

- 1. (5%) Explain why the value of S° for graphite is greater than that for diamond at 298K (use table2). Would this inequality hold at 0 K?**

A diamond is essentially a macroscopic molecule with no internal rotational or translational states. Graphite consists of large sheets of carbon that are bound to each other via intermolecular interactions. Consequently, it also has no significant internal rotational or translational states. But, since the sheets in graphite are free to slide past one another with relative ease (graphite makes a fine dry-lubricant), it does exhibit modes that are less like vibrations and more like translations, and we expect then that 1 mol of graphite will have greater entropy than diamond at a fixed temperature and pressure. In the absence of frozen-in residual entropy, the molar entropies of the two substances will both approach zero as the temperature is lowered.

- 2. (10%) One mole of an ideal gas at 298K expands isothermally from 1.0 to 2.0 L (a) reversibly and (b) against a constant external pressure of 12.2 bar.**

Calculate

ΔS_{sys} , ΔS_{surr} , and ΔS_{univ} in both cases. Are your results consistent with the nature of the process?

Sol.

(a) Use equation 8.10:

$$\Delta S_{\text{sys}} = nR \ln \frac{V_2}{V_1} = (1 \text{ mol})(8.314 \text{ J mol}^{-1} \text{ K}^{-1}) \ln \frac{2.0 \text{ L}}{1.0 \text{ L}} = 5.76 \text{ J K}^{-1}$$

Since the process is reversible, $\Delta S_{\text{univ}} = 0$ and

$$\Delta S_{\text{surr}} = -\Delta S_{\text{sys}} = -5.76 \text{ J K}^{-1}.$$

(b) Use the ideal gas equation to find the volume of the final state:

$$\begin{aligned} V_2 &= \frac{nRT}{P_2} \\ &= \frac{(1 \text{ mol})(0.08314 \text{ bar L mol}^{-1} \text{ K}^{-1})(298 \text{ K})}{12.2 \text{ bar}} \\ &= 2.03 \text{ L} \end{aligned}$$

Next, use equation 8.10 to find the entropy change of the system:

$$\Delta S_{\text{sys}} = nR \ln \frac{V_2}{V_1} = (1 \text{ mol})(8.314 \text{ J mol}^{-1} \text{ K}^{-1}) \ln \frac{2.03 \text{ L}}{1.0 \text{ L}} = \mathbf{5.89 \text{ K}^{-1}}$$

To find the entropy change in the surroundings, use $\Delta S_{\text{surr}} = - \frac{q_{\text{sys}}}{T}$:

$$\begin{aligned} \Delta S_{\text{surr}} &= - \frac{q_{\text{sys}}}{T} \\ &= - \frac{\Delta H_{\text{sys}}}{T} \text{ (constant P, expansion work only)} \\ &= - \frac{\Delta U_{\text{sys}} + P\Delta V}{T} \\ &= - \frac{P\Delta V}{T} \text{ } (\Delta U_{\text{sys}} \text{ is zero for an ideal gas at constant } T) \\ &= - \frac{(12.2 \text{ bar})(2.03 \text{ L} - 1.0 \text{ L})(100 \text{ J bar}^{-1})}{298 \text{ K}} \\ &= - 4.2 \text{ J K}^{-1} \end{aligned}$$

Overall,

$$\begin{aligned} \Delta S_{\text{univ}} &= 5.8 \text{ J K}^{-1} + (- 4.2 \text{ J K}^{-1}) \\ &= 1.6 \text{ J K}^{-1} \end{aligned}$$

- 3. (10%) One mole of an ideal gas is first heated at constant pressure from T to 3T and then it is cooled back to T at constant volume.**
- (a) Determine an expression for ΔS for the overall process.**
- (b) Show that the overall process is equivalent to an isothermal expansion of the gas at T from V to 3V, where V is the original volume.**
- (c) Show that the value of ΔS for the process in part (a) is the same as that for part (b).**

Sol.

