

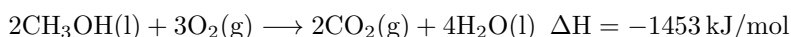
Non Sibi High School

Andover's Chem 550/580: Advanced Chemistry

Chapter 13, Review Quiz 1 Answers

1

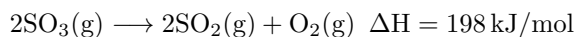
If 125 kilograms of methanol is burned according to the combustion equation below, how much heat will be released?



$$125 \text{ kg} \left(\frac{1000 \text{ g}}{1 \text{ kg}} \right) \left(\frac{1 \text{ mol CH}_3\text{OH}}{32.04 \text{ g}} \right) \left(\frac{-1453 \text{ kJ}}{2 \text{ mol CH}_3\text{OH}} \right) = -2.83 \times 10^6 \text{ kJ}$$

2

If 3.55 kJ of heat are absorbed during the decomposition reaction below, how many milliliters of sulfur trioxide gas, measured at 22°C and 712 mmHg, will decompose?



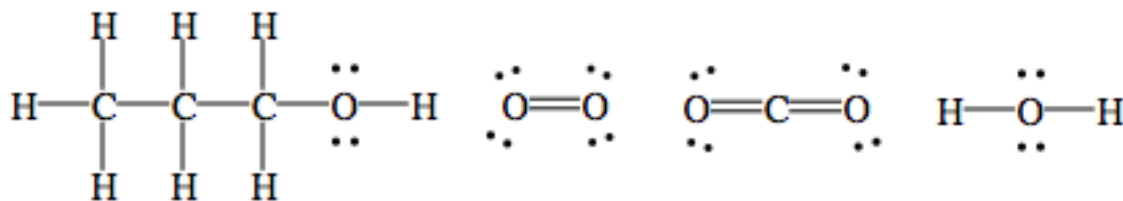
$$3.55 \text{ kJ} \left(\frac{2 \text{ mol SO}_3}{198 \text{ kJ}} \right) = 0.03586 \text{ mol SO}_3$$

$$\frac{0.03586 \text{ mol} \times 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \times (22 + 273) \text{ K}}{\frac{712}{760} \text{ atm}} = 0.927 \text{ L} \left(\frac{1000 \text{ mL}}{1 \text{ L}} \right) = 927 \text{ mL SO}_3$$

3

For the reaction $2\text{C}_3\text{H}_7\text{OH}(g) + 9\text{O}_2(g) \longrightarrow 6\text{CO}_2(g) + 8\text{H}_2\text{O}(g)$, estimate ΔH using average bond energies.

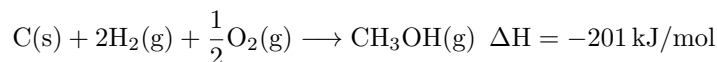
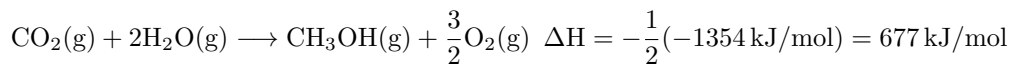
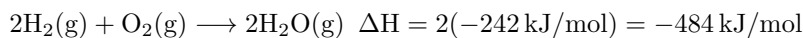
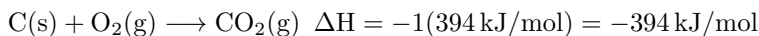
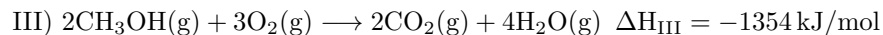
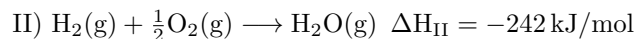
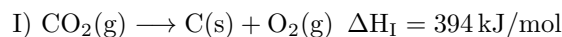
Lewis structures:



$$\begin{aligned}
 \Delta H \text{ (estimated)} &= \\
 &(2 \times 7)(\text{C} - \text{H}) + (2 \times 2)(\text{C} - \text{C}) + (2 \times 1)(\text{C} - \text{O}) + (2 \times 1)(\text{O} - \text{H}) + (9 \times 1)(\text{O} = \text{O}) \\
 &\quad - (6 \times 2)(\text{C} = \text{O}) - (8 \times 2)(\text{O} - \text{H}) \\
 &= 14(414) + 4(347) + 2(360) + 2(464) + 9(498) - 12(745) - 16(464) \\
 &= -3050 \text{ kJ/mol}
 \end{aligned}$$

