

Joe Perfect

Practice Test Chapter 19 Thermodynamics

Part I: Circle the letter of the best answer(s). No calculators are allowed on the multiple choice portion of the practice test!

Directions: Circle the best answer (or answers) for each question.

Target 1: I can define the term *spontaneous reaction* and apply it by identifying spontaneous processes.

- Which one of the following is always positive when a spontaneous process occurs?
a) ΔS_{system} b) $\Delta S_{\text{surroundings}}$ **c) $\Delta S_{\text{universe}}$** d) $\Delta H_{\text{universe}}$ e) $\Delta H_{\text{surroundings}}$
- Which one of the following cannot be determined when studying the thermodynamics of a reaction?
a) The speed of a reaction.
b) The direction of a reaction.
c) The extent of a reaction.
d) In which direction a reaction is spontaneous.
e) The temperature at which a reaction will be spontaneous.
- Which of the following is/are **false**? (Note: There could be more than one false statement. If there is, circle them all!)
a) A decrease in the entropy of a system occurs if $\Delta S_{\text{system}} > 0$.
b) The Gibbs free-energy difference between products and reactants for a reaction at equilibrium is > 0 . $\Delta G = 0$ at equilibrium
c) Reactions for which energy is required to initiate the process cannot be classified as spontaneous processes.
d) The standard free energy of formation of any pure substance at standard state conditions is zero. \rightarrow only elements are zero...
e) A reaction that is spontaneous in the forward reaction only is called an irreversible reaction (assuming the conditions of the reaction do not change.) \rightarrow not compds.
- Which of the following statements is true?
a) Processes that are spontaneous in one direction are spontaneous in the opposite direction.
b) Processes are spontaneous because they occur at an observable rate. FALSE... the rxn. may be so slow that you cannot see it occurring.
c) Spontaneity can depend on the temperature.
d) All of the statements are true.
 \rightarrow A reaction may occur at one temperature but not at another!

Target 2: I can define entropy and describe how it is related to randomness or disorder of a system.

5. Entropy . . .

- a) is defined as the energy of a system.
- b) is defined as the energy change of a system.
- ☒ c) is a state function.
- d) increases as a substance changes from a liquid to a solid.
- e) decreases as a substance sublimates.

→ solid to gas

6. Which molecule below should have the highest gas-phase entropy at 25°C?

- a) H₂
- ☒ b) C₂H₆
- c) C₂H₂
- d) CH₄
- e) C₂H₂

most atoms

7. Which of the following reaction(s) will have a positive ΔS° ? (There could be more than one answer. If there is, circle them all!)

- a) $2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{H}_2\text{O}(\text{g})$
- b) $2 \text{NO}_2(\text{g}) \rightarrow \text{N}_2\text{O}_4(\text{g})$
- c) $\text{H}^+(\text{aq}) + \text{F}^-(\text{aq}) \rightarrow \text{HF}(\text{aq})$
- d) $\text{Ba}^{2+}(\text{aq}) + 2 \text{F}^-(\text{aq}) \rightarrow \text{BaF}_2(\text{s})$
- ☒ e) $2 \text{HgO}(\text{s}) \rightarrow 2 \text{Hg}(\text{l}) + \text{O}_2(\text{g})$

Target 3: I can state the first and second law of thermodynamics.

8. Which 2 statements below best represent the First Law of Thermodynamics?

- ☒ a) Energy is conserved in any process. *at least any physical or chemical process...*
- b) The energy change of the universe is always increasing. *not true during a nuclear reaction*
- c) The rate of a reaction is dependent upon the energy change of a system.
- ☒ d) $\Delta E = q + w$
- e) $q = m \cdot C_p \cdot \Delta T$

9. Which 2 statements best represent the Second Law of Thermodynamics?

- ☒ a) In any spontaneous process, the entropy of the universe increases.
- b) All nonspontaneous processes have a $\Delta S > 0$.
- c) All spontaneous processes have a $\Delta S = 0$.
- ☒ d) For reversible processes: $\Delta S_{\text{universe}} = \Delta S_{\text{system}} + \Delta S_{\text{surroundings}} = 0$
- e) For reversible processes: $\Delta S_{\text{universe}} = \Delta S_{\text{system}} + \Delta S_{\text{surroundings}} < 0$