(a) Step A: heating at constant pressure.

$$\Delta S_A = C_P \ln \frac{T_2}{T_1} = C_P \ln \frac{3T}{T} = C_P \ln 3$$

Step B: cooling at constant volume.

$$\Delta S_B = C_V \ln \frac{T_1}{T_2} = C_V \ln \frac{T}{3T} = C_V \ln \frac{1}{3} = - C_V \ln 3$$

Overall,

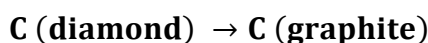
$$\Delta S = \Delta S_A + \Delta S_B = C_P \ln 3 + (-C_V \ln 3) = \left(\frac{5}{2}RT - \frac{3}{2}RT \right) \ln 3 = nR \ln 3$$

(b) At constant T ,

$$\begin{aligned}\Delta S &= nR \ln \frac{V_2}{V_1} \\ &= nR \ln \frac{3V}{V} \\ &= nR \ln 3\end{aligned}$$

(c) The entropy changes for part (a) and part(b) are the same.

4. (5%) Calculate ΔG° for the process (use table 2)



Is the formation of graphite from diamond favored at 25°C ? If so, why is it that diamonds do not become graphite on standing?

Sol.

The standard free energy change is given by:

$$\Delta G_{\text{rxn}}^\circ = \Delta G_{\text{f}}^\circ(\text{graphite}) - \Delta G_{\text{f}}^\circ(\text{diamond})$$

You can look up the standard free energy of formation values in Appendix 3 of the text.

$$\Delta G_{\text{rxn}} = (1)(0) - (1)(2.87 \text{ kJ mol}^{-1}) = -2.87 \text{ kJ mol}^{-1}$$

Thus, the formation of graphite from diamond is favored under standard-state conditions at 25°C . However, the rate of the diamond to graphite conversion is very slow (due to a high activation energy) so that it will take millions of years before the process is complete.

5. (5%) Water freezes spontaneously at -5°C and 1 atm, but ice has a lower entropy than liquid water. Explain how a spontaneous process can lead to a decrease in entropy.

Sol.

For a reaction to be spontaneous, ΔG must be negative. If ΔS is negative, as it is in this case, then the reaction must be exothermic (why?). When water freezes, it gives off heat (exothermic). Consequently, the entropy of the surroundings increases and $\Delta S_{\text{universe}} > 0$.

6. (5%) A heat engine operates between 210°C and 35°C . Calculate the

minimum amount of heat that must be withdrawn from the hot source to obtain 2000 J of work.

Sol.

According to the discussion on pp. 438-439, the heat transfers and temperatures are related by $\frac{q_c}{q_h} = \frac{T_c}{T_h}$, where the equality applies for reversible processes. Also,

$w_{\text{delivered}} = q_h - q_c$ or $q_c = q_h - w_{\text{delivered}}$. Substitute and solve:

$$\begin{aligned}\frac{q_h - 2000 \text{ J}}{q_h} &= \frac{(273 + 35) \text{ K}}{(273 + 210) \text{ K}} \\ q_h - 2000 \text{ J} &= 0.638 q_h \\ q_h (1 - 0.638) &= 2000 \text{ J} \\ q_h &= 5520 \text{ J}\end{aligned}$$

The minimum amount of energy withdrawn as heating is 5520 J.

- 7. (5%) Calculate the entropy change for the conversion of a 100.0 g sample of ice at -20°C to water at 37°C.**

Sol.