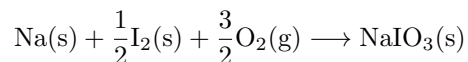
4

Calculate ΔH for the reaction $\text{C(s)} + 2\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \longrightarrow \text{CH}_3\text{OH}(\text{g})$ using the following three reactions:



5

Write the balanced formation reaction, including physical states, for solid sodium iodate, NaIO_3 .



| Compound | ΔH_f° (kJ/mol) |
|---------------------|-----------------------------|
| NO(g) | 90. |
| NO ₂ (g) | 33 |

6

Calculate ΔH° for the reaction $2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \longrightarrow 2\text{NO}_2(\text{g})$ using the following information:

$$\Delta H^\circ = 2(33) - 2(90) - 1(0) = -114 \text{ kJ/mol}$$

7

The specific heat of magnesium metal is $1.05 \text{ J/g}\cdot^\circ\text{C}$. How much heat in kilojoules is lost when a 225 gram sample of magnesium metal is cooled from 625°C to 125°C ?

$$q = 225 \text{ g} \left(\frac{1.05 \text{ J}}{\text{g}\cdot^\circ\text{C}} \right) (125 - 625)^\circ\text{C} \left(\frac{1 \text{ kJ}}{1000 \text{ J}} \right) = -118 \text{ kJ}$$

8

In an insulated calorimeter, a 475 gram piece of tin metal originally at 132°C was added to 135 grams of water originally at 19°C . The final temperature of the tin-water mixture was 36°C . Determine the specific heat of tin.

$$\begin{aligned} q_{\text{Sn lost}} &= -q_{\text{water gained}} \\ 475 \text{ g} (s_{\text{Sn}})(36 - 132)^\circ\text{C} &= -135 \text{ g} \left(\frac{4.18 \text{ J}}{\text{g}\cdot^\circ\text{C}} \right) (36 - 19)^\circ\text{C} \\ s_{\text{Sn}} &= 0.21 \text{ J/g}\cdot^\circ\text{C} \end{aligned}$$

9

The specific heat of tungsten metal is $0.13 \text{ J/g}\cdot^\circ\text{C}$. In an insulated calorimeter, a 955 gram piece of tungsten metal originally at 375°C was added to 725 grams of water originally at 18°C . Determine the final temperature of the tungsten-water mixture.

$$\begin{aligned} q_{\text{W lost}} &= -q_{\text{water gained}} \\ 955 \text{ g} \left(\frac{0.13 \text{ J}}{\text{g}\cdot^\circ\text{C}} \right) (t_{\text{final}} - 375^\circ\text{C}) &= -725 \text{ g} \left(\frac{4.18 \text{ J}}{\text{g}\cdot^\circ\text{C}} \right) (t_{\text{final}} - 18^\circ\text{C}) \\ t_{\text{final}} &= 32^\circ\text{C} \end{aligned}$$

10

In an insulated calorimeter, 18.2 grams of solid cesium hydroxide at 22.3°C was dissolved in 135.7 grams of water also at 22.3°C, after which the final temperature of the mixed solution was 36.9°C. If the specific heat of the mixed solution was 3.87 J/g·°C, determine ΔH for the dissolving process $\text{CsOH}(s) \rightarrow \text{CsOH}(aq)$ in kJ/mol CsOH.

