Reaction Enthalpies

- Communicating Enthalpy Changes
(ΔH_r notation, balanced equation, potential energy diagrams)
- Hess' Law
- Enthalpy Changes using Formation Reactions
- Thermal Stability
- Multi-Step Energy Calculations

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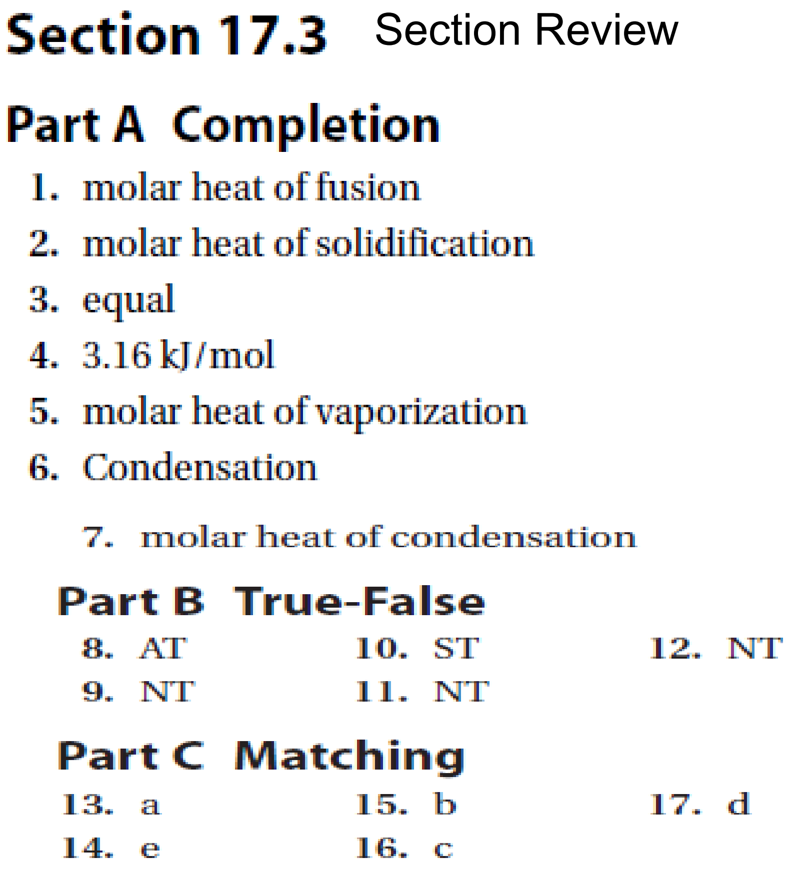

Section 17.3 Section Review

Part A Completion

1. molar heat of fusion
2. molar heat of solidification
3. equal
4. 3.16 kJ/mol
5. molar heat of vaporization
6. Condensation
7. molar heat of condensation

