

Unit 3 Practice Test: Thermo
AP Chemistry

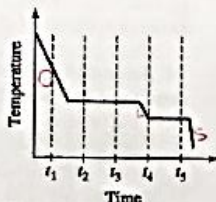
Name Key
Date _____ Block _____

MULTIPLE CHOICE – NO CALCULATOR ALLOWED

1. The cooling curve shows how the temperature of a sample varies with time as the sample goes through phase changes. The sample starts as a gas, and heat is removed at a constant rate. At which time does the sample contain the most liquid?

- A. t_1
B. t_2
C. t_3

D. t_4
E. t_5



2. A 100 g sample of a metal was heated to 100°C and then quickly transferred to an insulated container holding 100 g of water at 22°C. The temperature of the water rose to reach a final temperature of 35°C. Which of the following can be concluded?

- A. The metal temperature changed more than the water temperature did; therefore the metal lost more thermal energy than the water gained.
B. The metal temperature changed more than the water temperature did, but the metal lost the same amount of thermal energy as the water gained.
C. The metal temperature changed more than the water temperature did; therefore the heat capacity of the metal must be greater than the heat capacity of the water.
D. The final temperature is less than the average starting temperature of the metal and the water; therefore the total energy of the metal and water decreased.

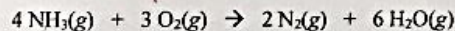
3. In an insulated cup of negligible heat capacity, 50. g of water at 40.°C is mixed with 30. g of water at 20.°C. The final temperature of the mixture is closest to

- A. 27°C
B. 30°C
C. 33°C
D. 38°C

$$50 \times 4.18 \times (T_f - 40) = 30 \times 4.18 \times (T_f - 20)$$

$$50 T_f - 2000 = 30 T_f + 600$$

$$80 T_f = 2600$$



Substance	ΔH_f°
$\text{NH}_3(\text{g})$	-46 kJ/mol
$\text{H}_2\text{O}(\text{g})$	-242 kJ/mol

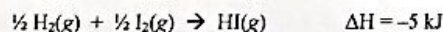
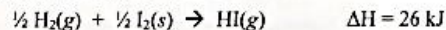
4. Based on the values for standard molar enthalpies of formation shown in the table, what is the value of ΔH° for the reaction represented above?

- A. -1400 kJ
B. -1270 kJ
C. -200 kJ
D. +1270 kJ

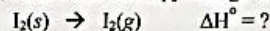
$$\Delta H_{\text{rxn}} = (2 \times 0 + 6 \times -242) - (4 \times -46 + 3 \times 0)$$

$$= -1452 + 184$$

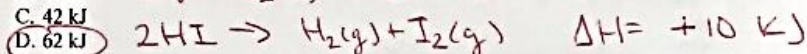
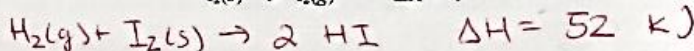
Use the information below to answer the following two questions.



5. Based on the information above, what is the enthalpy change for the sublimation of iodine?

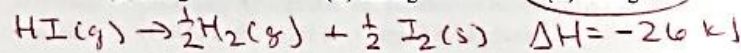


- A. 21 kJ
B. 31 kJ
C. 42 kJ
D. 62 kJ



6. What is the approximate mass of $\text{HI}(\text{g})$ that must decompose into $\text{H}_2(\text{g})$ and $\text{I}_2(\text{s})$ in order to release 500. kJ of energy?

- A. 250 g
B. 650 g
C. 1300 g
D. 2500 g



$$-500 \text{ kJ} \times \frac{1 \text{ mol HI}}{-26 \text{ kJ}} \times \frac{128 \text{ g HI}}{1 \text{ mol HI}} \approx 500 \times 5$$

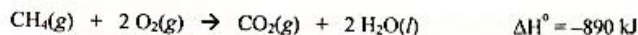
The following two questions refer to an experiment to determine the value of the heat of fusion of ice. A student used a calorimeter consisting of a polystyrene cup and a thermometer. The cup was weighed, then filled halfway with warm water, then weighed again. The temperature of the water was measured, and some ice cubes were added to the cup. The mixture was gently stirred as the ice melted, and the lowest temperature reached by the water in the cup was recorded. The cup and its contents was weighed again.

7. The purpose of weighing the cup and its contents again at the end of the experiment was to
A. determine the mass of ice that was added
B. determine the mass of water that evaporated
C. verify the mass of water that was cooled
D. verify the mass of the calorimeter cup

8. Suppose that during the experiment, a significant amount of water adhered to the ice cubes that were added to the cup. How does this affect the calculated value for the heat of fusion of ice?