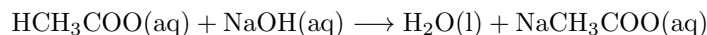
$$q_{\text{rxn lost}} = -q_{\text{soln gained}} = -(18.2 + 135.7)\text{g} \left(\frac{3.87 \text{ J}}{\text{g} \cdot ^\circ\text{C}} \right) (36.9 - 22.3)^\circ\text{C} \left(\frac{1 \text{ kJ}}{1000 \text{ J}} \right) = -8.696 \text{ kJ}$$

$$n_{\text{CsOH}} = 18.2 \text{ g} \left(\frac{1 \text{ mol}}{149.9 \text{ g}} \right) = 0.1214 \text{ mol CsOH}$$

$$\Delta H = \frac{-8.696 \text{ kJ}}{0.1214 \text{ mol CsOH}} = -71.6 \text{ kJ/mol CsOH}$$

11

In an insulated calorimeter, 55.7 mL of 1.91 M acetic acid was mixed with 62.6 mL of 1.83 M sodium hydroxide, with both solutions originally at 18.2°C. The final temperature of the mixed solutions was 30.1°C. The density of the mixed solutions was 1.03 g/mL and the specific heat of the mixed solutions was 3.96 J/g·°C. Write a balanced molecular equation, including physical states, and determine ΔH for the neutralization reaction in kJ/mol of water formed.



$$q_{\text{rxn lost}} = -q_{\text{soln gained}} = -(55.7 + 62.6)\text{mL} \left(\frac{1.03 \text{ g}}{1 \text{ mL}} \right) \left(\frac{3.96 \text{ J}}{\text{g} \cdot ^\circ\text{C}} \right) (30.1 - 18.2)^\circ\text{C} \left(\frac{1 \text{ kJ}}{1000 \text{ J}} \right) = -5.742 \text{ kJ}$$

$$55.7 \text{ mL} \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) \left(\frac{1.91 \text{ mol HCH}_3\text{COO}}{1 \text{ L}} \right) \left(\frac{1 \text{ mol H}_2\text{O}}{1 \text{ mol HCH}_3\text{COO}} \right) = 0.1064 \text{ mol H}_2\text{O}$$

$$62.6 \text{ mL} \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) \left(\frac{1.83 \text{ mol NaOH}}{1 \text{ L}} \right) \left(\frac{1 \text{ mol H}_2\text{O}}{1 \text{ mol NaOH}} \right) = 0.1147 \text{ mol H}_2\text{O}$$

HCH₃COO produces less H₂O, so HCH₃COO is the limiting reagent and 0.1064 mol H₂O is formed.

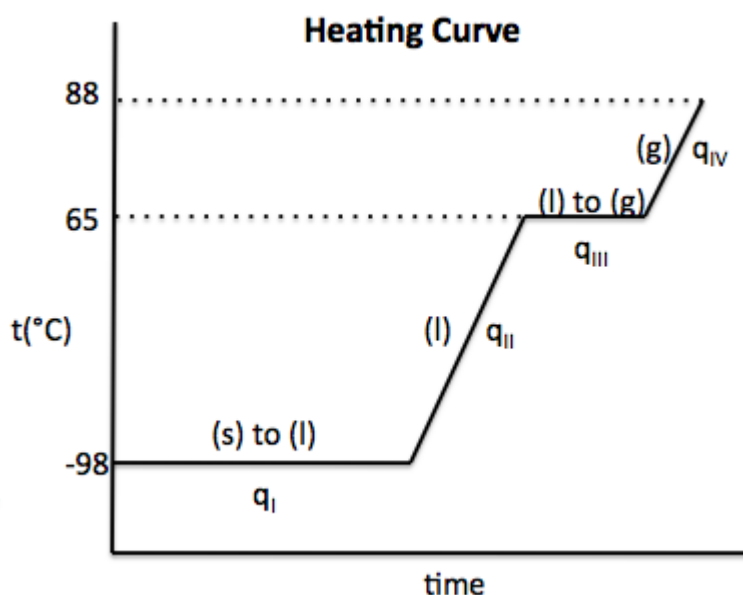
$$\Delta H = \frac{-5.742 \text{ kJ}}{0.1064 \text{ mol H}_2\text{O}} = -54.0 \text{ kJ/mol H}_2\text{O}$$

12

Consider the following data for methanol, CH₃OH:

melting point = -98°C
 boiling point = 65°C
 $\Delta H_{\text{fusion}} = 3.2 \text{ kJ/mol}$
 $\Delta H_{\text{vaporization}} = 38 \text{ kJ/mol}$
 specific heat of liquid methanol = $2.5 \text{ J/g}\cdot^{\circ}\text{C}$
 specific heat of methanol vapor = $1.7 \text{ J/g}\cdot^{\circ}\text{C}$

Sketch a heating curve that depicts solid methanol at -98°C being heated to 88°C and then calculate the total amount of heat in kilojoules absorbed when 77 grams of methanol undergoes this process.



