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Practice Test - Chapter 5 - Thermochemistry

Target 1. I can give examples of different forms of energy. I can also state common units of energy and convert between these units.

1. Choose the correct type of energy for each of the following: Choices: ~~potential energy, kinetic energy, thermal energy, internal energy, electrostatic potential energy~~

KINETIC ENERGY

a) The energy of motion.

POTENTIAL ENERGY

b) The energy of an object due to its relative position.

ELECTROSTATIC P.E.

c) The energy due to the attractions or repulsions of charged particles such as protons and electrons.

THERMAL ENERGY

d) The energy that a substance possess because of its temperature.

INTERNAL ENERGY

e) The sum of all the PE and KE of a system.

2. Circle all of the following which are units of energy:

joule kilogram gram calorie Calorie kilojoule kg•m²/s² m²/s²

same as a "joule"

3. Fill in the blanks: 1 calorie = 4.184 J = 0.001 Cal

Target 2. I can define the first law of thermodynamics and write the associated equation.

4. The First Law of Thermodynamics states that ... energy is conserved!

Energy lost by system is gained by surroundings (and visa-versa).

5. Which equation represents the First Law of Thermodynamics?

- a. $\Delta E = q + w$
b. $\Delta H = q_p$
c. $q = m \cdot C_p \cdot \Delta T$
d. $KE = \frac{1}{2} mv^2$

Target 3. I can describe how the change in internal energy of a system is related to the changes of heat and work between the system and its surroundings.

6. A positive "q" means that the system _____ heat. A positive "w" means that work is done _____ the system.
- a. loses, by b. loses, on c. gains, by **d. gains, on**
7. What is ΔE for a system when it does 230 kJ of work on its surroundings and 130 kJ of heat is removed from the system?
- a. +100 kJ b. -100 kJ c. +360 kJ **d. -360 kJ**

$$\Delta E = q + w = -130 \text{ kJ} + (-230 \text{ kJ}) = -360 \text{ kJ}$$

Target 4. I can define enthalpy and relate the enthalpy change in a process to the heat added (endothermic) to or lost (exothermic) by the system during the process.

8. Enthalpy is the same thing as heat when measured at a constant _____. $\Delta H = q_p$
- a. volume b. temperature **c. pressure** d. elevation
9. A negative ΔH indicates that a reaction is ... **EXOTHERMIC! loses heat**
- a. endothermic.
b. absorbing heat.
c. occurring spontaneously.
d. nonspontaneous.
e. losing heat.
10. Assume that a chemical reaction is occurring in a plastic bag and that you are holding the bag in your hands. If your hands feel cold it is because ...
- a. the chemical reaction is exothermic and absorbing heat from your hands.
b. the chemical reaction is endothermic and absorbing heat from your hands.
c. the chemical reaction is exothermic and releasing heat into your hands.
d. the chemical reaction is endothermic and releasing heat into your hands.

Target 5. I can define the term *state function* and be provide examples of state functions.

11. A state function is defined as a property of a system ...
- a. which never changes.
b. which only changes for intensive properties and not for extensive properties.
c. which only changes for extensive properties and not for intensive properties.
d. determined at standard conditions (25°C and 1 atm).
e. determined by its state and not by how it got to that state.
12. Circle all of the following which are considered STATE functions:

mass **volume** **kinetic energy** **potential energy** work **enthalpy** height

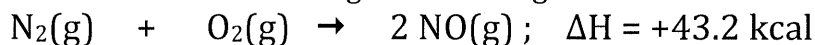
m V KE PE w H h

capital letters
"

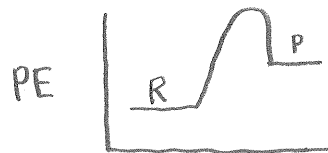
Target 6. I can differentiate between exothermic and endothermic processes. I can also sketch an exothermic and endothermic energy diagram and label the various parts.

14. Which ONE of the following statements is false?
- The reaction vessel cools when an endothermic reaction occurs.
 - An exothermic reaction is characterized by a negative value of ΔH .
 - Heat is evolved when an exothermic reaction occurs.
 - Heat is added to the system by an endothermic reaction.
 - An endothermic reaction causes the surroundings to ~~absorb~~ ^{lose} heat.

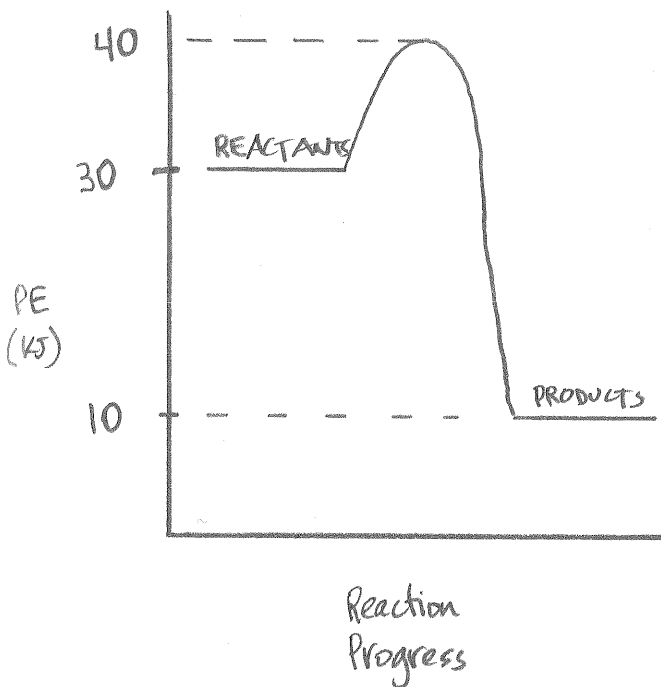
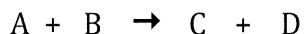
15. Which ONE statement concerning the following is correct?