10. The entropy of the universe is _____.

- a) constant
- b) continually decreasing
- ☒ c) continually increasing
- d) zero
- e) the same as the energy, E

Target 4: I can calculate the ΔS° for any reaction from tabulated absolute entropy values, S° using . . . $\Delta S^\circ_{\text{rxn}} = \sum S^\circ_{\text{products}} - \sum S^\circ_{\text{reactants}}$

11. Calculate the standard entropy change, ΔS° , for the following reaction.



Substance	$\text{Al}_2\text{O}_3(\text{s})$	$\text{H}_2(\text{g})$	$\text{Al}(\text{s})$	$\text{H}_2\text{O}(\text{g})$	$\text{H}_2\text{O}(\text{l})$
$S^\circ, \text{J/mol}\cdot\text{K}$	51	131	28	189	70

- a) 179 J/K b) -151 J/K c) 151 J/K d) -225 J/K e) 225 J/K

$$\Delta S^\circ_{\text{rxn}} = [2(28) + 3(189)] - [(1)(51) + 3(131)] \approx (56 + 270) - (51 + 390) = 626 - 441 = 185$$

closest
to
choice "A"

Target 5: I can define free energy and the sign conventions associated with free energy. I can explain free energy in terms of enthalpy and entropy.

12. Gibbs free energy is . . . (Circle all that apply.)

- ☒ a) the energy associated with a chemical reaction that can be used to do work.
☒ b) a state function.
☐ c) the energy not used by a chemical reaction under standard conditions.
☐ d) never released by a chemical reaction as it is stored in the chemical bonds of the reactants and products.
☐ e) measured in $\text{kJ/g}\cdot^\circ\text{C}$.

13. If the sign of ΔG is _____, then the reaction _____. (Circle all that apply.)

- ☐ a) positive, is at equilibrium
☒ b) negative, is spontaneous in the forward direction
☒ c) positive, is nonspontaneous in the forward direction
☒ d) positive, has a bigger ΔH term than a $T\Delta S$ term in the equation
 $\Delta G = \Delta H - T\Delta S$
☐ e) negative, cannot occur

14. The sign of ΔG is always negative if . . .

- ☐ a) ΔH is positive and ΔS is positive.
☐ b) ΔH is positive and ΔS is negative.
☒ c) ΔH is negative and ΔS is positive.
☐ d) ΔH is negative and ΔS is negative.
☐ e) a reaction does not occur.

If both conditions (ΔH & ΔS) are favorable, reaction is spontaneous, no matter what the temp. !!

15. Chemical reactions will always occur if the reaction is EXO and the entropy ↑ during the reaction.
- endothermic, increases
 - endothermic, decreases
 - exothermic, decreases
 - ☒ exothermic, increases

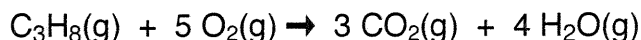
Target 6: I can calculate the free energy change at a standard temperature and pressure, G° , for any process from tabulated values for the standard free energies of reactants and products.

16. Which of the following are conventions used in establishing Standard Free Energies (G°)? Circle all that apply.

- ☒ The pressure of gases are 1 atmosphere.
- ☒ The concentrations of solutions are 1 molar.
- ☒ Elements in their standard states have a ΔG°_f of zero.
- d) The temperature is 25 Kelvin. (25°C)
- e) Solids are ionic.

↳ Solids need to be PURE... not necessarily ionic.