Part B True-False

8. AT
9. NT
10. ST
11. NT
12. NT

Part C Matching

13. a
14. e
15. b
16. c
17. d

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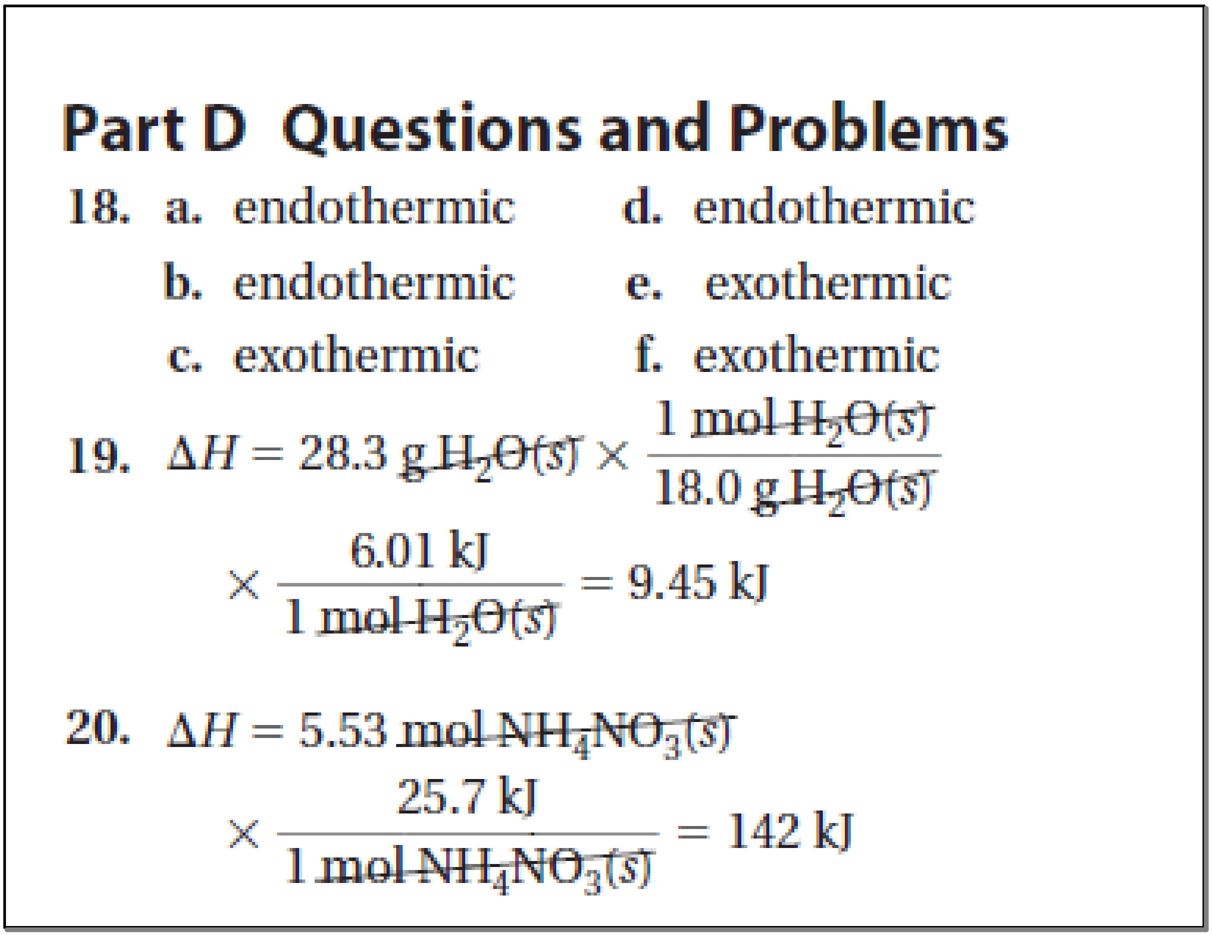

Part D Questions and Problems

18. a. endothermic d. endothermic
b. endothermic e. exothermic
c. exothermic f. exothermic

19. $$\Delta H = 28.3\ \cancel{\text{g H}_2\text{O}(s)} \times \frac{1\ \cancel{\text{mol H}_2\text{O}(s)}}{18.0\ \cancel{\text{g H}_2\text{O}(s)}} \times \frac{6.01\ \text{kJ}}{1\ \cancel{\text{mol H}_2\text{O}(s)}} = 9.45\ \text{kJ}$$

20. $$\Delta H = 5.53\ \cancel{\text{mol NH}_4\text{NO}_3(s)} \times \frac{25.7\ \text{kJ}}{1\ \cancel{\text{mol NH}_4\text{NO}_3(s)}} = 142\ \text{kJ}$$

