0. Graphite is thermodynamically less stable that 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 $\Delta H°_f C(s,graphite)$?
    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 J/g·°C; lead, 0.159 J/g·°C; water,4.18J/g·°C; glass, 0.502 J/g·°C. If the sameamount 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 (C_3H_8). 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 $g{\cdot}mL^{-1}$)

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 H_2O 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:

$CaCO3(s) + 2HCl(aq) \rightarrow CaCl2(aq) + H2O(\ell) + CO2(g)$

The standard enthalpies of formation are:
for CaCl2(aq) : −877.1 kJ/mol;
for $H_2O(\ell)$: −285.83 kJ/mol;
for $CO_2(g)$: −393.51 kJ/mol;
for $CaCO_3(s)$: −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:

$NO_2(g) \rightarrow NO(g) + O(g)$

Using:

$O_2(g) \rightarrow 2O(g)$ ΔH° = +498.4 kJ/mol
$3/2\ O_2(g) \rightarrow O_3(g)$ ΔH° = +142.7 kJ/mol
$NO(g) + O_3(g) \rightarrow NO_2(g) + O_2(g)$ ΔH° = -200.0 kJ/mol

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 CO_2?

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 H2? The balanced equation is:

$$2\ H_2O(g) \rightarrow 2\ H_2(g) + O_2(g)$$

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 $NH_3(g)$:

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

If this reaction is carried out in a system against a constant 0.75 atm pressure (i.e., a piston) at 0 °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 systemlost 210 kJ of energy as heat.

16.1. For the four chemical reactions

I) $3O_2(g) \rightarrow 2O_3(g)$
II) $2H_2O(g) \rightarrow 2H_2(g) + O_2(g)$
III) $H_2O(g) \rightarrow H_2O(\ell)$
IV) $2H_2O(\ell) + O_2(g) \rightarrow 2H_2O_2(\ell)$

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. $N_2(g)$
2. $N_2(s)$
3. $N_2(aq)$ (nitrogen dissolved in water)
4. $N_2(\ell)$

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 kJ · mol^{-1}.

1. -5.02 kJ · K^{-1} · mol^{-1}
2. -44.0 J · K^{-1} · mol^{-1}
3. -31.6 J · K^{-1} · mol^{-1}
4. $+31.6$ J · K^{-1} · mol^{-1}
5. $+5.02$ kJ · K^{-1} · 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 crystaline 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 O_2.
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 BH_2F molecules
2. 20 CCl_4 molecules
3. 9 $CHCl_3$ molecules
4. 100 BF_3 molecules

23. What is the positional entropy at absolute zero for two moles of CF_3Cl?
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 H_2 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, ΔH_{vap} = 38.56 kJ/mol and ΔS_{vap} = 109.7 J/mol · 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 ΔH = 393.5 $kJ \cdot mol^{-1}$ and ΔS = 2.023 $kJ \cdot mol^{-1} \cdot K^{-1}$ for dry ice sublimating and that the temperature in the room was 25 °C, by how much in total has has your roommate increased the entropy of the universe?
1. -0.198 $kJ \cdot K^{-1}$
2. 0 $kJ \cdot K^{-1}$
3. 0.105 $kJ \cdot K^{-1}$
4. 0.303 $kJ \cdot K^{-1}$
5. 2.40 $kJ \cdot K^{-1}$
6. 6.89 $kJ \cdot K^{-1}$
7. -6.00 $kJ \cdot K^{-1}$

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 ΔS° = 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?

	$S°_m$	$\Delta H°_f$	$\Delta G°_f$
	$J \cdot mol^{-1} \cdot K^{-1}$	$kJ \cdot mol^{-1}$	$kJ \cdot mol^{-1}$
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 $CS_2(\ell)$ is 65.3 $kJ \cdot mol^{-1}$. This means that at 298 K

1. $CS_2(\ell)$ will not spontaneously form C(s) + 2 S(s).
2. $CS_2(\ell)$ is thermodynamically stable.
3. No catalyst can be found to decompose $CS_2(\ell)$ into its elements.
4. $CS_2(\ell)$ is thermodynamically unstable.
5. $CS_2(\ell)$ has a negative entropy.

30.1. Assuming that $\Delta H°_r$ and $\Delta S°_r$ are independent of temperature, what is the cutoff temperature at which the reaction below is spontaneous?

$CaC_2(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + C_2H_2(g)$

	$S°_m$	$\Delta H°_f$
	$J \cdot mol^{-1} \cdot K^{-1}$	kJ·mol-1
$CaC_2(s)$	69.96	-59.8
HCl(aq)	56.5	-167.16
$CaCl_2(aq)$	59.8	-877.1
$C_2H_2(g)$	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