Assume that the process occurs at a constant pressure of 1 bar. Use the internet to find specific heat capacities of water ice and liquid water near 0°C. For example, $C_S(\text{ice}) = 2.050 \text{ J g}^{-1}\text{°C}^{-1}$ and $C_S(\text{liquid}) = 4.1813 \text{ J g}^{-1}\text{°C}^{-1}$. Assume these heat capacities are temperature-independent and be sure to include the entropy change for the phase change:

$$\begin{aligned}\Delta S &= mC_S(\text{ice}) \ln \frac{(273 + 0) \text{ K}}{(273 + 20) \text{ K}} + \frac{n \times \Delta H_{fus}}{273 \text{ K}} + mC_S(\text{liquid}) \ln \frac{(273 + 37) \text{ K}}{(273 + 0) \text{ K}} \\ &= (100.0 \text{ g})(2.050 \text{ J g}^{-1}\text{K}^{-1}) \ln \frac{(273 + 0)}{(273 - 20)} + \frac{(100 \text{ g})(334 \text{ J g}^{-1})}{273} \\ &\quad + (100.0 \text{ g})(4.1813 \text{ J g}^{-1}\text{K}^{-1}) \ln \frac{273 + 37}{273 + 0} \\ &= 191 \text{ J K}^{-1}\end{aligned}$$

- 8. (10%) Calculate the changes in entropy and Gibbs free energy for the following processes: (a) the mixing of 1 mole of nitrogen and 1 mole of oxygen, and (b) the mixing of 3 moles argon, 1 mole of helium, and 3 moles of hydrogen. Both parts (a) and (b) are carried out under conditions of constant**

temperature (298K) and constant pressure. Assume ideal behavior.

Sol.

Use equations 8.36 and 8.37

$$(a) x_{N_2} = \frac{1.00 \text{ mol}}{1.00 \text{ mol} + 1.00 \text{ mol}} = 0.500 \quad x_{O_2} = \frac{1.00 \text{ mol}}{1.00 \text{ mol} + 1.00 \text{ mol}} = 0.500$$

$$\begin{aligned}\Delta S_{\text{mix}} &= -R(n_{N_2} \ln x_{N_2} + n_{O_2} \ln x_{O_2}) \\ &= - (8.314 \text{ J mol}^{-1} \text{ K}^{-1})((1.00 \text{ mol})(\ln 0.5) + (1.00 \text{ mol})(\ln 0.5)) \\ &= 12 \text{ J mol}^{-1} \text{ K}^{-1}\end{aligned}$$

$$\begin{aligned}\Delta G_{\text{mix}} &= -T\Delta S_{\text{mix}} \\ &= - (298 \text{ K})(12 \text{ J mol}^{-1} \text{ K}^{-1}) \\ &= -3600 \text{ J mol}^{-1} = -3.6 \text{ kJ mol}^{-1}\end{aligned}$$

(b)

$$x_{\text{Ar}} = \frac{3.00 \text{ mol}}{3.00 \text{ mol} + 1.00 \text{ mol} + 3.00 \text{ mol}} = 0.429$$

$$x_{\text{He}} = \frac{1.00 \text{ mol}}{3.00 \text{ mol} + 1.00 \text{ mol} + 3.00 \text{ mol}} = 0.143$$

$$x_{\text{H}_2} = \frac{3.00 \text{ mol}}{3.00 \text{ mol} + 1.00 \text{ mol} + 3.00 \text{ mol}} = 0.429$$

$$\begin{aligned}\Delta S_{\text{mix}} &= -R(n_{\text{Ar}} \ln x_{\text{Ar}} + n_{\text{He}} \ln x_{\text{He}} + n_{\text{H}_2} \ln x_{\text{H}_2}) \\ &= - (8.314 \text{ J mol}^{-1} \text{ K}^{-1})((3.00 \text{ mol})(\ln 0.429) + (1.00 \text{ mol})(\ln 0.143) + (3.00 \text{ mol})(\ln 0.429)) \\ &= 58.4 \text{ J mol}^{-1} \text{ K}^{-1}\end{aligned}$$

$$\begin{aligned}\Delta G_{\text{mix}} &= -T\Delta S_{\text{mix}} \\ &= - (298 \text{ K})(58.4 \text{ J mol}^{-1} \text{ K}^{-1}) \\ &= -17,400 \text{ J mol}^{-1} = -17.4 \text{ kJ mol}^{-1}\end{aligned}$$

- 9. (10%) Which of the following is not accompanied by an increase in the entropy of the system: (a) mixing of two gases at the same temperature and pressure, (b) mixing of ethanol and water, (c) discharging a battery, and (d) expansion of a gas followed by compression to its original temperature, pressure, and volume?**

Sol.

