 °C. The amount of heat evolved is 1,420 kJ.

- Write and balance the chemical equation for the combustion reaction.
- Calculate the standard change in internal energy for the combustion of 1.00 mole hydrogen to liquid water.
- Calculate the standard enthalpy change per mole of hydrogen for the same reaction as in part (a).
- Calculate the standard enthalpy of formation per mole of hydrogen, using data for the standard enthalpies of formation of liquid water (Table T1).

Solution



b. $\Delta U^\circ = q_v = \frac{-1420 \text{ kJ}}{10 / (1.008 \times 2) \text{ g/mol}} = -286.3 \text{ kJ/(mol of } H_2)$

c. $\Delta H^\circ_{rxn} = \Delta U^\circ + \Delta n_{gas} RT = -286.3 \text{ kJ/(mol of } H_2) + (2 \text{ mol. of } H_2)(-3)(8.314 \text{ J/K/mol})(298 \text{ K}) = -580 \text{ kJ/mol}$

d. $\Delta H^\circ_{rxn} = 2 \Delta H^\circ_f \{H_2\}$ so $\Delta H^\circ_f / 2$.

A consequence of the constant-volume condition is that the heat released corresponds to q_v and thus to the internal energy change ΔU_{sys} rather than to ΔH_{sys} under constant pressure conditions.

$$\Delta U_{sys} = q_v$$

This comes from the definition of enthalpy

$$H_{sys} = U_{sys} + PV$$

and associated change

$$\Delta H_{sys} = \Delta U_{sys} + \Delta(PV)$$

or using the chain rule

$$\Delta H_{sys} = \Delta U_{sys} + P\Delta V + V\Delta P$$

which simplifies to

$$\Delta H_{sys} = \Delta U_{sys} + V\Delta P$$

If we assuming ideal gas law for the gases

$$PV = nRT$$

or

$$\Delta P = \frac{\Delta n RT}{V}$$

substituting this into the equation for enthalpy change gives the enthalpy change under *constant volume* conditions in terms of changing number of moles

$$\Delta H_{sys} = \Delta U_{sys} + \Delta n_g RT$$

or

$$\Delta H_{sys} = q_v + \Delta n_g RT$$

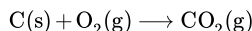
where Δn_g is the change in the number of moles of gases in the reaction.

Q49

Carbon dioxide (CO₂) is a common byproduct of the combustion of fossil fuels. Estimate the standard enthalpy of formation (ΔH_f°) of carbon dioxide at 25°C, use [Table T1](#).

Solution

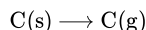
First write the equation:



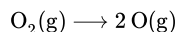
The bond enthalpies in [Table T3](#) allows us to calculate ΔH° because 2 C=O and 2 C-O bonds are formed:

$$\Delta H^\circ \approx 2 \text{ mol } (-192.0 \text{ kJ mol}^{-1}) + 2 \text{ mol } (-85.5) = -555 \text{ kJ}$$

Next we need to write the equations for "stripping" the atoms from there standard state to single atoms. Each of the atomization has an enthalpy that was found in [Table T1](#):



$$\Delta H^\circ = 1 \text{ mol } (716.7 \text{ kJ mol}^{-1}) = 716.7 \text{ kJ}$$



$$\Delta H^\circ = 2 \text{ mol } (249.2 \text{ kJ mol}^{-1}) = 498.4 \text{ kJ}$$

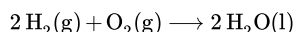
Combine the results of all three equations to calculate ΔH_f° of 1 mol of CO₂:

$$-555 \text{ kJ} + 716.7 \text{ kJ} + 498.4 \text{ kJ} = 660.1 \text{ kJ}$$

This solution should be the same if finding the standard enthalpy of formation by using [Table T1](#) which would be -393.5 kJ for one mole of CO₂. Since this question is not clear, it is better if the students answering this question to just use [Table T1](#) which will just use the standard heat of formation for and then have the sum of products minus the sum of reactants = the answer. So the answer would be -393.5 kJ.

Q51

Given the table of average bond enthalpies shown below, estimate the enthalpy change ΔH° for the following reaction:



Bond	Bond Enthalpy (kJ/mol)
H-H	436
H-O	463
O=O	498

Solution

To calculate the total enthalpy change of the reaction, we use Hess's Law. Thus, the total enthalpy of formation is equal to the sum of the enthalpy of formation of the reactants and the enthalpy of formation of the products.