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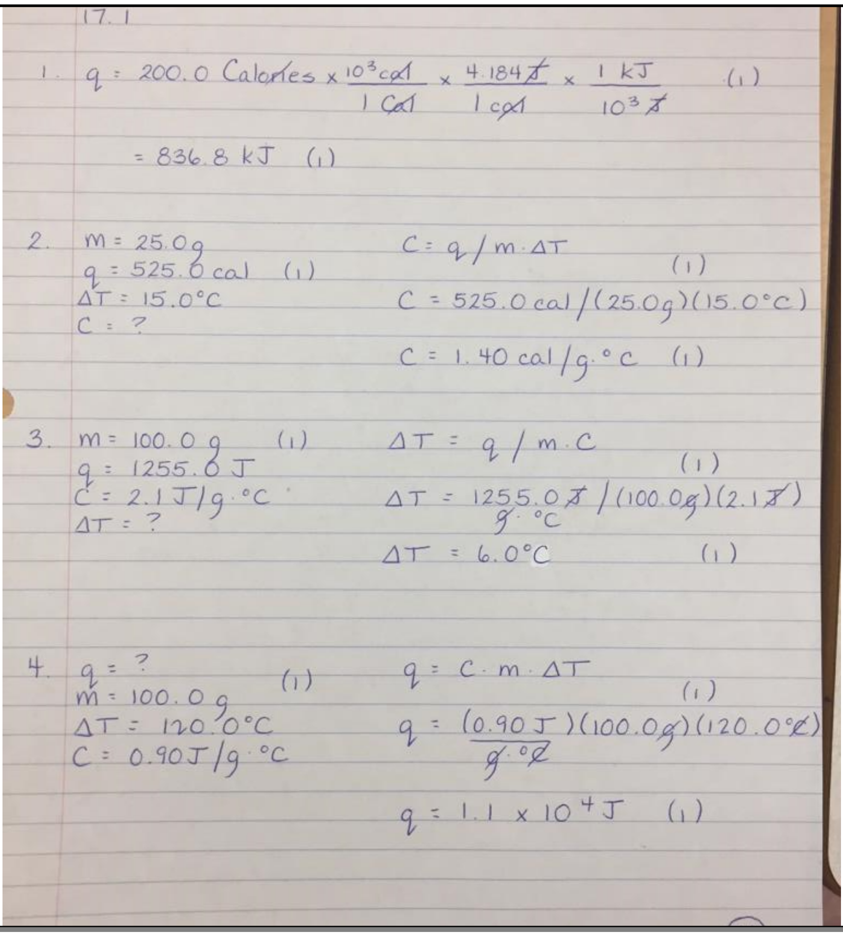

17.1

1. $q = 200.0\ \text{Calories} \times \frac{10^3\ \cancel{\text{cal}}}{1\ \cancel{\text{Cal}}} \times \frac{4.184\ \cancel{\text{J}}}{1\ \cancel{\text{cal}}} \times \frac{1\ \text{kJ}}{10^3\ \cancel{\text{J}}}$ (1)

$= 836.8\ \text{kJ}$ (1)

2. $m = 25.0\ \text{g}$
$q = 525.0\ \text{cal}$ (1)
$\Delta T = 15.0°\text{C}$
$C = ?$

$C = q / m \cdot \Delta T$ (1)
$C = 525.0\ \text{cal} / (25.0\ \text{g})(15.0°\text{C})$
$C = 1.40\ \text{cal}/\text{g}\cdot°\text{C}$ (1)

3. $m = 100.0\ \text{g}$ (1)
$q = 1255.0\ \text{J}$
$C = 2.1\ \text{J}/\text{g}\cdot°\text{C}$
$\Delta T = ?$

$\Delta T = q / m \cdot C$ (1)
$\Delta T = 1255.0\ \cancel{\text{J}} / (100.0\ \cancel{\text{g}})(2.1\ \frac{\cancel{\text{J}}}{\cancel{\text{g}}\cdot°\text{C}})$
$\Delta T = 6.0°\text{C}$ (1)

4. $q = ?$ (1)
$m = 100.0\ \text{g}$
$\Delta T = 120.0°\text{C}$
$C = 0.90\ \text{J}/\text{g}\cdot°\text{C}$

$q = C \cdot m \cdot \Delta T$ (1)
$q = \left(\frac{0.90\ \text{J}}{\cancel{\text{g}}\cdot\cancel{°\text{C}}}\right)(100.0\ \cancel{\text{g}})(120.0\ \cancel{°\text{C}})$
$q = 1.1 \times 10^4\ \text{J}$ (1)