(d) will not lead to an increase in entropy of the system. The gas is returned to

its original state. The entropy of the system does not change.

10. (10%) Suppose that your friend told you of the following extraordinary event.

A block of metal with a mass of 500 g was seen rising spontaneously from the table on which it was resting to a height of 1.00 cm above the table. She stated that the metal had absorbed thermal energy from the table that was then used to raise itself against the acceleration of gravity.

(a) Dose this process violate the first law of thermodynamics?

(b) Dose it violate the second law? Assume that the room temperature was 298 K and that the table was large enough that its temperature was unaffected by this transfer of energy. (Hint: calculate the decrease in entropy as a result of this process, and then estimate the probability for the occurrence of such a process. The acceleration due to gravity is 9.81 m s^{-2} .)

Sol.

(a) The mechanism for this miraculous-looking phenomenon is the following: a set of vibrating particles in the table oscillate upwards and in unison with sufficient energy to “bat” the block into the air, which for some unspecified and irrelevant reason remains stationary at the apex of its flight. Define the table as the surroundings and the block and surrounding air as the system.

Since there are no volume or pressure changes,

$$\Delta H_{\text{sys}} = \Delta U_{\text{sys}} = mgh = (0.500 \text{ kg})(9.81 \text{ m s}^{-2})(0.0100 \text{ m}) = 0.0491 \text{ J}$$

While the data given in the problem enables one to calculate the energy change of the system, there is insufficient information to compute the energy change of the surroundings. So, no determination regarding violation of the first law of thermodynamics can be made.

(b) To assess violation of the second law of thermodynamics, we must

determine whether $\Delta S_{\text{univ}} = \Delta S_{\text{sys}} + \Delta S_{\text{surr}}$ is greater than or less than zero

for the process. Assume that no energy is lost to acoustical excitation (sound) or to thermal excitation of the block or surrounding air. Under these assumptions, we can conclude that $\Delta S_{\text{sys}} = 0$ since the internal macrostate of the system is unchanged.

The surroundings (table) are assumed to be an infinite-capacity heat bath, in which case

$$\Delta S_{surr} = -\frac{\Delta H_{sys}}{T} = -\frac{0.0491J}{298K} = -1.65 \times 10^{-4} J K^{-1}$$

And

$$\Delta S_{univ} = 0 + (-1.65 \times 10^{-4} J K^{-1}) = -1.65 \times 10^{-4} J K^{-1}$$

Since the entropy change of the universe is negative, the process violates the second law of thermodynamics.

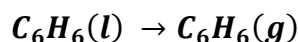
Use equation 8.3 in the text to find the relative proportion of microstates for the two macrostates:

$$\Delta S_{surr} = 1.65 \times 10^{-4} = k_B \ln \frac{W_f}{W_i}$$

$$\frac{W_f}{W_i} = e^{-(1.65 \times 10^{-4} J K^{-1})/k_B} = e^{-(1.65 \times 10^{-4} J K^{-1})/(1.38 \times 10^{-23} J K^{-1})} = e^{-1.2 \times 10^{19}}$$

The number of microstates in the final macrostate (hovering block) is negligible compared the number of microstates of the initial macrostate (block on table). If the system is in either state, you can bet money that it is the initial one!

11. (10%) The standard enthalpy of formation and the standard entropy gaseous benzene are $82.93 \text{ kJ mol}^{-1}$ and $269.2 \text{ J mol}^{-1} \text{ K}^{-1}$, respectively. Calculate ΔH° , ΔS° , and ΔG° for the following process at 25°C . Comment on your answers. (use table1)



Sol.