I) solid methanol at -98°C to liquid methanol at -98°C :

$$q_{\text{I}} = 77 \text{ g} \left(\frac{1 \text{ mol}}{32.04 \text{ g}} \right) \left(\frac{3.2 \text{ kJ}}{1 \text{ mol}} \right) = 7.69 \text{ kJ}$$

II) liquid methanol at -98°C to liquid methanol at 65°C :

$$q_{\text{II}} = 77 \text{ g} \left(\frac{2.5 \text{ J}}{\text{g}\cdot^{\circ}\text{C}} \right) (65 - (-98))^{\circ}\text{C} \left(\frac{1 \text{ kJ}}{1000 \text{ J}} \right) = 31.4 \text{ kJ}$$

III) liquid methanol at 65°C to methanol vapor at 65°C :

$$q_{\text{III}} = 77 \text{ g} \left(\frac{1 \text{ mol}}{32.04 \text{ g}} \right) \left(\frac{38 \text{ kJ}}{1 \text{ mol}} \right) = 91.3 \text{ kJ}$$

IV) methanol vapor at 88°C to methanol vapor at 88°C :

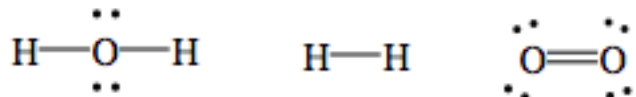
$$q_{IV} = 77 \text{ g} \left(\frac{1.7 \text{ J}}{\text{g} \cdot ^\circ \text{C}} \right) (88 - 65)^\circ \text{C} \left(\frac{1 \text{ kJ}}{1000 \text{ J}} \right) = 3.01 \text{ kJ}$$

$$q_{\text{total}} = 7.69 + 31.8 + 91.3 + 3.01 = 133 \text{ kJ}$$

13

Given the reaction $2\text{H}_2\text{O}(\text{g}) \rightarrow 2\text{H}_2(\text{g}) + \text{O}_2(\text{g})$ $\Delta H = 484 \text{ kJ/mol}$, use the table of average bond energies to calculate the H–H bond energy.

Lewis structures:



$$\Delta H = (2 \times 2)(\text{O} - \text{H}) - 2(\text{H} - \text{H}) - 1(\text{O} = \text{O})$$

$$484 \text{ kJ/mol} = 4(464 \text{ kJ/mol}) - 2(\text{H} - \text{H}) - 1(498 \text{ kJ/mol})$$

$$\text{bond energy of H} - \text{H} = 437 \text{ kJ/mol}$$

14

Given the reaction $2\text{C}_2\text{H}_2(\text{g}) + 5\text{O}_2(\text{g}) \rightarrow 4\text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$ $\Delta H^\circ = -2602 \text{ kJ/mol}$, use the information below to calculate the standard enthalpy of formation, ΔH_f° , for $\text{C}_2\text{H}_2(\text{g})$:

| Compound | ΔH_f° (kJ/mol) |
|--------------------------------|-----------------------------|
| $\text{CO}_2(\text{g})$ | -394 |
| $\text{H}_2\text{O}(\text{l})$ | -286 |

$$\Delta H^\circ = -2602 \text{ kJ/mol} = 4(-394 \text{ kJ/mol}) + 2(-286 \text{ kJ/mol}) - 2(\Delta H_f^\circ \text{ for } \text{C}_2\text{H}_2(\text{g})) - 5(0 \text{ kJ/mol})$$

$$\Delta H_f^\circ \text{ for } \text{C}_2\text{H}_2(\text{g}) = 227 \text{ kJ/mol}$$

15

ΔH for the dissolving process $\text{KClO}_3(\text{s}) \rightarrow \text{KClO}_3(\text{aq})$ is $+41.4 \text{ kJ/mol}$ KClO_3 . In an insulated calorimeter, 14.1 grams of solid KClO_3 at 24.6°C was dissolved

in 102.5 grams of water also at 24.6°C. If the specific heat of the mixed solution was 3.91 J/g·°C, determine the final temperature in the calorimeter.