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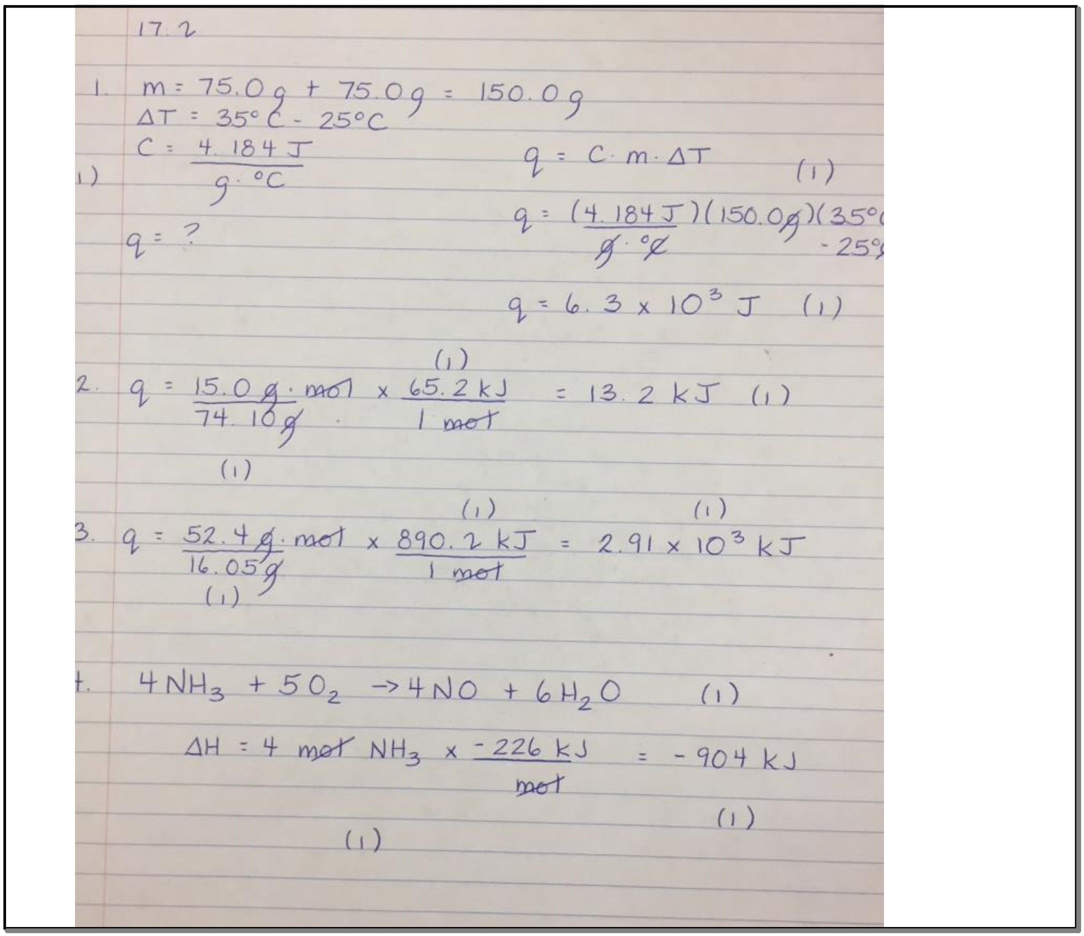

17.2

1. $m = 75.0\,g + 75.0\,g = 150.0\,g$

$\Delta T = 35°C - 25°C$

$C = \frac{4.184\,J}{g \cdot °C}$

1)

$q = ?$

$q = C \cdot m \cdot \Delta T$ (1)

$q = \left(\frac{4.184\,J}{\cancel{g} \cdot \cancel{°C}}\right)(150.0\,\cancel{g})(35°$ $- 25°$

$q = 6.3 \times 10^3\,J$ (1)

2. $q = \frac{15.0\,\cancel{g} \cdot \cancel{mol}}{74.10\,\cancel{g}} \times \overset{(1)}{\frac{65.2\,kJ}{1\,\cancel{mol}}} = 13.2\,kJ$ (1)

(1)

3. $q = \frac{52.4\,\cancel{g} \cdot \cancel{mol}}{16.05\,\cancel{g}} \times \overset{(1)}{\frac{890.2\,kJ}{1\,\cancel{mol}}} = \overset{(1)}{2.91 \times 10^3\,kJ}$

(1)

4. $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$ (1)

$\Delta H = 4\,\cancel{mol}\,NH_3 \times \frac{-226\,kJ}{\cancel{mol}} = -904\,kJ$

(1)

(1)

