

0. Graphite is thermodynamically less stable than diamond under standard conditions.
1. True
 2. False

1. Which statement would be the best interpretation of the First Law of Thermodynamics?
1. The total amount of energy in the universe is increasing.
 2. The total amount of entropy in the universe is increasing.
 3. The total amount of energy in the universe is constant.
 4. The total amount of matter in the universe is constant.

2. Enthalpy (H) is best defined by which of the following statements?
1. the capacity of a system to influence the entropy of its surroundings
 2. pressure-volume work
 3. a measure of a system's energetic degeneracy
 4. none of these describe enthalpy

- 3.1. Water changing from a solid to a liquid is
1. an endothermic change.
 2. an exothermic change.
 3. Neither an exothermic or an endothermic change

- 3.2. Heat flow is considered positive when heat flows (into, out of) a system; work is considered positive when work is done (by, on) a system.

1. out of; by
2. into; on
3. out of; on
4. into; by

4. Definition: state functions

An extensive state function's value is (dependent on/independent of) the amount of a given substance. An intensive state function's value is (dependent on/independent of) the amount of a given substance.

1. dependent on, dependent on
2. independent of, independent of
3. dependent on, independent of
4. independent of, dependent on

5. What is the enthalpy change for $\Delta H^\circ_f(\text{C(s, graphite)})$?

1. -1300 kJ/mol
2. 31 kJ/mol
3. 717 kJ/mol
4. 0 kJ/mol
5. -575 kJ/mol

6. Consider the following specific heats: copper, $0.384 \text{ J/g}\cdot^\circ\text{C}$; lead, $0.159 \text{ J/g}\cdot^\circ\text{C}$; water, $4.18 \text{ J/g}\cdot^\circ\text{C}$; glass, $0.502 \text{ J/g}\cdot^\circ\text{C}$. If the same amount of heat is added to identical masses of each of these substances, which substance attains the highest temperature? (Assume that they all have the same initial temperature.)

1. lead
2. water
3. glass
4. copper

- 7.1. You set up a bomb calorimetry experiment using 1 liter of water as your heat sink and

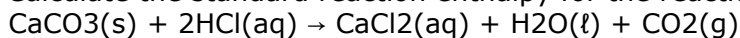
combusting a 4.409 g sample of propane (C₃H₈). If the initial and final temperature are 24.90 °C and 77.96 °C respectively, what is the approximate molar enthalpy of combustion of ethene? (Assume the calorimeter itself absorbs no heat. Assume the density of water is 1 g·mL⁻¹)

1. -222.0 kJ · mol⁻¹
2. -222, 000 kJ · mol⁻¹
3. -22.20 kJ · mol⁻¹
4. -22, 200 kJ · mol⁻¹
5. -2, 220 kJ · mol⁻¹

7.2. A 0.10 g piece of chocolate cake is combusted with oxygen in a bomb calorimeter. The temperature of 4,000 g of H₂O in the calorimeter is raised by 0.32 K. (The specific heat of the water is 1.0 cal/g·K and the heat of vaporization of water is 540 cal/g.) What is ΔE for the combustion of chocolate cake? Assume no heat is absorbed by the calorimeter.

1. -460 kcal/g
2. -532 kcal/g
3. -13.3 kcal/g
4. -12.8 kcal/g
5. -3900 kcal/g

8. Calculate the standard reaction enthalpy for the reaction of calcite with hydrochloric acid:



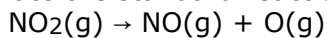
The standard enthalpies of formation are:

for CaCl₂(aq) : -877.1 kJ/mol;
for H₂O(l) : -285.83 kJ/mol;
for CO₂(g) : -393.51 kJ/mol;
for CaCO₃(s) : -1206.9 kJ/mol;
and for HCl(aq) : -167.16 kJ/mol.

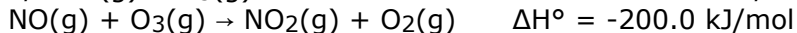
1. -72.7 kJ/mol
2. -165 kJ/mol
3. -38.2 kJ/mol
4. -98.8 kJ/mol
5. -15.2 kJ/mol
6. -116 kJ/mol
7. -215 kJ/mol

9. Calculation: Hess's Law and combined reaction enthalpies

Calculate the standard reaction enthalpy for the reaction:



Using:



1. +592 kJ/mol
2. +555 kJ/mol
3. +307 kJ/mol
4. +355 kJ/mol
5. +192 kJ/mol

10.1. How much internal energy would be associated with the rotational motion of 1 mole of CO₂?

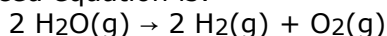
1. RT
2. 1/2 RT

3. $3/2 RT$
4. $2 RT$

10.2. What would be the total vibrational energy of .0833 moles of cetane ($C_{16}H_{34}$), assuming cetane is non-linear?

1. $72 RT$
2. $144 RT$
3. $12 RT$
4. $6 RT$

11. What energy change is associated with the reaction to obtain 1.00 mole of H_2 ? The balanced equation is:



and the relevant bond energies are:

H-H : 436 kJ/mol; H-O : 467 kJ/mol;
O-O : 146 kJ/mol; O=O : 498 kJ/mol.