$$\begin{aligned}\text{Step 1. } \Delta H_1 &= \text{the bond enthalpies of } 2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \\ &= 2(\text{H}-\text{H}) + (\text{O}=\text{O}) \\ &= 2(436) \text{ kJ/mol} + 498 \text{ kJ/mol} \\ &= +1370 \text{ kJ/mol}\end{aligned}$$

$$\begin{aligned}\text{Step 2. } \Delta H_2 &= \text{the bond enthalpies of } 2\text{H}_2\text{O}(\text{l}) \\ &= 2 \times 2(\text{H}-\text{O}) \\ &= 4(463) \text{ kJ/mol} \\ &= -1852 \text{ kJ/mol}\end{aligned}$$

$$\begin{aligned}\text{Step 3. } \Delta H^\circ &= \Delta H_1 + \Delta H_2 \\ &= 1370 \text{ kJ/mol} - 1852 \text{ kJ/mol} \\ &= -482 \text{ kJ/mol}\end{aligned}$$

*Note in Step 1, ΔH_1 is positive because the process of breaking the bonds into individual atoms is **always** an endothermic process while in Step 2, ΔH_2 is negative because the process of creating bonds out of the individual atoms is **always** an exothermic process.*

Q59

Two quantum states are separated by an energy of $0.6 \times 10^{-21} \text{ J}$. Calculate the relative populations of the two quantum states at a temperature of 33°C .

Solution

The equation for the relative populations of two quantum states, P_2/P_1 is given by the equation

$$\frac{P_2}{P_1} = e^{-(E_2 - E_1)/k_B T}$$

Where $E_2 - E_1$ is the energy difference between the two states, k_B is the Boltzmann constant, and T is the temperature.

Plug in the given values into the equation to reach the final answer.

$$\begin{aligned}\frac{P_2}{P_1} &= e^{-(0.6 \cdot 10^{-21} \text{ J}) / (1.38 \cdot 10^{-23} \text{ K}^{-1} (298 \text{ K}))} \\ \frac{P_2}{P_1} &= 0.86\end{aligned}$$

Q63

By utilizing the Harmonic Oscillator Model, calculate the relative population of the first energy state and the ground state, both at 278.15 K , for H_2 . The force constant for H_2 is $510 \frac{\text{N}}{\text{m}}$

Solution

A few pieces of information need to be understood before going about solving this equation. First, the question asks for the “relative population” of the first energy state at a given temperature (278.15 K) to the energy state at 0 K . Recall that energy states are basically a way of saying that every chemical species (molecule, atom, etc.) can “have” a specific (discrete) amount of energy. Each of these values are a “state”, and the lowest-value state can be considered the “ground state”, and all states above are “excited states”. Also recall that the states can be named by quantum number; i.e. if n were the quantum number, then the ground state would be $n = 0$ and the first state after that, an excited state, has a quantum number $n = 1$. The next step is to understand what “relative population is”. Basically, this refers to the probability that the species, in this case H_2 , will be found in one energy state versus another. The equation that describes the probability of a species being in energy level n is as follows:

$$P(n) = C e^{-\varepsilon_n / k_B T}$$

Where C is a constant, n is quantum number (energy level), k_b is Boltzmann’s constant, T is temperature, and ε is the energy of the molecule, which can be determined as such:

$$\begin{aligned}\varepsilon_n &= \left(n + \frac{1}{2}\right) h\nu \\ h\nu &= \frac{h}{2\pi} \sqrt{\frac{k}{\mu}}\end{aligned}$$

n , as already seen, is the quantum number. $h\nu$ in the equation for ε is defined as the product of planck's constant (h) divided by 2π and the $\sqrt{\frac{k}{\mu}}$, where k is the force constant, which is given in the problem, and the μ is the *reduced mass*, which can be calculated as follows:

$$\mu = \frac{m_1 m_2}{m_1 + m_2}$$

Where m is just the mass of each of the atoms. Since the mass of both hydrogens are 1 amu (u),

$$\mu = \frac{1 \times 1}{1 + 1} = \frac{1}{2}u$$

$$\frac{1}{2}u = 8.3027 \times 10^{-28} \text{ kg}$$

Because the question is only interested in *relative* population, the actual number of molecules there are doesn't really matter, since *relative* means to take the ratio between the probability of one energy state to another. Since the two energy states in comparison for this problem is $n = 1, n = 0$, (remember ground state is $n = 0$), the equation is as follows.

$$\frac{P(1) = C e^{-\varepsilon_1 / k_B 278.15 \text{ K}}}{P(0) = C e^{-\varepsilon_0 / k_B 278.15 \text{ K}}}$$