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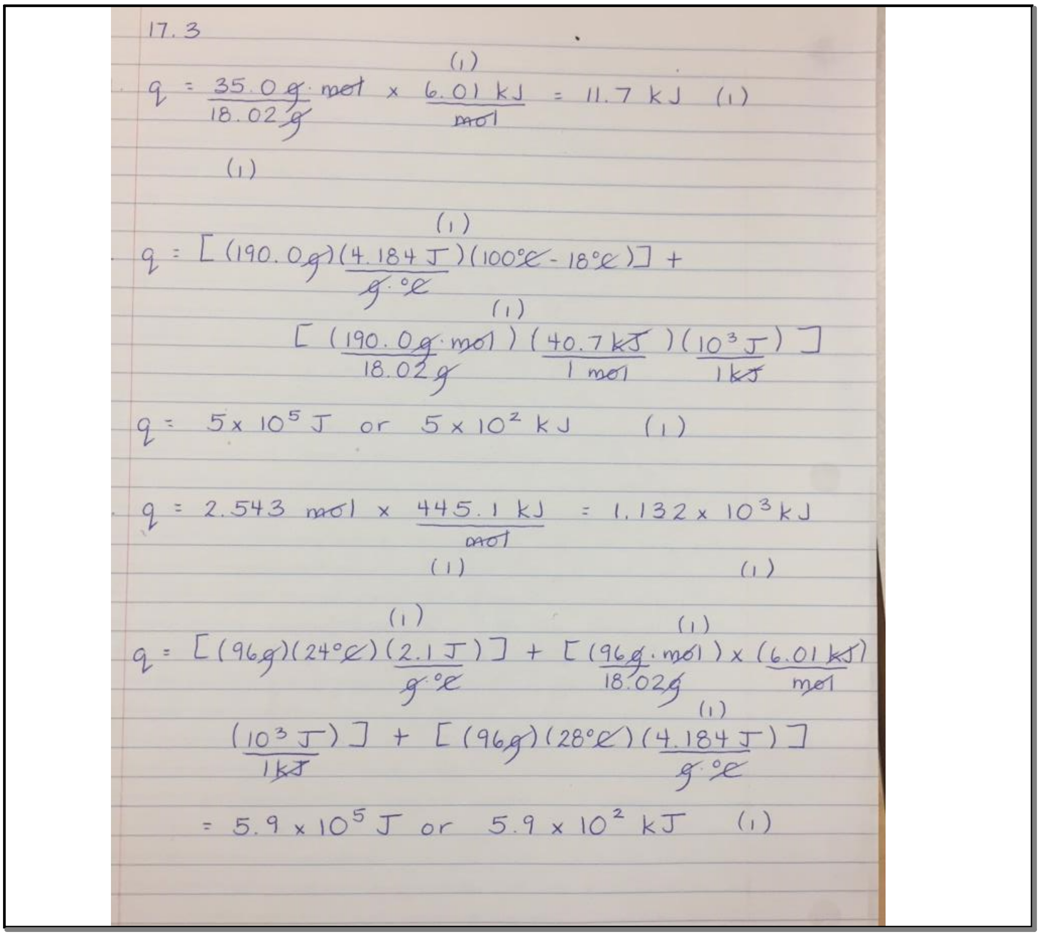

17.3

$q = \frac{35.0\,\cancel{g} \cdot \cancel{mol}}{18.02\,\cancel{g}} \times \overset{(1)}{\frac{6.01\,kJ}{\cancel{mol}}} = 11.7\,kJ$ (1)

(1)

$q = \overset{(1)}{[(190.0\,\cancel{g})\left(\frac{4.184\,J}{\cancel{g} \cdot \cancel{°C}}\right)(100\cancel{°C} - 18\cancel{°C})]} +$

$\overset{(1)}{\left[\left(\frac{190.0\,\cancel{g} \cdot \cancel{mol}}{18.02\,\cancel{g}}\right)\left(\frac{40.7\,\cancel{kJ}}{1\,\cancel{mol}}\right)\left(\frac{10^3\,J}{1\,\cancel{kJ}}\right)\right]}$

$q = 5 \times 10^5\,J$ or $5 \times 10^2\,kJ$ (1)

$q = 2.543\,\cancel{mol} \times \frac{445.1\,kJ}{\cancel{mol}} = 1.132 \times 10^3\,kJ$

(1) (1)

$q = \overset{(1)}{\left[(96\,\cancel{g})(24\cancel{°C})\left(\frac{2.1\,J}{\cancel{g} \cdot \cancel{°C}}\right)\right]} + \overset{(1)}{\left[\left(\frac{96\,\cancel{g} \cdot \cancel{mol}}{18.02\,\cancel{g}}\right) \times \left(\frac{6.01\,\cancel{kJ}}{\cancel{mol}}\right)\right.}$