$$q_{\text{rxn gained}} = 14.1 \text{ g KClO}_3 \left(\frac{1 \text{ mol KClO}_3}{122.6 \text{ g KClO}_3} \right) \left(\frac{41.4 \text{ kJ}}{1 \text{ mol KClO}_3} \right) \left(\frac{1000 \text{ J}}{1 \text{ kJ}} \right) = 4761 \text{ J}$$

$$q_{\text{soln lost}} = -q_{\text{rxn gained}} = -4761 \text{ J} = (14.1 + 102.5) \text{ g} \left(\frac{3.91 \text{ J}}{\text{g} \cdot ^\circ\text{C}} \right) (t_{\text{final}} - 24.6^\circ\text{C})$$

$$t_{\text{final}} = 14.2^\circ\text{C}$$

16

Consider the reaction $\text{Sr}(\text{OH})_2(\text{aq}) + 2\text{HNO}_3(\text{aq}) \longrightarrow 2\text{H}_2\text{O}(\text{l}) + \text{Sr}(\text{NO}_3)_2(\text{aq})$ $\Delta H^\circ = -112 \text{ kJ/mol}$. In an insulated calorimeter, 65.4 mL of 2.96 M strontium hydroxide at 22.5°C was mixed with 72.6 mL of 2.84 M nitric acid also at 22.5°C. If the density of the mixed solution was 1.06 g/mL and the specific heat of the mixed solution was 3.89 J/g·°C, determine the final temperature in the calorimeter.

$$65.4 \text{ mL} \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) \left(\frac{2.96 \text{ mol Sr}(\text{OH})_2}{1 \text{ L}} \right) \left(\frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol Sr}(\text{OH})_2} \right) = 0.3872 \text{ mol H}_2\text{O}$$

$$72.6 \text{ mL} \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) \left(\frac{2.84 \text{ mol HNO}_3}{1 \text{ L}} \right) \left(\frac{2 \text{ mol H}_2\text{O}}{2 \text{ mol HNO}_3} \right) = 0.2062 \text{ mol H}_2\text{O}$$

HNO₃ produces less H₂O, so HNO₃ is the limiting reagent and 0.2062 mol H₂O is formed in the reaction.

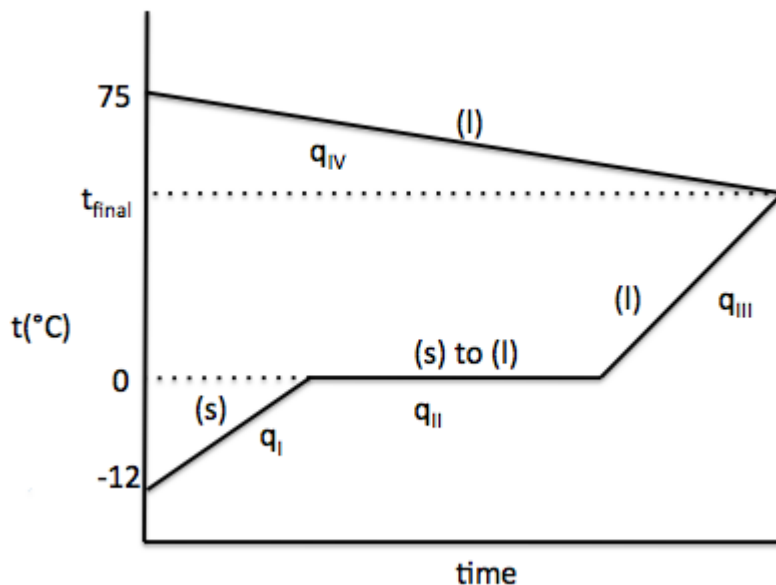
$$q_{\text{rxn lost}} = 0.2062 \text{ mol H}_2\text{O} \left(\frac{-112 \text{ kJ}}{2 \text{ mol H}_2\text{O}} \right) \left(\frac{1000 \text{ J}}{1 \text{ kJ}} \right) = -1.155 \times 10^4 \text{ J}$$

$$q_{\text{soln gained}} = -q_{\text{rxn lost}} = 1.155 \times 10^4 \text{ J} = (65.4 + 72.6) \text{ mL} \left(\frac{1.06 \text{ g}}{1 \text{ mL}} \right) \left(\frac{3.89 \text{ J}}{\text{g} \cdot ^\circ\text{C}} \right) (t_{\text{final}} - 22.5^\circ\text{C})$$

$$t_{\text{final}} = 42.8^\circ\text{C}$$

17

If 8.5 grams of ice at -12°C is added to an insulated calorimeter containing 65 grams of water at 75°C, and all the ice melts, sketch a heating/cooling curve for the process and determine the final temperature of the liquid water in the calorimeter.