By simplification,

$$= e^{\frac{\varepsilon_1 - \varepsilon_0}{K_B T}}$$

And, substituting in the formula for energy(ε),

$$= e^{\frac{\left[\left(n + \frac{1}{2} \right) h\nu - \frac{1}{2} h\nu \right]}{K_B T}}$$

$$= e^{n h\nu / K_B T}$$

Here, $n = 1$ (because the 0 has already been taken into consideration) and $h\nu$ is as follows:

$$h\nu = \frac{h}{2\pi} \sqrt{\frac{k}{\mu}}$$

$$h\nu = \frac{6.62607 \times 10^{-34} \frac{\text{kgm}^2}{\text{s}^2}}{2\pi} \sqrt{\frac{510 \frac{\text{N}}{\text{m}}}{8.3027 \times 10^{-28} \text{ kg}}}$$

$$h\nu = \frac{6.62607 \times 10^{-34} \text{ Js}^{-1}}{2\pi} \sqrt{\frac{510 \frac{\text{N}}{\text{m}}}{8.3027 \times 10^{-28} \text{ kg}}}$$

$$h\nu = 8.2650 \times 10^{-20} \text{ J}$$

All the variables have been found, so all there is left is to substitute and solve.

$$= e^{\frac{-(1) (8.2650 \times 10^{-20} \text{ J})}{1.3806 \times 10^{-23} \frac{\text{m}^2 \text{ kg}}{\text{s}^2 \text{ K}} (278.15 \text{ K})}}$$

$$= e^{\frac{-(1) (8.2650 \times 10^{-20} \text{ J})}{1.3806 \times 10^{-23} \frac{\text{J}}{\text{K}} (278.15 \text{ K})}}$$

$$\approx e^{-21.5229}$$

$$\approx 4.495 \times 10^{-10}$$

Thus, the probability of the H_2 molecules being in the 1st energy state when compared to the vibrational state of the lowest energy is $\approx 4.495 \times 10^{-10}$. This extremely low probability makes sense, because intuition says that the compound will most likely be found at its lowest energy.

Abstract: Find difference in energy, ε , divide by Boltzmann's constant (K_B) and Temperature. Plug as e^x and solve.

Q69

A container holds 2 L of gas under 5.00 atm and a ball floating in 10 L of NaOH. As the volume of the gas expands to 20 L, the ball is turned upside down. This turn is caused by the temperature increase of the NaOH after the gas expands. Assuming that no heat is lost, the density of NaOH is 2.13

g/cm^3 and the specific heat is 4.184 J/g K , calculate the increase in temperature of the NaOH.

Solution

The "work" completed by the ball to turn equals the negative value of the work the gas absorbs.

$$w = -P(V_2 - V_1)$$

$$w = -(5 \text{ atm})(20 - 2) \text{ L}$$

$$w = 90 \text{ atm L } (101.325 \text{ J/1 atm L})$$

$$w = 9119.25 \text{ J}$$

The work carried out by the ball by the NaOH (in order to turn the ball) correlates to the increase in 10 L of NaOH temperature .

$$10 \text{ L of NaOH } (1000 \text{ mL/1 L}) = 10000 \text{ mL} = 10000 \text{ cm}^3 (2.13 \text{ g/cm}^3) = 21300 \text{ g of NaOH}$$

$$q = mC_p\Delta T$$

$$\Delta T = (q \text{ NaOH}) / (\text{Specific Heat NaOH} * \text{g NaOH})$$

$$\Delta T = (9119.25 \text{ J}) / (4.184 \text{ J/K g})(21300 \text{ g})$$

$$\Delta T = 0.102 \text{ K gained by the NaOH to turn the ball}$$

Q69

Some gas in a piston expands against a constant pressure of 1.2 atm from a volume of 3 L to 18 L. The piston turns an egg beater submerged in 150 g of water. If the water was originally at 25°C , what is its temperature once the gas stops expanding? Assume that all of heat goes into the water and the specific heat capacity of water is $4.184 \text{ J K}^{-1} \text{ g}^{-1}$.

Solution

From the first law of thermodynamics

$$\Delta U = q + w$$

In this case, $\Delta U = 0$, so

$$q = w = P_{ext} \Delta V = P_{ext} (V_f - V_i)$$

$$q = (1.2 \text{ atm})(18 \text{ L} - 3 \text{ L}) = 14.4 \text{ L atm}$$

Convert From L atm to joules using the conversion factor

$$14.4 \text{ L atm} \times 101.325 \text{ J L}^{-1} \text{ atm}^{-1} = 1459 \text{ J}$$

Now connect the heat transferred to the temperature increased via the specific heat c_{sp} via

$$q = mc_{sp}\Delta T$$

$$\Delta T = \frac{q}{mc_{sp}}$$

$$\Delta T = \frac{1459 \text{ J}}{(150 \text{ g})(4.184 \text{ J K}^{-1} \text{ g}^{-1})} = 2.325 \text{ K}$$

Since initial temperatures is $25^\circ = 298 \text{ K}$ so

$$T_f = 298 \text{ K} + 2.325 \text{ K}$$

$$T_f = 300.325 \text{ K} = 27.325^\circ \text{C}$$

12.E: Thermodynamic Processes (Exercises) is shared under a [CC BY-NC-SA 4.0](https://creativecommons.org/licenses/by-nc-sa/4.0/) license and was authored, remixed, and/or curated by LibreTexts.