$\left.\left(\frac{10^3\,J}{1\,\cancel{kJ}}\right)\right] + \overset{(1)}{\left[(96\,\cancel{g})(28\cancel{°C})\left(\frac{4.184\,J}{\cancel{g} \cdot \cancel{°C}}\right)\right]}$

$= 5.9 \times 10^5\,J$ or $5.9 \times 10^2\,kJ$ (1)

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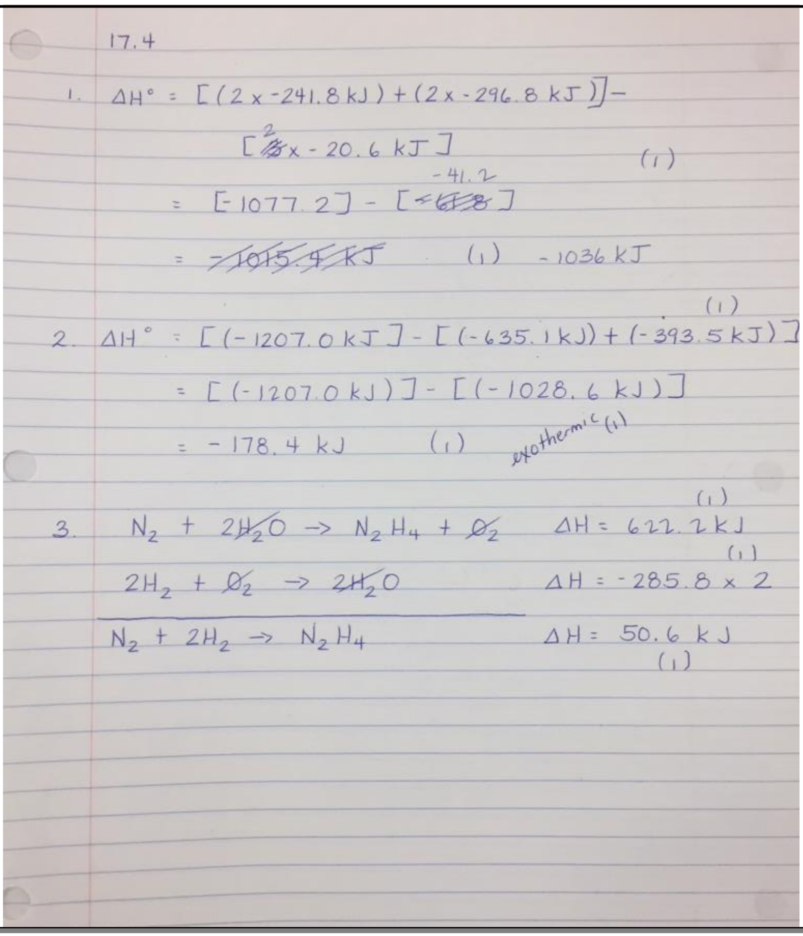

17.4

1. $\Delta H^\circ = [(2 \times -241.8\text{ kJ}) + (2 \times -296.8\text{ kJ})] - [\overset{2}{\cancel{3}} \times -20.6\text{ kJ}]$ (1)

$= [-1077.2] - [\overset{-41.2}{\cancel{-61.8}}]$

$= \cancel{-1015.4\text{ kJ}}$ (1) -1036 kJ

2. $\Delta H^\circ = [(-1207.0\text{ kJ}] - [(-635.1\text{ kJ}) + (-393.5\text{ kJ})]$ (1)

$= [(-1207.0\text{ kJ})] - [(-1028.6\text{ kJ})]$

$= -178.4\text{ kJ}$ (1) exothermic (1)

3. $N_2 + 2\cancel{H_2O} \rightarrow N_2H_4 + \cancel{O_2}$ $\Delta H = 622.2\text{ kJ}$ (1)

$2H_2 + \cancel{O_2} \rightarrow \cancel{2H_2O}$ $\Delta H = -285.8 \times 2$ (1)

$N_2 + 2H_2 \rightarrow N_2H_4$ $\Delta H = 50.6\text{ kJ}$ (1)