1. +425 kJ
2. +249 kJ
3. -425 kJ
4. +498 kJ
5. -498 kJ
6. -436 kJ

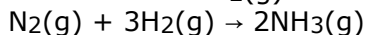
12.1. 1.95 mol of an ideal gas at 300 K and 3.00 atm expands from 16 L to 28 L and a final pressure of 1.20 atm in two steps:

(1) the gas is cooled at constant volume until its pressure has fallen to 1.20 atm, and
(2) it is heated and allowed to expand against a constant pressure of 1.20 atm until its volume reaches 28 L.

Which of the following is CORRECT?

1. $w = -4.57$ kJ for the overall process
2. $w = -6.03$ kJ for the overall process
3. $w = -4.57$ kJ for (1) and $w = -1.46$ kJ for (2)
4. $w = 0$ for the overall process
5. $w = 0$ for (1) and $w = -1.46$ kJ for (2)

12.2. 0.500 mole of $N_2(g)$ reacts with 1.50 moles $H_2(g)$ to produce $NH_3(g)$:



If this reaction is carried out in a system against a constant 0.75 atm pressure (i.e., a piston) at $0^\circ C$, calculate the magnitude of the P V work.

1. 2.27×10^3 J
2. 22.4 J
3. 22.7 J
4. 22.4×10^3 J
5. 4.54×10^3 J

13. Which of the following are not forms of internal energy?

1. motion of molecules
2. kinetic energy
3. heat
4. potential energy
5. chemical bonds

14. Which of the following statements concerning calorimetry is/are true?

- I) Bomb calorimeters hold the volume of the system constant.
II) The calorimeter itself does not absorb heat.
III) $\Delta H = q$ in a bomb calorimeter.

1. I, II
2. I only
3. II only
4. I, II, III
5. I, III
6. II, III
7. III only

15. A system did 150 kJ of work and its internal energy increased by 60 kJ. How much energy did the system gain or lose as heat?

1. The system gained 60 kJ of energy as heat.
2. The system gained 90 kJ of energy as heat.
3. The system gained 210 kJ of energy as heat.
4. The system lost 90 kJ of energy as heat.
5. The system lost 210 kJ of energy as heat.

16.1. For the four chemical reactions

- I) $3\text{O}_2(\text{g}) \rightarrow 2\text{O}_3(\text{g})$
II) $2\text{H}_2\text{O}(\text{g}) \rightarrow 2\text{H}_2(\text{g}) + \text{O}_2(\text{g})$
III) $\text{H}_2\text{O}(\text{g}) \rightarrow \text{H}_2\text{O}(\ell)$
IV) $2\text{H}_2\text{O}(\ell) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}_2(\ell)$

which one(s) is/are likely to exhibit a positive ΔS ?

1. All have a positive ΔS .
2. I, III and IV only
3. I and II only
4. III and IV only
5. II only

16.2. When Δn_{gas} for a reaction is large and positive ΔS for the reaction is likely to be (large/small) and (positive/negative).

1. large, positive
2. large, negative
3. small, positive
4. small, negative

17. Which of the following would have the largest absolute entropy?

1. $\text{N}_2(\text{g})$
2. $\text{N}_2(\text{s})$
3. $\text{N}_2(\text{aq})$ (nitrogen dissolved in water)
4. $\text{N}_2(\ell)$

18. Calculate the standard entropy of fusion of ethanol at its melting point 159 K. The standard molar enthalpy of fusion of ethanol at its melting point is $5.02 \text{ kJ} \cdot \text{mol}^{-1}$.

1. $-5.02 \text{ kJ} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}$
2. $-44.0 \text{ J} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}$
3. $-31.6 \text{ J} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}$
4. $+31.6 \text{ J} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}$
5. $+5.02 \text{ kJ} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}$

- 19.1. Which of the following statements is FALSE?
1. The total amount of energy and matter in the Universe is constant.
 2. Breaking chemical bonds is an endothermic process.
 3. It is more efficient to use a primary energy source than a secondary energy source.
 4. Entropy must be conserved in all chemical reactions.
- 19.2. Which of these statements about thermodynamics is NOT TRUE?
1. The entropy of the system is always increasing.
 2. Energy is conserved in chemical reactions.
 3. Heat given off to the surroundings is negative in sign.
 4. ΔV , ΔS , and ΔH are examples of changes in thermodynamic state functions.
 5. Work done on the system is positive in sign.
20. The law states that a substance that is perfectly crystalline at 0 K has an entropy of zero. This law is called
1. None of these
 2. the first law of thermodynamics.
 3. the third law of thermodynamics.
 4. the second law of thermodynamics.
 5. the zeroth law of thermodynamics.
21. In terms of absolute entropy, which of the following is/are true?
- I) W has both a real and a theoretical value.
 - II) The Boltzmann constant *can* be used when calculating S for molar quantities.
 - III) W for 1 mole of CO is greater than W for 1 mole of O₂.
1. III only
 2. I and III
 3. I, II and III
 4. II only
 5. I and II
 6. I only
 7. II and III
22. Which of the following systems would likely have the greatest residual entropy at absolute zero?
1. 10 BH₂F molecules
 2. 20 CCl₄ molecules
 3. 9 CHCl₃ molecules
 4. 100 BF₃ molecules
23. What is the positional entropy at absolute zero for two moles of CF₃Cl?
1. 11.5 J/K
 2. 5.7 J/K
 3. 0 J/K
 4. 8.3 J/K
 5. 23 J/K
- 24.1. What is the ΔS_{surr} when a balloon filled with 8 grams of hydrogen is combusted at 8 °C? The heat of combustion for a mole of H₂ is 242 kJ/mol.
1. 121000 J/K
 2. 861 J/K
 3. -861 J/K
 4. -121000 J/K
 5. -30.25 J/K