17. Calculate the standard entropy change, ΔG° , for the following reaction.



Substance	$\text{C}_3\text{H}_8(\text{g})$	$\text{CO}_2(\text{g})$	$\text{H}_2\text{O}(\text{g})$
$G^\circ_f, \text{kJ/mol}$	-23	-394	-229

- ☒ a) -2,075 kJ b) -1,510 kJ c) +1,510 kJ d) -225 kJ e) 0 kJ

$$\Delta G^\circ = 3(-394) + 4(-229) - (-23) \approx 3(-400) + 4(-225) + 25 = -1200 - 900 + 25$$

Target 7: I can estimate ΔG° at any temperature, given ΔS° and ΔH° .

18. Consider the following reaction: $\text{C}_2\text{H}_6(\text{g}) \rightarrow \text{C}_2\text{H}_4(\text{g}) + \text{H}_2(\text{g})$

The ΔH° is 137 kJ and the ΔS° is 120 J/K. This reaction will be _____.

- spontaneous at all temperatures.
- ☒ spontaneous only at high temperatures.
- spontaneous only at low temperatures.
- nonspontaneous at all temperatures.
- unreliable.

$-2100 + 25$
 ≈ -2075
 \uparrow
 choice
 "A"

- As a general rule, if ΔH is POSITIVE (not favorable) & ΔS is POSITIVE (favorable), then the reaction is only spontaneous at HIGH temp's. 😊 (This is analogous to ice melting.)

19. When KNO_3 dissolves in water at room temperature, ΔH is positive for the dissolution (dissolving) process. Given this information, what can you conclude?

- a) $\Delta G > 0$ for the dissolution process.
- b) $\Delta G = 0$ for the dissolution process.
- c) The dissolving of salts in water is always a spontaneous process.
- ☒ d) $\Delta S > 0$ for the dissolution process.
- e) $\Delta S < 0$ for the dissolution process.

Since $\Delta H = +$, the only way possible for process to occur is that $\Delta S > 0$!

Target 8: I can calculate the free energy change under nonstandard conditions, ΔG , given ΔG° .

20. Consider the following equation: $\Delta G = \Delta G^\circ + RT \ln Q$

Under what conditions is $\Delta G = \Delta G^\circ$?

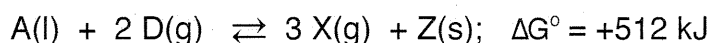
- a) When the temperature is 1 Kelvin.
- b) When the temperature is 1°C .
- ☒ c) When the conditions are standard conditions.
- d) Whenever $R = 8.31 \text{ J/mol}\cdot\text{K}$.
- e) When the reaction is spontaneous.

Target 9: I can calculate ΔG° from K and perform the reverse operation.

21. If ΔG° for a reaction is greater than zero, then _____.

- a) $K = 0$
- b) $K = 1$
- ☒ c) $K < 1$
- d) $K > 1$
- e) More info is needed.

22. Consider the following reaction at 25°C :



The equilibrium mixture _____.

- ☒ a) will consist almost exclusively of A and D
- b) will consist almost exclusively of A and Z
- c) will consist almost exclusively of X and Z
- d) will consist of significant amounts of A, D, X, and Z
- e) has a composition predictable only if one knows T and ΔH°

non spontaneous ... shifted way to left ... reactants favored!

23. Circle all TRUE statements relating ΔG° and the equilibrium constant, K .

- ☒ a) The more negative ΔG° , the larger the K .
- b) If ΔG° is positive, then $K > 1$.
- c) If a reaction is spontaneous, the ΔG° is positive and $K < 1$.
- ☒ d) If a reaction is spontaneous, the ΔG° is negative and $K > 1$.
- ☒ e) If $\Delta G^\circ = 0$, the $K = 1$.

Part 2: You may use a calculator! Show all of your work on separate sheets of paper.