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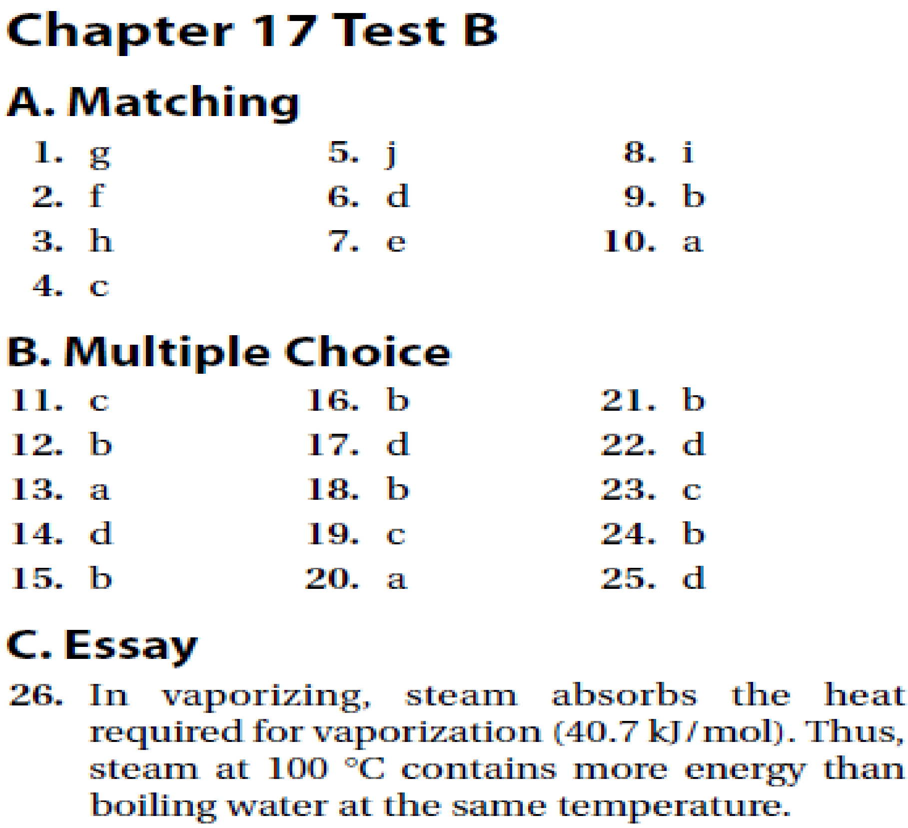

Chapter 17 Test B

A. Matching

1. g
2. f
3. h
4. c
5. j
6. d
7. e
8. i
9. b
10. a

B. Multiple Choice

11. c
12. b
13. a
14. d
15. b
16. b
17. d
18. b
19. c
20. a
21. b
22. d
23. c
24. b
25. d

C. Essay

26. In vaporizing, steam absorbs the heat required for vaporization (40.7 kJ/mol). Thus, steam at 100 °C contains more energy than boiling water at the same temperature.

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D. Problems

27. Specific heat $= \dfrac{75\text{ J}}{18\text{ g}\times 25°\text{C}}$

$= 0.17\text{ J/(g} \times °\text{C)}$

28. $\Delta H = m \times \text{C} \times \Delta T$

$$= \left(55.0\ \cancel{\text{mL}} \times 1.00\,\frac{\text{g}}{\cancel{\text{mL}}}\right) \times \left(4.18\,\frac{\text{J}}{\text{g}\times°\text{C}}\right) \times (33.0°\text{C} - 24.0°\text{C})$$

$$= 55.0\ \cancel{\text{g}} \times 4.18\,\frac{\text{J}}{\cancel{\text{g}}\times\cancel{°\text{C}}} \times 9.0\cancel{°\text{C}}$$

$$= 2.1 \times 10^3\text{ J} = 2.1\text{ kJ}$$

29. $8.00\text{ g}\,\cancel{\text{O}_2} \times \dfrac{1\ \cancel{\text{mol O}_2}}{32.0\text{ g}\,\cancel{\text{O}_2}} \times \dfrac{1411\text{ kJ}}{3\ \cancel{\text{mol O}_2}} = 118\text{ kJ}$

30. $75.0\ \cancel{\text{kJ}} \times \dfrac{1\ \cancel{\text{mol}}}{6.01\ \cancel{\text{kJ}}} \times \dfrac{18.0\text{ g}\ \cancel{\text{H}_2\text{O}}}{1\ \cancel{\text{mol}}\ \cancel{\text{H}_2\text{O}}} = 225\text{ g}$

