

### Answers to Review for Unit 3: Thermochemistry

#### Answers to multiple choice:

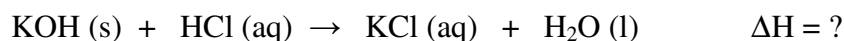
1. c	11. d	21. a	31. d	41. c
2. b	12. c	22. d	32. c	42. b
3. d	13. a	23. a	33. a	43. a
4. b	14. d	24. a	34. b	44. a
5. c	15. d	25. c	35. c	45. d
6. d	16. b	26. c	36. a	46. a
7. c	17. c	27. c	37. c	47. a
8. a	18. a	28. b	38. d	48. b
9. a	19. d	29. c	39. d	49. d
10. c	20. c	30. d	40. b	50. d

**Answers 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 data 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.



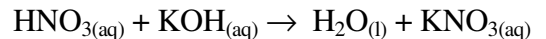
- The temperature of the water in the HCl (the system) increases after the potassium hydroxide has dissolved. This is an exothermic reaction (the reaction releases heat), so the sign for  $\Delta H$  will be negative.
- Calculate the amount of heat ( $Q$ ) released when the KOH reacts:

$$\begin{aligned} \text{Given: } m &= 200.0 \text{ g} & Q &= m \cdot c \cdot \Delta T \\ c &= 4.184 \text{ J/g}^\circ\text{C} & &= 200.0 \text{ g} \times 4.184 \text{ J/g}^\circ\text{C} \times 6.9^\circ\text{C} \\ \Delta T &= T_2 - T_1 & &= 5773.92 \text{ J} \\ &= 31.4^\circ\text{C} - 24.5^\circ\text{C} & &= 5.774 \text{ kJ} \\ &= 6.9^\circ\text{C} & & \end{aligned}$$

- 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  $\Delta H/\text{mol}$ , divide  $\Delta H$  by the number of moles of KOH:

$$\begin{aligned} n_{\text{KOH}} &= \frac{m}{\text{MM}_{\text{KOH}}} & \Delta H/\text{mole} &= \frac{\Delta H_{\text{reaction}}}{\text{moles KOH reacted}} \\ &= \frac{5.260 \text{ g}}{56.11 \text{ g/mol}} & &= \frac{-5.774 \text{ kJ}}{0.09374 \text{ mol}} \\ &= 0.09374 \text{ mol of KOH} & &= -61.6 \text{ kJ/mol KOH (3 sig digs)} \end{aligned}$$

2. An experiment was conducted using the reaction:



The following data were collected:

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:**

Total volume of mixture = 115.0 mL

$$m = 115.0 \text{ g}$$

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

$$\Delta T = T_2 - T_1$$

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

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

$$Q = m \cdot c \cdot \Delta T$$

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

$$= 3993.63 \text{ J}$$

$$= 3.994 \text{ kJ}$$

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

$$n_{\text{HNO}_3} = C \times V$$

$$= 1.3 \text{ mol/L} \times 0.0550 \text{ L}$$

$$= 0.0715 \text{ mol HNO}_3$$

$$\Delta H/\text{mole} = \frac{\Delta H \text{ reaction}}{\text{moles HNO}_3 \text{ reacted}}$$

$$= \frac{-3.994 \text{ kJ}}{0.0715 \text{ mol}}$$

$$= -55.9 \text{ kJ/mol HNO}_3$$

$$= -55.9 \text{ kJ/mol HNO}_3$$

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

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  $\frac{1}{2}$   
equation (2) x  $\frac{3}{2}$   
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:  $2 \text{HNO}_3(l) + \text{Mg}(\text{OH})_2(s) \rightarrow \text{Mg}(\text{NO}_3)_2(aq) + 2 \text{H}_2\text{O}(l)$

$$\Delta H^\circ \text{ rxn} = \sum \Delta H^\circ_f \text{ products} - \sum \Delta H^\circ_f \text{ reactants}$$

$$= [ \Delta H^\circ_f \text{Mg}(\text{NO}_3)_2(aq) + 2 (\Delta H^\circ_f \text{H}_2\text{O}(l)) ] - [ 2 (\Delta H^\circ_f \text{HNO}_3(l)) + \Delta H^\circ_f \text{Mg}(\text{OH})_2(s) ]$$

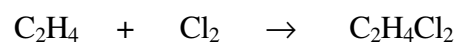
$$= [ -875.0 \text{ kJ} + 2 (-285.8 \text{ kJ}) ] - [ 2 (-174.1 \text{ kJ}) + (-924.5 \text{ kJ}) ]$$

$$= [ -1446.6 \text{ kJ} ] - [ -1272.7 \text{ kJ} ]$$

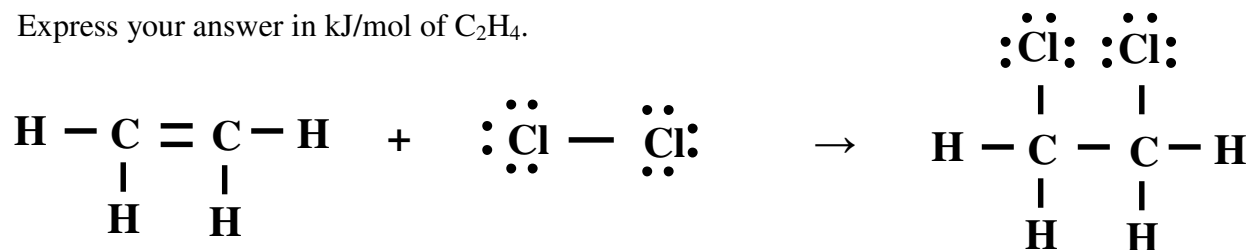
$$= -173.9 \text{ kJ}$$



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



Express your answer in kJ/mol of  $\text{C}_2\text{H}_4$ .



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

$$= \Sigma [ 4 (\text{C} - \text{H}) + 1 (\text{C} = \text{C}) + 1 (\text{Cl} - \text{Cl}) ] - [ 1 (\text{C} - \text{C}) + 4 (\text{C} - \text{H}) + 2 (\text{C} - \text{Cl}) ]$$

$$= \Sigma [ 4 (338 \text{ kJ}) + 1 (607 \text{ kJ}) + 1 (243 \text{ kJ}) ] - [ 1 (347 \text{ kJ}) + 4 (338 \text{ kJ}) + 2 (397 \text{ kJ}) ]$$

$$= 2202 \text{ kJ} - 2493 \text{ kJ}$$

$$= -291 \text{ kJ for 1 mole of } \text{C}_2\text{H}_4$$