I) ice at -12°C to ice at 0°C :

$$q_{\text{I}} = 8.5 \text{ g} \left(\frac{2.03 \text{ J}}{\text{g} \cdot ^{\circ}\text{C}} \right) (0 - (-12))^{\circ}\text{C}$$

II) ice at 0°C to liquid water at 0°C :

$$q_{\text{II}} = 8.5 \text{ g} \left(\frac{1 \text{ mol}}{18.02 \text{ g}} \right) \left(\frac{6.01 \text{ kJ}}{1 \text{ mol}} \right) \left(\frac{1000 \text{ J}}{1 \text{ kJ}} \right)$$

III) liquid water at 0°C to liquid water at t_{final} :

$$q_{\text{III}} = 8.5 \text{ g} \left(\frac{4.18 \text{ J}}{\text{g} \cdot ^{\circ}\text{C}} \right) (t_{\text{final}} - 0^{\circ}\text{C})$$

IV) liquid water at 75°C to liquid water at t_{final} :

$$q_{\text{IV}} = 65 \text{ g} \left(\frac{4.18 \text{ J}}{\text{g} \cdot ^{\circ}\text{C}} \right) (t_{\text{final}} - 75^{\circ}\text{C})$$

$$q_{\text{I}} + q_{\text{II}} + q_{\text{III}} = -q_{\text{IV}}, \text{ solve for } t_{\text{final}}$$

:

$$t_{\text{final}} = 56^{\circ}\text{C}$$

18

Calculate the lattice energy of potassium bromide using the information below. Show all relevant reactions, including states of matter.

$\Delta H_{\text{sublimation}}$ of potassium = 88 kJ/mol

$\Delta H_{\text{vaporization}}$ of Br_2 = 31 kJ/mol

Br_2 bond energy = 193 kJ/mol

first ionization energy of potassium = 419 kJ/mol

first electron affinity of bromine = -325 kJ/mol

ΔH_f° of solid potassium bromide = -394 kJ/mol

1a) $\text{K(s)} \rightarrow \text{K(g)}$ $\Delta H_{\text{sublimation}} = 88 \text{ kJ/mol}$

1b) $\frac{1}{2}\text{Br}_2(\text{l}) \rightarrow \frac{1}{2}\text{Br}_2(\text{g})$ $\frac{1}{2}\Delta H_{\text{vaporization}} = \frac{1}{2}(31 \text{ kJ/mol})$

2) $\frac{1}{2}\text{Br}_2(\text{g}) \rightarrow \text{Br(g)}$ $\frac{1}{2}(\text{Br}_2 \text{ bond energy}) = \frac{1}{2}(193 \text{ kJ/mol})$

3) $\text{K(g)} \rightarrow \text{K}^+(\text{g}) + \text{e}^-$ $I_1 = 419 \text{ kJ/mol}$

4) $\text{Br(g)} + \text{e}^- \rightarrow \text{Br}^-(\text{g})$ $\text{EA}_1 = -325 \text{ kJ/mol}$

5) $\text{K}^+(\text{g}) + \text{Br}^-(\text{g}) \rightarrow \text{KBr(s)}$ - L.E.

6) $\text{K(s)} + \frac{1}{2}\text{Br}_2(\text{l}) \rightarrow \text{KBr(s)}$ $\Delta H_f^\circ = -394 \text{ kJ/mol}$

$$88 \text{ kJ/mol} + \frac{1}{2}(31 \text{ kJ/mol}) + \frac{1}{2}(193 \text{ kJ/mol}) + 419 \text{ kJ/mol} + (-325 \text{ kJ/mol}) + (-\text{L.E.}) = -394 \text{ kJ/mol}$$

L.E. of KBr = 688 kJ/mol



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