31. $2[\text{C}(s) + \text{O}_2(g) \rightarrow \text{CO}_2(g)]\ \ 2(\Delta H = -393.5\text{ kJ})$
$2\text{CO}_2(g) \rightarrow 2\text{CO}(g) + \text{O}_2(g)\ \ \Delta H = +565.7\text{ kJ}$
$2\text{C}(s) + \text{O}_2(g) \rightarrow 2\text{CO}(g)$
$\Delta H = -787.0\text{ kJ} + 565.7\text{ kJ} = -221.3\text{ kJ}$

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32. $\text{C}_6\text{H}_6 + 15/2\text{O}_2(g) \rightarrow 6\text{CO}_2(g) + 3\text{H}_2\text{O}(l)$
$\Delta H = ?$
$\Delta H^0 = \Delta H_f^0\ (\text{products}) - \Delta H_f^0\ (\text{reactants})$
$= [6(-393.5\text{ kJ}) + 3(-285.8\text{ kJ})] - [(48.50\text{ kJ}) + (0.0\text{ kJ})]$
$= -3266.9\text{ kJ}$

Mar 26-9:12 AM

COMMUNICATING ENTHALPY CHANGES

① Using ΔH_r notation:
- for chemical reactions not well known, the chemical equation must accompany the enthalpy change. The molar enthalpy of reaction (or change in enthalpy) follows the equation. **For exothermic reactions the $\Delta H_r < 0$.** negative #

Ex. $Mg_{(s)} + 1/2\ O_{2(g)} \Longrightarrow MgO_{(s)}\quad \Delta H_r = -601.6\text{kJ}$

② The Enthalpy Change **(ΔH_r)** may be included as a term in the balanced equation:

(i) In endothermic reactions - energy is reported as a reactant and is transformed in the reaction.

Ex. $H_2O_{(l)} + 285.8\text{kJ} \Rightarrow H_{2(g)} + 1/2O_{2(g)}$

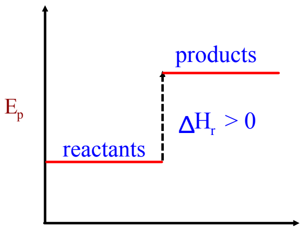
(ii) In exothermic reactions - energy is reported as a product since it is being produced.

Ex. $Mg_{(s)} + 1/2\ O_{2(g)} \Rightarrow MgO_{(s)} + 601.6\text{kJ}$

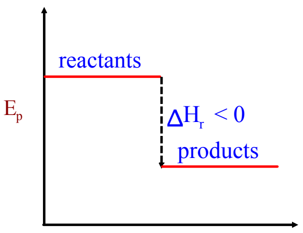
Nov 12-7:20 PM

③ POTENTIAL ENERGY DIAGRAMS
- may be used to express enthalpy change (ΔH_r)
- shows the potential energy of the reactants and products of a chemical reaction.
- shows the difference between the initial and final energies as the enthalpy change. (ΔH_r)

Endothermic Rxn



Exothermic Rxn



see Fig 11-8 p 373 (also 11-15,16,17)

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Predicting Energy Changes using Hess' Law

Hess' Law - (Heat of Summation)
- allows for the determination of the enthalpy change of a reaction with direct use of calorimetry.

Rules:

- if a chemical equation is reversed, then the sign of the ΔH_r changes
- if the coefficients of a chemical equation are altered by multiplying or dividing by a constant factor, then the ΔH_r is altered in the same way

Ex. $C_{(s)} + O_{2(g)} \Rightarrow CO_{2(g)}$ $\Delta H = -393.5$ kJ

Oct 28-8:16 AM

Enthalpies of Formation to Predict ΔH_r

$$\boxed{\Delta H_r = \underset{\text{products}}{\Delta H_{fp}} - \underset{\text{reactants}}{\Delta H_{fr}}}$$

$$\Delta H_r = \Sigma n H_{fp} - \Sigma n H_{fr}$$

knowing that $\Delta H = nH$

Nov 21-10:16 AM

Multi-Step Energy Calculations

Step 1: *Find H general* - equation
-use Hess' law
-from equation
-use calorimetry

Step 2: *Find n (specific)*
-use mass (molar mass)
-use concentration
-use $n = \Delta H / H$

Step 3: *Find ΔH (specific), mass,* ΔT, etc.

Nov 25-7:03 PM