## Answers to Review for Unit 3: Thermochemistry

1. c	11. d	21. a	31. d	41. c
2. b	12. c	22. d	32. c	42. b
3. d	13. a	23. a	33. a	43. a
4. b	14. d	24. a	34. b	44. a
5. c	15. d	25. c	35. c	45. d
6. d	16. b	26. c	36. a	46. a
7. c	17. c	27. с	37. c	47. a
8. a	18. a	28. b	38. d	48. b
9. a	19. d	29. c	39. d	49. d
10. c	20. c	30. d	40. b	50. d

## Answers to multiple choice:

**Answers for Calculations**. Be sure to show an equation, substitution step and your final answer. Include all units. Round your final answer to the correct number of sig digs.

1. In an experiment, 5.260 g of potassium hydroxide is mixed with a dilute solution of hydrochloric acid. A neutralization (double displacement) reaction occurs. The hydrochloric acid is in excess. The following date are obtained:

Initial temperature of hydrochloric acid	24.5 °C
Volume of hydrochloric acid	200.0 mL
Final temperature of solution after mixing	31.4 °C

Calculate the molar heat of reaction ( $\Delta H$ ) per mole of potassium hydroxide.

KOH (s) + HCl (aq)  $\rightarrow$  KCl (aq) + H<sub>2</sub>O (l)  $\Delta H = ?$ 

- The temperature of the water in the HCl (the system) increases after the potassium hydroxide has dissolved. This is an exothermic reaction (the reaction releases heat), so the sign for  $\Delta H$  will be negative.
- Calculate the amount of heat (Q) released when the KOH reacts:

Given: $m = 200.0 \text{ g}$	$O = m \cdot c \cdot \Delta T$
$c = 4.184 \text{ J/g}^{\circ}\text{C}$	$= 200.0 \text{ g x } 4.184 \text{ J/g}^{\circ}\text{C} \text{ x } 6.9 ^{\circ}\text{C}$
$\Delta T = T_2 - T_1$	= 5773.92 J
= 31.4  °C - 24.5  °C	= 5.774  kJ
$= 6.9 ^{\circ}\mathrm{C}$	

- The heat absorbed by the system (5.774 kJ) is equal to the heat released when KOH reacted; therefore, when 5.260 g of KOH is dissolved, it releases 5.774 kJ of thermal energy.
- $Q = -\Delta H$  so  $\Delta H = -5.774 \text{ kJ}$
- To find  $\Delta$ H/mol, divide  $\Delta$ H by the number of moles of KOH:

n <sub>KOH</sub>	$= \underline{m}$ . MM <sub>KOH</sub>	$\Delta$ H/mole = $\Delta$ H reaction moles KOH reacted
	$= \frac{5.260 \text{ g}}{56.11 \text{ g/mol}}$ = 0.09374 mol of KOH	$= \frac{-5.774 \text{ kJ}}{0.09374 \text{ mol}}$
	- 0.07574 mor or Kom	= -61.6 kJ/mol KOH (3 sig digs)

2. An experiment was conducted using the reaction:

$$HNO_{3(aq)} + KOH_{(aq)} \rightarrow H_2O_{(l)} + KNO_{3(aq)}$$

The following data were collected:

	55 0T
volume of 1.3-M HNO <sub>3</sub>	55.0 mL
initial temperature of HNO <sub>3</sub>	23.5 °C
volume of 1.3-M KOH	60.0 mL
initial temperature of KOH	23.5 °C
final temperature after mixing	31.8 °C

Calculate the heat of reaction,  $\Delta H$ , expressed in kJ per mole of HNO<sub>3</sub>.