1. Answer the following questions that relate to the chemistry of nitrogen.

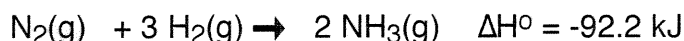
a) Two atoms of nitrogen combine to form a nitrogen molecule, as represented by the following equation:



Using the table of average bond energies below, determine the enthalpy change for the reaction.

bond	ave. bond energy (kJ mol ⁻¹)
-----	-----
N - N single bond	160
N = N double bond	420
N \equiv N triple bond	950

b) The reaction between nitrogen and hydrogen to form ammonia is below:



Predict the sign of the standard entropy change for the reaction. Justify your answer.

c) The value of ΔG° for the reaction represented in part (b) is negative at low temperatures but positive at high temperatures. Explain.

d) When $\text{N}_2\text{(g)}$ and $\text{H}_2\text{(g)}$ are placed in a sealed container at a low temperature, no measurable amount of $\text{NH}_3\text{(g)}$ is produced. Explain.



The reaction represented above is one that contributes significantly to the formation of photochemical smog.

a) Calculate the quantity of heat released when 73.1 g of NO(g) is converted to $\text{NO}_2\text{(g)}$.

b) For the reaction at 25°C, the value of the standard free energy change is -70.4 kJ.

i- Calculate the value of the equilibrium constant, K_{eq} , for the reaction at 25°C.

- ii- Indicate whether the value of ΔG° would become more negative, less negative, or remain unchanged as the temperature is increased. Justify your answer.

- c) Use the data in the table below to calculate the value of the standard molar entropy, S° , for $O_2(g)$ at $25^\circ C$.

Standard Molar Entropy ($J K^{-1} mol^{-1}$)	
NO(g)	210.8
NO ₂ (g)	240.1

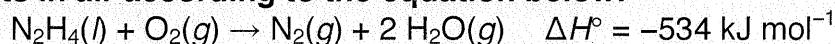
- d) Use the data in the table below to calculate the bond energy, in $kJ mol^{-1}$, of the nitrogen-oxygen bond in NO₂. Assume that the bonds in NO₂ molecule are equivalent (ie. have the same energy).

Bond Energy ($kJ mol^{-1}$)	
N-O bond in NO	607
O-O bond in O ₂	495
N-O bond in NO ₂	?

3. Hydrazine is an inorganic compound with the formula N₂H₄.

- (a) Draw the Lewis electron-dot diagram for the N₂H₄ molecule. Assume each nitrogen is bonded to 2 hydrogen atoms and to the other nitrogen atom.
- (b) On the basis of the diagram you completed in part (a), do all six atoms in the N₂H₄ molecule lie in the same plane? Explain.
- (c) The normal boiling point of N₂H₄ is $114^\circ C$, whereas the normal boiling point of C₂H₆ is $-89^\circ C$. Explain, in terms of the intermolecular forces present in each liquid, why the boiling point of N₂H₄ is so much higher than that of C₂H₆.
- (d) Write a balanced chemical equation for the reaction between N₂H₄ and H₂O that explains why a solution of hydrazine in water has a pH greater than 7.

N₂H₄ reacts in air according to the equation below:



- (e) Is the reaction an oxidation-reduction, acid-base, or decomposition reaction? Justify your answer.

- (f) Predict the sign of the entropy change, ΔS , for the reaction. Justify your prediction.
- (g) Indicate whether the statement written in the box below is true or false. Justify your answer.

The large negative ΔH° for the combustion of hydrazine results from the large release of energy that occurs when the strong bonds of the reactants are broken.

4. Answer the following questions about glucose, $C_6H_{12}O_6$, an important biochemical energy source.

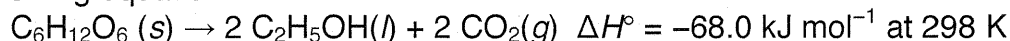
- (a) Write the empirical formula of glucose.

In many organisms, glucose is oxidized to carbon dioxide and water, as represented by the following equation.



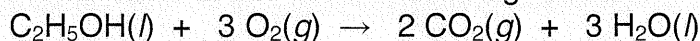
A 2.50 g sample of glucose and an excess of $O_2(g)$ were placed in a calorimeter. After the reaction was initiated and proceeded to completion, the total heat released by the reaction was calculated to be 39.0 kJ.