- A. The calculated value is too large because less warm water had to be cooled.
B. The calculated value is too large because more cold water had to be heated.
C. The calculated value is too small because less ice was added than the student had assumed.
D. There is no effect on the calculated value because the water that adhered to the ice cubes was at 0°C.

242
6
1452



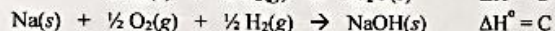
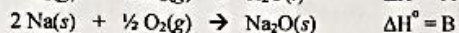
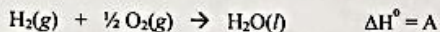
ΔH_f° for $\text{CO}_2(\text{g}) = X \text{ kJ/mol}$ and ΔH_f° for $\text{H}_2\text{O}(\text{l}) = Y \text{ kJ/mol}$

9. The value of ΔH° for the combustion of methane gas is equal to -890 kJ as shown above. The standard enthalpy of formation of $\text{CO}_2(\text{g})$ is equal to X kilojoules per mole. The standard enthalpy of formation of $\text{H}_2\text{O}(\text{l})$ is equal to Y kilojoules per mole. In terms of X and Y , the standard enthalpy of formation of methane gas is equal to

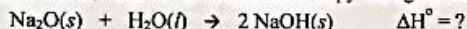
- A. $X + 2Y + 890$
 B. $X + 2Y - 890$
 C. $X + Y + 890$
 D. $X + Y - 890$

$$-890 = (X + 2Y) - (H_f(\text{CH}_4) + 2 \times 0)$$

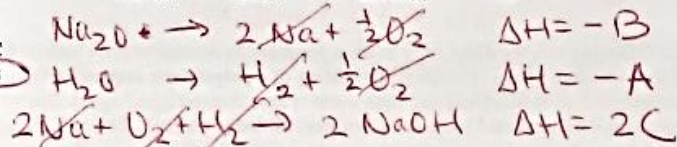
$$-890 = X + 2Y - H_f(\text{CH}_4)$$



10. Based on the information above, what is the standard enthalpy change for the following reaction:



- A. $A + B - C$
 B. $A + B - 2C$
 C. $2C - A - B$
 D. $C - A - B$

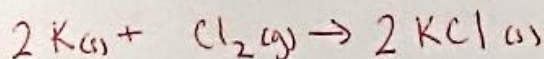


11. Which of the following reactions is not thermodynamically favored at low temperatures but becomes favored as the temperature increases?

Reaction	ΔH° (kJ/mol _{rxn})	ΔS° (J/(mol _{rxn} ·K))
(A) $2 \text{CO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{CO}_2(\text{g})$	-566	-173
(B) $2 \text{H}_2\text{O}(\text{g}) \rightarrow 2 \text{H}_2(\text{g}) + \text{O}_2(\text{g})$	484	90.0
(C) $2 \text{N}_2\text{O}(\text{g}) \rightarrow 2 \text{N}_2(\text{g}) + \text{O}_2(\text{g})$	-164	149
(D) $\text{PbCl}_2(\text{s}) \rightarrow \text{Pb}^{2+}(\text{aq}) + 2 \text{Cl}^-(\text{aq})$	23.4	-12.5

12. It is observed that the reaction producing KCl from its elements goes essentially to completion. Which of the following is a true statement about the thermodynamic favorability of the reaction?

- A. The reaction is favorable and driven by an enthalpy change only.
 B. The reaction is favorable and driven by an entropy change only.
 C. The reaction is favorable and driven by both enthalpy and entropy changes.
 D. The reaction is unfavorable due to both enthalpy and entropy changes.



13. Which of the following equations represents a reaction for which the standard entropy change is positive ($\Delta S^\circ > 0$)?

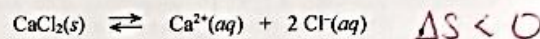
- A. $3 \text{O}_2(\text{g}) \rightarrow 2 \text{O}_3(\text{g})$
 B. $2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{H}_2\text{O}(\text{l})$
 C. $\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$
 D. $\text{I}_2(\text{g}) + 2 \text{K}(\text{s}) \rightarrow 2 \text{KI}(\text{s})$

14. An endothermic chemical reaction is thermodynamically favorable under standard conditions at 298 K. Therefore the value of ΔS° for the reaction must be

- A. equal to zero
 B. equal to $\Delta H^\circ/298 \text{ K}$
 C. greater than $\Delta H^\circ/298 \text{ K}$
 D. less than $\Delta H^\circ/298 \text{ K}$

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

- + +



15. For the process of dissolving solid calcium chloride in water, the entropy change might be expected to be positive. However, ΔS for the process is actually negative. Which of the following best helps to account for this decrease in entropy?