We use data in Appendix 2 of the text to calculate ΔH° and ΔS° .

$$\Delta H^\circ = \Delta H_{vap} = \Delta H_f^\circ[C_6H_6(g)] - \Delta H_f^\circ[C_6H_6(l)]$$

$$\Delta H^\circ = 82.93 \text{ kJ mol}^{-1} - 49.04 \text{ kJ mol}^{-1} = 33.89 \text{ kJ mol}^{-1}$$

$$\Delta S^\circ = S^\circ[C_6H_6(g)] - S^\circ[C_6H_6(l)]$$

$$\Delta S^\circ = 269.2 \text{ kJ mol}^{-1} - 172.8 \text{ kJ mol}^{-1} = 96.4 \text{ kJ mol}^{-1}$$

We can now calculate ΔG° at 298 K.

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

$$\Delta G^\circ = 33.89 \text{ kJ mol}^{-1} - (298 \text{ K})(96.4 \text{ J mol}^{-1} \text{ K}^{-1}) \times \frac{1 \text{ kJ}}{1000 \text{ J}}$$

$$\Delta G^\circ = 5.2 \text{ kJ mol}^{-1}$$

ΔH° is positive because this is an endothermic process. We also expect ΔS° to be positive because this is a liquid \rightarrow vapor phase change. ΔG° is positive because we are at a temperature that is below the boiling point of benzene (80.1°C).

12. (5%) Decomposition reactions are usually endothermic, whereas combination reactions are usually exothermic. Give a qualitative explanation for these trends.

Sol.

Energy must be supplied to break a chemical bond. By the same token, energy is released when a bond is formed.

13. (5%) Equal amounts of heat are added to two containers. One contains 1.0 mol of carbon dioxide and the other contains 1.0 mol of ammonia. If both are originally at the same temperature, which container (if any) will have the higher temperature at the end? Use molecular reasoning to justify your answer.

Sol.

For a given amount of thermal energy transferred to a pair of objects, the one with the larger heat capacity will experience the smaller change in temperature. According to the equipartition theorem, ammonia will have a larger heat capacity than carbon dioxide since it has one more rotational degree of freedom and two more vibrational degrees of freedom than carbon dioxide. The carbon dioxide will have the higher temperature.

14. (10%) For the reaction $\text{NH}_3(g) + \text{HCl}(g) \rightarrow \text{NH}_4\text{Cl}(s)$, calculate ΔG° at 25°C and 1000°C. At what temperature (if any) does ΔG° become zero? Assume that ΔH° and ΔS° are independent of temperature.

($\Delta H_f^\circ(\text{NH}_4\text{Cl}) = -314.4 \text{ kJ mol}^{-1}$, $\Delta H_f^\circ(\text{NH}_3) = -45.9 \text{ kJ mol}^{-1}$, $\Delta H_f^\circ(\text{HCl}) = -92.3 \text{ kJ mol}^{-1}$, $S^\circ(\text{NH}_4\text{Cl}) = 94.6 \text{ J mol}^{-1} \text{ K}^{-1}$, $S^\circ(\text{NH}_3) = 192.8 \text{ J mol}^{-1} \text{ K}^{-1}$, $S^\circ(\text{HCl}) = 187.0 \text{ J mol}^{-1} \text{ K}^{-1}$)

Sol.

Strategy.

First determine ΔH° and ΔS° using equation 7.41 and 8.26, respectively, and the data in Appendix 2. The ΔG° can be calculated at the two temperatures using equation 8.35.

Solution.