- (b) Calculate the value of ΔH° , in kJ mol^{-1} , for the combustion of glucose.
- (c) When oxygen is not available, glucose can be oxidized by fermentation. In that process, ethanol and carbon dioxide are produced, as represented by the following equation.



The value of the equilibrium constant, K_p , for the reaction at 298 K is 8.9×10^{39} .

- (i) Calculate the value of the standard free-energy change, ΔG° , for the reaction at 298 K. Include units with your answer.
 - (ii) Calculate the value of the standard entropy change, ΔS° , in $\text{J K}^{-1} \text{ mol}^{-1}$, for the reaction at 298 K.
 - (iii) Indicate whether the equilibrium constant for the fermentation reaction increases, decreases, or remains the same if the temperature is increased. Justify your answer.
- (d) Using your answer for part (b) and the information provided in part (c), calculate the value of ΔH° for the following reaction.



CHAPTER 19 PRACTICE TEST ANSWER KEY

- ① (a) -950 kJ ; You are forming $\text{N}\equiv\text{N}$ in the N_2 molecule. The bond energy given in the table is the energy required to BREAK a bond. Since we are MAKING a bond, the sign of bond energy is opposite... -950 kJ .
- (b) ΔS° is NEGATIVE! You are forming fewer gas particles than you have in the reactants... so it is becoming more orderly... which means $-\Delta S^\circ$.
- (c) Since the reaction is exothermic (as indicated by $-\Delta H$) and ΔS° is negative, the reaction will only be spontaneous at low temps. At low temp's, $T\Delta S^\circ$ is smaller than ΔH° $\therefore \Delta G^\circ$ is negative in the equation $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$
- (d) Even though at low temp's the reaction is spontaneous, the reaction must be so slow that you cannot observe it taking place. At low T 's, the number of colliding particles have activation energy is really small.

(2) (a) $\times \text{ kJ} = \left| \frac{73.1 \text{ g NO}}{30.0 \text{ g NO}} \right| \left| \frac{1 \text{ mol NO}}{2 \text{ mol NO}} \right| \left| \frac{114.1 \text{ kJ}}{1} \right| = 139 \text{ kJ}$

(b) i- $\Delta G^\circ = -RT \ln K$

$$-70.4 \text{ kJ} = (-0.008314 \frac{\text{kJ}}{\text{mol}\cdot\text{K}})(298 \text{ K}) \ln K$$

$$\ln K = 28.4$$

$$K = e^{28.4} = 2.2 \times 10^{12}$$

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

(neg) (neg)
(fav.) (not fav.)

Spontaneous at low T's.
As Temp ↑, the reaction becomes less spontaneous, ∴ ΔG becomes more positive or less negative!

→ Another way to look at it is using $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$.
If temp. increases, you are subtracting a more negative "TΔS" term, which makes ΔG more positive (less negative).

(c) $\Delta S^\circ_{\text{rxn}} = \sum S^\circ_{\text{prod}} - \sum S^\circ_{\text{react}}$

$$-146.5 \text{ J/K} = (2 \text{ mol})(240.1 \text{ J/mol}\cdot\text{K}) - [(2 \text{ mol})(205.1 \text{ J/mol}\cdot\text{K}) + S^\circ_{\text{O}_2}]$$

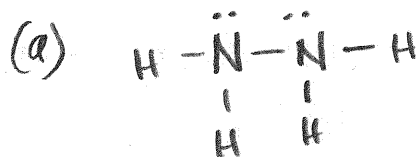
$$S^\circ_{\text{O}_2} = 205.1 \text{ J/mol}\cdot\text{K}$$

(d) $\Delta H^\circ_{\text{rxn}} = \text{Bond Energy (Broken)} - \text{Bond Energy (Formed)}$

$$-114.1 \text{ kJ} = \left[2(607 \frac{\text{kJ}}{\text{mol}}) + 1(495 \frac{\text{kJ}}{\text{mol}}) \right] - 4(\text{N-O})$$