- (A) The particles in solid calcium chloride are more ordered than the aqueous ions in solution.
 (B) Water molecules in the hydration shells of Ca^{2+} and Cl^- are more ordered than they are in pure water.
 (C) The average distance between Ca^{2+} and Cl^- ions in solution is greater than the average distance between the ions in the crystal lattice.
 (D) The value of ΔH for the dissolution of CaCl_2 is positive, which drives the reaction forward.

16. A certain reaction is thermodynamically favored at temperatures below 400. K, but it is not favored at temperatures above 400. K. The value of ΔH° for the reaction is -20 kJ/mol . What is the value of ΔS° for the reaction? (Assume that ΔH° and ΔS° do not change with temperature.)

- (A) $-50 \text{ J mol}^{-1} \text{ K}^{-1}$
 (B) $-20 \text{ J mol}^{-1} \text{ K}^{-1}$
 (C) $-0.050 \text{ J mol}^{-1} \text{ K}^{-1}$
 (D) $50 \text{ J mol}^{-1} \text{ K}^{-1}$

$$\Delta G = \Delta H - T \Delta S = 0 @ 400$$

$$-20 - 400 \Delta S = 0$$

$$\text{Pb}(\text{s}) \rightleftharpoons \text{Pb}(\text{l}) \quad \Delta H = -20.5 \text{ kJ}$$

$$= -20.5 \text{ kJ} = -\frac{20.5}{400} = \Delta S$$

17. Which of the following is true for the process represented above at 327°C and 1 atm? The normal melting point of Pb is 327°C .

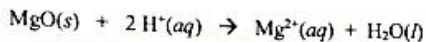
- A. $\Delta H = 0$
 B. $T \Delta S = 0$
 C. $\Delta H = T \Delta S$
 D. $\Delta S < 0$

@ equilibrium
 $\therefore \Delta G = 0$

$$\Delta G = \Delta H - T \Delta S = 0$$

$$\Delta H = T \Delta S$$

FREE RESPONSE - CALCULATOR IS ALLOWED



1. A student was assigned the task of determining the enthalpy change for the reaction between solid MgO and aqueous HCl represented by the net ionic equation above. The student uses a polystyrene cup calorimeter and performs four trials. Data for each trial are shown in the table below.

Trial	Volume of 1.0 M HCl (mL)	Mass of MgO(s) Added (g)	Initial Temperature of Solution (°C)	Final Temperature of Solution (°C)
1	100.0	0.25	25.5	26.5
2	100.0	0.50	25.0	29.1
3	100.0	0.25	26.0	28.1
4	100.0	0.50	24.1	28.1

a) Which is the limiting reactant in all four trials, HCl or MgO? Justify your answer.

The temperature change depended on the amount of MgO added, indicating it was the limiting reactant

b) The data in one of the trials is inconsistent with the data in the other three trials. Identify the trial with inconsistent data and draw a line through the data from that trial in the table above. Explain how you identified the inconsistent data.

Trial 1

The temperature change should be directly proportional to the amount of MgO added.

In trial 1, the temp change is $\frac{1}{2}$ of what was obtained in trial 3

For parts (c) and (d), use the data from one of other three trials (i.e., not from the trial you identified in part (b) above.) Assume the calorimeter has a negligible heat capacity and that the specific heat of the contents of the calorimeter is $4.18 \text{ J g}^{-1} \text{ }^\circ\text{C}^{-1}$. Assume that the density of the HCl(aq) is 1.0 g/mL .

c) Calculate the magnitude of q , the thermal energy change, when the MgO was added to 1.0 M HCl(aq). Express your answer in units of kilojoules.

ex: trial 4

$$q = (100 \text{ mL} \times \frac{1 \text{ g}}{1 \text{ mL}} + 0.50 \text{ g}) (4.18) (4.0)$$

$$= 1680 \text{ J} = \boxed{1.680 \text{ kJ}}$$

d) Determine the student's experimental value of ΔH° for the reaction between MgO and HCl in units of $\text{kJ/mol}_{\text{rxn}}$.

$$\Delta H = \frac{-1.680 \text{ kJ}}{0.50 \text{ g} \times \frac{1 \text{ mol MgO}}{40.31 \text{ g}}} = \boxed{-135 \text{ kJ/mol}}$$

e) Enthalpies of formation for substances involved in the reaction are shown in the table below. Using the information in the table, determine the accepted value of ΔH° for the reaction between MgO(s) and HCl(aq).