6. 3444 J/K
7. 30.25 J/K
8. -3444 J/K

24.2. A system releases 3456 J into its surroundings which are at 125 °C. By how much does this increase the entropy of the surroundings?

1. 138.24 J/K
2. 8.68 J/K
3. 11.60 J/K
4. this would decrease the entropy of the surroundings

25. For the vaporization of ethanol, $\Delta H_{\text{vap}} = 38.56 \text{ kJ/mol}$ and $\Delta S_{\text{vap}} = 109.7 \text{ J/mol} \cdot \text{K}$. What is the boiling point of ethanol?

1. 273.502 K
2. 0.352 K
3. 351.5 K
4. 2842 K
5. 2.84 K

26. Your roommate left 150 g of dry ice out on the counter last night and all of it sublimated. Given that $\Delta H = 393.5 \text{ kJ}\cdot\text{mol}^{-1}$ and $\Delta S = 2.023 \text{ kJ}\cdot\text{mol}^{-1}\cdot\text{K}^{-1}$ for dry ice sublimating and that the temperature in the room was 25 °C, by how much in total has your roommate increased the entropy of the universe?

1. $-0.198 \text{ kJ}\cdot\text{K}^{-1}$
2. $0 \text{ kJ}\cdot\text{K}^{-1}$
3. $0.105 \text{ kJ}\cdot\text{K}^{-1}$
4. $0.303 \text{ kJ}\cdot\text{K}^{-1}$
5. $2.40 \text{ kJ}\cdot\text{K}^{-1}$
6. $6.89 \text{ kJ}\cdot\text{K}^{-1}$
7. $-6.00 \text{ kJ}\cdot\text{K}^{-1}$

27. Which of the following statements is always true?

1. If the number of moles of gas does not change in a chemical reaction, then $\Delta S^\circ = 0$.
2. A reaction for which ΔS° is positive is spontaneous.
3. If ΔH° and ΔS° are both positive, then ΔG° will decrease when the temperature increases.
4. An exothermic reaction is spontaneous.

28.1. What happens to CuO(s) with respect to its elements when the temperature is raised?

	S°_{m}	$\Delta H^\circ_{\text{f}}$	$\Delta G^\circ_{\text{f}}$
	$\text{J}\cdot\text{mol}^{-1}\cdot\text{K}^{-1}$	$\text{kJ}\cdot\text{mol}^{-1}$	$\text{kJ}\cdot\text{mol}^{-1}$
Cu(s)	33.15		
O ₂ (g)	205.14		
CuO(s)	42.63	-157.3	-129.7

The compound is

1. more stable at higher temperatures.
2. less stable at higher temperatures.
3. Unable to determine

28.2. A reaction for which ΔH is negative and ΔS is negative

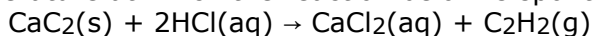
1. could become spontaneous at low temperatures.

2. is spontaneous at any temperature.
3. could become spontaneous at high temperatures.
4. is not spontaneous at any temperature.

29. The standard free energy of formation of $\text{CS}_2(\ell)$ is $65.3 \text{ kJ}\cdot\text{mol}^{-1}$. This means that at 298 K

1. $\text{CS}_2(\ell)$ will not spontaneously form $\text{C}(\text{s}) + 2 \text{S}(\text{s})$.
2. $\text{CS}_2(\ell)$ is thermodynamically stable.
3. No catalyst can be found to decompose $\text{CS}_2(\ell)$ into its elements.
4. $\text{CS}_2(\ell)$ is thermodynamically unstable.
5. $\text{CS}_2(\ell)$ has a negative entropy.

30.1. Assuming that ΔH°_r and ΔS°_r are independent of temperature, what is the cutoff temperature at which the reaction below is spontaneous?



	S°_m	ΔH°_f
	$\text{J}\cdot\text{mol}^{-1}\cdot\text{K}^{-1}$	$\text{kJ}\cdot\text{mol}^{-1}$
$\text{CaC}_2(\text{s})$	69.96	-59.8
$\text{HCl}(\text{aq})$	56.5	-167.16
$\text{CaCl}_2(\text{aq})$	59.8	-877.1
$\text{C}_2\text{H}_2(\text{g})$	200.94	226.73

1. 635 K
2. 390 K
3. 573 K
4. 745 K
5. Nonspontaneous at all temperatures
6. 416 K
7. Spontaneous at all temperatures