- The temperature of the water in the HNO<sub>3</sub> (the system) increases after the potassium hydroxide has dissolved. This is an exothermic reaction (the reaction releases heat), so the sign for  $\Delta$ H will be negative.
- Calculate the amount of heat (Q) released when the KOH and HNO<sub>3</sub> react:

## Given:

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Total volume of mixture = 115.0 mL

m = 115.0 \text{ g}

c = 4.184 \text{ J/g}^{\circ}\text{C}

\Delta T = T_2 - T_1

= 31.8 ^{\circ}\text{C} - 23.5 ^{\circ}\text{C}

= 8.3 ^{\circ}\text{C}

Q = m · c · \Delta T

= 115.0 \text{ g x } 4.184 \text{ J/g}^{\circ}\text{C x } 8.3 ^{\circ}\text{C}

= 3993.63 \text{ J}

= 3.994 \text{ kJ}
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- The heat absorbed by the system (3.994 kJ) is equal to the heat released when the KOH and  $HNO_3$  reacted
- $Q = -\Delta H$  so  $\Delta H = -3.994 \text{ kJ}$
- To find  $\Delta$ H/mol, divide  $\Delta$ H by the number of moles of HNO<sub>3</sub>:

n <sub>HNO3</sub>	$= C \times V$	$\Delta$ H/mole = $\Delta$ H reaction
		moles HNO <sub>3</sub> reacted
	$= 1.3 \text{ mol/L} \times 0.0550 \text{ L}$	
		= <u>- 3.994 kJ</u>
	$= 0.0715 \text{ mol HNO}_3$	0.0715 mol
		= - 55.9 kJ/mol HNO <sub>3</sub>

=  $-56 \text{ kJ/mol HNO}_3$  (2 sig digs from 1.3 M)

3.	$\Delta H^\circ = -2728 \text{ kJ}$	process:	equation (1) x 2, equation (2) x 3, equation (3) flipped
4.	$\Delta H^{\circ} = -699.5 \text{ kJ}$	process:	equation (4) flipped equation (1) x <sup>1</sup> / <sub>2</sub> equation (2) x 3/2 equation (3) x 3
		alternatively:	equation (4) flipped and multiplied by 2 equation (1) as is equation (2) x 3 equation (3) x 6 then divide the final equation and $\Delta H$ by 2
5.	$\Delta H^{\circ} = + 221.4 \text{ kJ}$	process:	equation (3) as is equation (1) as is equation (2) flipped equation (4) flipped and multiplied by 2

6. For the reaction:  $2 \text{ HNO}_{3(l)} + \text{Mg(OH)}_{2(s)} \rightarrow \text{Mg(NO}_{3})_{2(aq)} + 2 \text{ H}_2O_{(l)}$   $\Delta H^\circ \text{rxn} = \Sigma \Delta H^\circ_f \text{ products} - \Sigma \Delta H^\circ_f \text{ reactants}$   $= [\Delta H^\circ_f \text{ Mg(NO}_{3})_{2(aq)} + 2 (\Delta H^\circ_f \text{ H}_2O_{(l)})] - [2 (\Delta H^\circ_f \text{ HNO}_{3(l)}) + \Delta H^\circ_f \text{ Mg(OH)}_{2(s)}]$  = [-875.0 kJ + 2 (-285.8 kJ)] - [2 (-174.1 kJ) + (-924.5 kJ)] = [-1446.6 kJ] - [-1272.7 kJ]= -173.9 kJ 7. Liquid butane,  $C_4H_{10}$ , burns to produce water vapour and carbon dioxide. The  $\Delta H$  for the reaction is -2662.8 kJ/mol butane burned. Use this value and  $\Delta H^{\circ}_{f}$  values on page 597 to calculate the heat of formation ( $\Delta H^{\circ}_{f}$ ) of butane.