Plugging the appropriate ΔH_f° values from Appendix 2 into equation 7.41

$$\begin{aligned}\Delta H^\circ &= \Delta H_f^\circ(\text{NH}_4\text{Cl}) - [\Delta H_f^\circ(\text{NH}_3) + \Delta H_f^\circ(\text{HCl})] \\ &= -314.4 \text{ kJ mol}^{-1} - [-45.9 \text{ kJ mol}^{-1} + (-92.3 \text{ kJ mol}^{-1})] \\ &= -176.2 \text{ kJ mol}^{-1}\end{aligned}$$

Similarly, plugging the appropriate S° into equation 8.23

$$\begin{aligned}\Delta S^\circ &= S^\circ(\text{NH}_4\text{Cl}) - [S^\circ(\text{NH}_3) + S^\circ(\text{HCl})] \\ &= 94.6 \text{ J mol}^{-1} \text{ K}^{-1} - [192.8 \text{ J mol}^{-1} \text{ K}^{-1} + 187.0 \text{ J mol}^{-1} \text{ K}^{-1}] \\ &= -285.2 \text{ J K}^{-1} \text{ mol}^{-1}\end{aligned}$$

Equation 8.38 then gives the following ΔG° at 25°C and at 1000°C:

$$\begin{aligned}\Delta G^\circ(25^\circ\text{C}) &= -176.2 \text{ kJ mol}^{-1} - (298.15 \text{ K})(-285.2 \text{ J mol}^{-1} \text{ K}^{-1})(1 \text{ kJ}/1000 \text{ J}) \\ &= \mathbf{-91.32 \text{ kJ mol}^{-1}}\end{aligned}$$

$$\begin{aligned}\Delta G^\circ(1000^\circ\text{C}) &= -176.2 \text{ kJ mol}^{-1} - (1273.15 \text{ K})(-285.2 \text{ J mol}^{-1} \text{ K}^{-1})(1 \text{ kJ}/1000 \text{ J}) \\ &= \mathbf{186.9 \text{ kJ mol}^{-1}}\end{aligned}$$

The temperature at which ΔG° is zero can be found using the same method used for the decomposition of CaCO_3 :

$$T = \frac{\Delta H^\circ}{\Delta S^\circ} = \frac{(-176.2 \text{ kJ mol}^{-1})(1000 \text{ J kJ}^{-1})}{-285.2 \text{ J mol}^{-1} \text{ K}^{-1}} = 618 \text{ K or } 347^\circ\text{C}$$

15. Consider the equation



If 2.0 moles of N_2 react with 6.0 moles of H_2 to form NH_3 , calculate the work done (in joules) against a pressure of 1.0 bar at 25°C. What is ΔU° for this

reaction? Assume the reaction goes to completion.

Sol.

We initially have 8 moles of gas (2 of nitrogen and 6 of hydrogen). Since our product is 4 moles of ammonia, there is a net loss of 4 moles of gas (8 reactant → 4 product). The corresponding volume loss is

$$V = \frac{nRT}{P} = \frac{(4.0 \text{ mol})(0.08314 \text{ L bar mol}^{-1} \text{ K}^{-1})(298 \text{ K})}{1 \text{ bar}} = 99 \text{ L}$$

$$w = -P\Delta V = -(1 \text{ bar})(-99 \text{ L}) = 99 \text{ L bar} \cdot \frac{100 \text{ J}}{1 \text{ L bar}} = 9.9 \times 10^3 \text{ J} = 9.9 \text{ kJ}$$

$$\Delta H = \Delta U + P\Delta V \text{ or } \Delta U = \Delta H - P\Delta V$$

Using ΔH as $-183.6 \text{ kJ} = (2 \times -91.8 \text{ kJ})$, (because the question involves the formation of 4 moles of ammonia, not 2 moles of ammonia for which the standard enthalpy is given in the question), and $-P\Delta V$ as 9.9 kJ (for which we just solved):