Substance	ΔH_f° (kJ/mol)
MgO(s)	-602
H ₂ O(l)	-286
H ⁺ (aq)	0
Mg ²⁺ (aq)	-467

$$\Delta H = (-467 + -286) - (-602 + 2 \times 0)$$

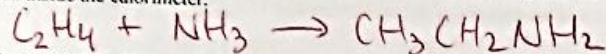
$$\Delta H = -753 + 602$$

$$\Delta H = \boxed{-151 \text{ kJ}}$$

f) The accepted value (in part (e)) and the experimental value (in part (d)) do not agree. If the calorimeter leaked heat energy to the environment, would it help account for the discrepancy between these two values? Explain.

Yes, if heat leaked out of the calorimeter, the measured ΔT would be less, and the calculated value for $q \neq \Delta H$ would be lower than expected.

2. Equimolar amounts of ethylene (C_2H_4) and ammonia (NH_3) were placed in a calorimeter. They reacted together to produce ethylamine, $CH_3CH_2NH_2$. The reaction proceeded to completion inside the calorimeter.



(a) Use the data in the table at right to calculate the value of ΔS° for this reaction. Include units in your answer.

$$\Delta S = 285 - (219 + 193)$$

$$\Delta S = -127 \text{ J/K}$$

Substance	S° ($J \text{ mol}^{-1} \text{ K}^{-1}$)
$CH_3CH_2NH_2(g)$	285
$C_2H_4(g)$	219
$NH_3(g)$	193

(b) ΔH° for the reaction is -49 kJ/mol . Use the data in the table at right to calculate the value of the Standard Enthalpy of Formation of ethylamine. Include units in your answer.

$$-49 = H_f - (52.30 - 46.19)$$

$$-49 = H_f - 6.11$$

$$H_f = -42.89 \text{ kJ/mol}$$

Substance	H° (kJ mol^{-1})
$CH_3CH_2NH_2(g)$?
$C_2H_4(g)$	52.30
$NH_3(g)$	-46.19

(c) Determine the value of ΔG° for this reaction at 25°C . Include units in your answer.

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

$$= -49 - 298(-0.127)$$

$$= -11.2 \text{ kJ/mol}$$

(d) This reaction is thermodynamically favorable at 25°C . Is this reaction driven by enthalpy only, driven by entropy only, or driven by both enthalpy and entropy? Justify your answer.

enthalpy only. Since entropy is decreasing, the fact that the reaction is exothermic ($\Delta H < 0$) helps make $\Delta G < 0$.

(e) Did the temperature inside the calorimeter increase, decrease, or remain the same as ethylamine was produced in this reaction? Justify your answer.

Temperature should increase since the reaction is exothermic.

(f) Would this reaction be thermodynamically favored to occur if it were carried out at 450 K ? Justify your answer with a calculation and an explanation.

$$\Delta G = \Delta H - T\Delta S = -49 - 450(-0.127)$$

$$= 8.15 \text{ kJ/mol}$$

No. ΔG would be positive at this temperature.

Reaction	Equation	ΔH° (kJ)	ΔS° (J/K)	ΔG° (kJ)
X	$C(s) + H_2O(g) \rightleftharpoons CO(g) + H_2(g)$	+131	+134	+91
Y	$CO_2(g) + H_2(g) \rightleftharpoons CO(g) + H_2O(g)$	+41	+42	+29
Z	$2 CO(g) \rightleftharpoons C(s) + CO_2(g)$?	?	?

3. Answer the following questions using the information related to reactions X, Y, and Z in the table above.

(a) For reaction Y at 298 K , which quantity is larger:

the total bond energy of the reactants or the total bond energy of the products

Explain.
The reaction is endothermic, so more energy is required to break the bonds in the reactants than the energy released when bonds are formed.

(b) Calculate the minimum temperature that would have to be reached in order for reaction Y to become thermodynamically favorable. Assume that the values of ΔH° and ΔS° do not change with temperature.

$$\Delta H = T\Delta S$$

$$T = 976 \text{ K}$$

$$41 = T(42/1000)$$

(c) Calculate the value of each of the following quantities for reaction Z at 298 K .

(i) ΔH°

$$= -131 - 41 = -172 \text{ kJ/mol} \quad \rightarrow \text{flip X + Y}$$

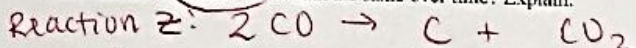
(ii) ΔS°

$$= -42 - 134 = -176 \text{ J/mol K}$$

(iii) ΔG°

$$= -91 - 29 = -120 \text{ kJ/mol}$$

(d) A sealed glass reaction vessel contains only $CO(g)$ and a small amount of $C(s)$. If a reaction occurs and the temperature is held constant at 298 K , will the pressure in the reaction vessel increase, decrease, or remain the same over time? Explain.



The decomposition of CO into $C + CO_2$ is spontaneous at 298 K , therefore the number of moles of gas in the flask would decrease, decreasing the pressure.