 $C_{4}H_{10}(l) + 13/2 \quad O_{2}(g) \rightarrow 4 \quad CO_{2}(g) + 5 \quad H_{2}O(g) \qquad \Delta H = -2662.8 \text{ kJ/mol butane}$   $\Delta H^{\circ} rxn = \Sigma \Delta H^{\circ}{}_{f} \text{ products} - \Sigma \Delta H^{\circ}{}_{f} \text{ reactants}$   $-2662.8 \text{ kJ} = [4 (\Delta H^{\circ}{}_{f} \text{ CO}_{2(g)}) + 5 (\Delta H^{\circ}{}_{f} \text{ H}_{2}O_{(g)})] - [\Delta H^{\circ}{}_{f} \text{ C}_{4}\text{H}_{10(l)}]$   $-2662.8 \text{ kJ} = [4 (-393.5 \text{ kJ}) + 5 (-241.8 \text{ kJ})] - [\Delta H^{\circ}{}_{f} \text{ C}_{4}\text{H}_{10(l)}]$   $-2662.8 \text{ kJ} = [-2783 \text{ kJ}] - [\Delta H^{\circ}{}_{f} \text{ C}_{4}\text{H}_{10(l)}]$   $120.2 \text{ kJ} = - [\Delta H^{\circ}{}_{f} \text{ C}_{4}\text{H}_{10(l)}]$   $\Delta H^{\circ}{}_{f} \text{ C}_{4}\text{H}_{10(l)} = -120.2 \text{ kJ}$ 

8. Use bond energies to estimate the enthalpy change ( $\Delta$ H) for the reaction when solid nitrogen triiodide (NI<sub>3</sub>, a contact explosive) decomposes to produce nitrogen gas and pure iodine vapour:

$$2 \operatorname{NI}_{3(s)} \rightarrow \operatorname{N}_{2(g)} + 3 \operatorname{I}_{2(v)}$$

Express your answer in kJ/mol of NI<sub>3</sub>.

$$\begin{array}{c} \mathbf{I} - \mathbf{N} - \mathbf{I} \\ \mathbf{I} \\ \mathbf{I} \\ \mathbf{I} \\ \mathbf{I} \end{array} + \begin{array}{c} \mathbf{I} - \mathbf{N} - \mathbf{I} \\ \mathbf{I} \\ \mathbf{I} \end{array} \rightarrow \begin{array}{c} \mathbf{N} \equiv \mathbf{N} \\ \mathbf{N} \equiv \mathbf{N} \\ \mathbf{I} \\ \mathbf{I} \end{array} + 3 \begin{array}{c} \mathbf{I} - \mathbf{I} \\ \mathbf{I} \\ \mathbf{I} \end{array}$$

 $\Delta H^{\circ}$  rxn =  $\Sigma$  bond energy reactants -  $\Sigma$  bond energy products

- =  $\Sigma [6(N-I)] [1(N \equiv N) + 3(I-I)]$
- =  $\Sigma [6(159 \text{ kJ})] [1(945 \text{ kJ}) + 3(151 \text{ kJ})]$
- = 954 kJ 1398 kJ
- = -444 kJ for 2 moles of NI<sub>3</sub> or -222 kJ per mole of NI<sub>3</sub>

9. Use bond energies to estimate  $\Delta H$  for the reaction:

 $C_2H_4 \hspace{.1in} + \hspace{.1in} Cl_2 \hspace{.1in} \rightarrow \hspace{.1in} C_2H_4Cl_2$ 

Express your answer in kJ/mol of C<sub>2</sub>H<sub>4</sub>.

Express your answer in kJ/mol of C<sub>2</sub>H<sub>4</sub>.  

$$H - C \equiv C - H + : Cl - Cl = Cl + H - C - C - H + H H H$$

$$H - H H H H H$$

 $\Delta H^{\circ}$  rxn =  $\Sigma$  bond energy reactants -  $\Sigma$  bond energy products

$$= \Sigma [4 (C - H) + 1 (C = C) + 1 (Cl - Cl)] - [1 (C - C) + 4 (C - H) + 2 (C - Cl)]$$
  
=  $\Sigma [4 (338 kJ) + 1 (607 kJ) + 1 (243 kJ)] - [1 (347 kJ) + 4 (338 kJ) + 2 (397 kJ)]$   
= 2202 kJ - 2493 kJ

= -291 kJ for 1 mole of C<sub>2</sub>H<sub>4</sub>