$$\Delta U = -183.6 \text{ kJ} + 9.9 \text{ kJ} = \mathbf{-173.7 \text{ kJ}}$$

Table1

Substance	Formula	ΔH_f° (kJ mol ⁻¹)	ΔG_f° (kJ mol ⁻¹)	S° (J mol ⁻¹ K ⁻¹)	\bar{C}_p (J mol ⁻¹ K ⁻¹)
Ethanoic acid (acetic acid)(l)	CH ₃ COOH	-484.3	-389.9	159.8	123.3
Ethanal (acetaldehyde)(g)	CH ₃ CHO	-166.2	-133	263.8	55.3
2-Propanone (acetone)(l)	CH ₃ COCH ₃	-246.8	-153.55	198.7	126.3
Ethyne (acetylene)(g)	C ₂ H ₂	227.4	209.9	200.9	44
Benzene(l)	C ₆ H ₆	49.1	124.5	173.4	136
Butane(g)	C ₄ H ₁₀	-124.7	-15.7	310.0	140.9
Ethanol(l)	C ₂ H ₅ OH	-277.6	-174.8	160.7	112.3
Ethane(g)	C ₂ H ₆	-84	-32	229.2	52.5
Ethene (ethylene)(g)	C ₂ H ₄	52.4	68.4	219.3	42.9
Methanoic acid (formic acid)(l)	HCOOH	-425	-361.4	129.0	99
Glucose(s)	C ₆ H ₁₂ O ₆	-1274.5	-910.56	212.1	—
Methane(g)	CH ₄	-74.6	-50.8	186.3	35.7
Methanol(l)	CH ₃ OH	-239.2	-166.6	126.8	81.1
Propane(g)	C ₃ H ₈	-103.8	-23.4	270.3	73.6
Sucrose(s)	C ₁₂ H ₂₂ O ₁₁	-2221.7	-1544.3	360.2	—

Table2.

Inorganic Substances				
Substance	ΔH_f° (kJ mol ⁻¹)	ΔG_f° (kJ mol ⁻¹)	S° (J mol ⁻¹ K ⁻¹)	\bar{C}_P° (J mol ⁻¹ K ⁻¹)
Ag(s)	0	0	42.6	27.2
Ag ⁺ (aq)	105.6	77.1	72.7	21.8
AgCl(s)	-127.0	-109.7	96.1	50.8
AgBr(s)	-100.4	-96.9	107.1	52.4
AgI(s)	-61.8	-66.3	115.5	56.8
AgNO ₃ (s)	-124.4	-33.4	140.9	93.1
Al(s)	0	0	28.3	24.2
Al ³⁺ (aq)	-531.0	-485.0	-321.7	—
Al ₂ O ₃ (s)	-1675.7	-1582.3	50.9	79
As(s)	0	0	35.15	24.6
AsO ₄ ³⁻ (aq)	-888.1	-648.4	-162.8	—
AsH ₃ (g)	66.4	68.9	222.8	38.1
H ₂ AsO ₄ (s)	-906.3	—	—	—
Au(s)	0	0	47.4	25.4
Au ₂ O ₃ (s)	80.8	163.2	125.5	—
AuCl(s)	-34.7	—	—	—
AuCl ₃ (s)	-117.6	—	—	—
B(s)	0	0	5.9	11.1
B ₂ O ₃ (s)	-1273.5	-1194.3	54.0	62.8
H ₃ BO ₃ (s)	-1094.3	-968.9	90	86.1
H ₃ BO ₃ (aq)	-1067.8	-963.3	159.8	—
Ba(s)	0	0	62.5	28.1
Ba ²⁺ (aq)	-537.6	-560.8	9.6	—
BaO(s)	-558.2	-528.4	70.3	—
BaCl ₂ (s)	-855	-806.7	123.7	75.1
BaSO ₄ (s)	-1464.4	-1353.1	132.2	—
BaCO ₃ (s)	-1213	-1134.4	112.1	86
Be(s)	0	0	9.5	16.4
BeO(s)	-609.4	-580.1	13.8	25.6
Br ₂ (l)	0	0	152.3	75.7
Br ⁻ (aq)	-121.6	-104.0	82.4	-141.8
HBr(g)	-36.2	-53.2	198.70	29.1
C(graphite)	0	0	5.69	8.5
C(diamond)	1.90	2.87	2.4	6.1
CO(g)	-110.5	-137.3	197.9	29.1