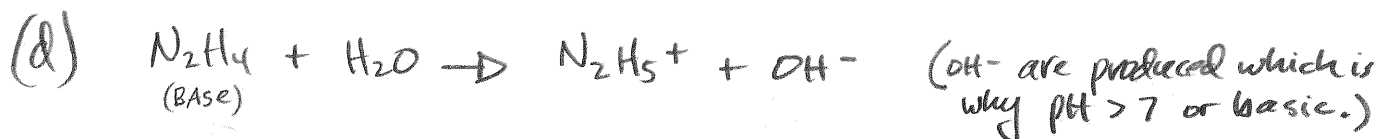
$$\text{N-O bond} = 456 \text{ kJ/mol}$$

③



(b) No. The geometry around each nitrogen is trigonal pyramidal. Therefore it is not possible to have all atoms in same plane.

(c) The N_2H_4 molecule is POLAR with a relatively large amount of intermolecular forces... Specifically London dispersion forces, dipole-dipole forces & hydrogen bonding. The C_2H_6 is NONPOLAR and only exhibits London dispersion forces. It takes more energy to overcome the IM forces in hydrazine, resulting in a higher boiling point.



(e) OXIDATION-REDUCTION REACTION. $\begin{cases} \text{N is oxidized } (-2 \rightarrow 0) \\ \text{O is reduced } (0 \rightarrow -2) \end{cases}$

(f) Sign of ΔS is POSITIVE. There are 3 moles of gas produced and only 1 mole of gas consumed. The net increase of 2 moles of gas results in an INCREASE of entropy.

(g) The statement is false for 3 reasons. \rightarrow Full CREDIT FOR GIVING EITHER ANSWER.

#1: Energy is not released when bonds are broken, but rather when bonds are FORMED!

#2: The bonds in the reactants are relatively weak compared to the bonds of the products.



$$(b) \quad x \frac{\text{kJ}}{\text{mol}} = \left| \frac{-39.0 \text{ kJ}}{2.50 \text{ g}} \right| \frac{180.16 \text{ g C}_6\text{H}_{12}\text{O}_6}{1 \text{ mol C}_6\text{H}_{12}\text{O}_6} = -2,810 \text{ kJ/mole}$$

$$(c) \quad \Delta G^\circ = -RT \ln K = -(8.31 \frac{\text{J}}{\text{mol} \cdot \text{K}})(298 \text{ K})(\ln 8.9 \times 10^{29})$$

$$\Delta G^\circ = -228,000 \text{ J/mol or } -228 \text{ kJ/mol}$$

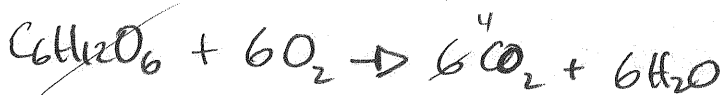
$$\text{ii-} \quad \Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

$$\Delta S^\circ = \frac{\Delta H^\circ - \Delta G^\circ}{T} = \frac{(-68.0 \text{ kJ/mol}) - (-228 \text{ kJ/mol})}{298 \text{ K}}$$

$$\Delta S^\circ = 0.537 \text{ kJ/K} \cdot \text{mol} = 537 \text{ J/K} \cdot \text{mol}$$

iii- Since ΔH is negative, as $T \uparrow$, the equilibrium is shifted left (according to LeChatelier's Principle). This means K DECREASES!

(d)

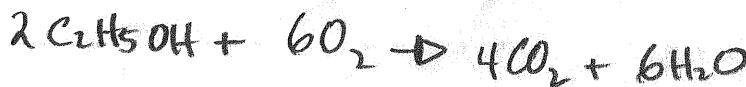


$$\Delta H^\circ = -2810 \frac{\text{kJ}}{\text{mol}}$$

flip equation
flip sign



$$\Delta H^\circ = 68 \text{ kJ/mol}$$



$$\Delta H^\circ = -2740 \frac{\text{kJ}}{\text{mole}}$$

cut equation in 1/2 (and ΔH !)



$$\Delta H^\circ = -1370 \frac{\text{kJ}}{\text{mole}}$$