- The reaction is exothermic.
- The products have an enthalpy loss.
- The sign of ΔH for the reaction is negative.
- PE of the products exceeds that of the reactants.
- The products have less enthalpy than the reactants.



16. Consider the following energy diagram:



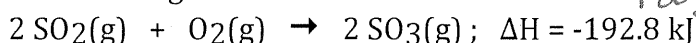
- a. Is the reaction exothermic or endothermic?

- b. Label the reactants and products.

- c. What is the value of ΔH ? -20 KJ

Target 7. I can make stoichiometric calculations based upon a thermochemical equation.

17. Consider the following reaction:

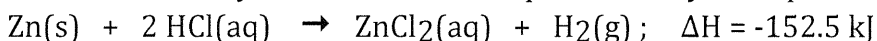


Assume enough SO_2 and O_2 were reacted and 200. grams of SO_3 were produced. Calculate the corresponding amount of heat produced.

- a. -3.86 kJ b. -15.1 kJ c. -75.5 kJ **(d) -241 kJ** e. -521 kJ

$$x \text{ kJ} = \frac{200 \text{ g SO}_3}{80 \text{ g SO}_3} \times \frac{1 \text{ mol SO}_3}{2 \text{ mol SO}_3} \times (-192.8 \text{ kJ}) = \frac{(100)(-193)}{80} = \frac{10}{8}(-193) = \frac{5}{4}(-193) = -240$$

18. The reaction of zinc with hydrochloric acid is represented by the equation:



In which of the following experiments would the temperature rise the most?

Zn 65.4 g/mol

- a. 0.65 g of Zn added to 1.0 L of 0.02 M HCl
(b) 1.3 g of Zn added to 1.0 L of 0.04 M HCl
 c. 1.3 g of Zn added to 1.0 L of 0.03 M HCl
 d. 2.6 g of Zn added to 1.0 L of 0.03 M HCl
 e. 5.2 g of Zn added to 1.0 L of 0.02 M HCl

	mol Zn	mol HCl
a.	0.01	.02
b.	0.02	.04
c.	0.02	.03
d.	0.04	.03
e.	0.04	.02

Use bal. eq. to determine which has MOST chem. reacting!

Target 8. I can differentiate between the terms *heat capacity* and *specific heat*.

19. Typical heat capacity units are _____ and specific heat capacity units are _____.

- (a)** J/°C, J/g·°C b. J/g·°C, J/°C c. °C/J, °C·J/g d. °C·J/g, °C/J

20. Which 2 statements are TRUE?

- (a)** Specific heat is specifically for 1 gram of a substance.
 b. Specific heat is an extensive property.
(c) The heat capacity of water is dependent upon the amount of water present.
 d. The specific heat of water is relatively low compared to other common substances.
 e. The specific heat of a substance is independent of its state.

Target 9. I can solve problems using $q = m \cdot C_p \cdot \Delta T$; calculate any one of the quantities given the other three.

21. What is the final temperature of 30. g of Al if 540. J of heat is added to a sample at 25.0°C? (specific heat of Al = 0.90 J/g·°C)

- a. 10°C b. 20°C c. 35°C **(d) 45°C** d. 135°C

$$q = (m)(C_p)(\Delta T); \quad 540. \text{ J} = (30 \text{ g})(0.90 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}})(\Delta T)$$

$$540 = 27 (\Delta T)$$

$$\Delta T = 20^\circ\text{C} + 25^\circ\text{C} = 45^\circ\text{C}$$

22. Assume that 100.0 g of water 30.0°C is mixed with 50.0 g of water at 0.0°C. What is the final temperature?

- a. 40.0 °C **b.** 20.0 °C c. 15.0 °C d. 10.0 °C

23. How much heat is lost when 10.0 g of iron cools from 129°C to 79°C? (The specific heat of iron is 0.450 J/g°C).

- a. 12.0 J b. 24 J c. 124 J d. 175 J **e.** 225 J

24. Consider the following specific heats of metals:

<u>Metal</u>	<u>Specific Heat (J/g°C)</u>
Lithium	3.56
Gallium	0.372
Nickel	0.444
Gold	0.129
Sodium	1.23

If the same amount of heat is added to 200 g samples of each of the metals, assume all metals are at the same temperature, which metal will attain the lowest temperature?

- highest Cp*
a. Lithium b. Gallium c. Nickel d. Gold e. Sodium

25. 500 mL of a 1.0 M NaOH are added to 500 mL of a 1.0 M HCl in a styrofoam cup calorimeter. The solution is stirred and the rise in temperature (ΔT_1) is measured. The experiment is repeated using 100 mL of each solution, and the rise in temperature (ΔT_2) is measured. It is found that _____.

- a. ΔT_2 is 5 times larger than ΔT_1
b. ΔT_1 is 5 times larger than ΔT_2
c. ΔT_2 is greater than ΔT_1
d. ΔT_2 is equal to ΔT_1
e. ΔT_1 is less than ΔT_2

10. I can describe the difference between a coffee cup calorimeter and a bomb calorimeter. I can also be able to solve simple bomb calorimetry problems.

26. When is ΔH approximately equal to ΔE ?

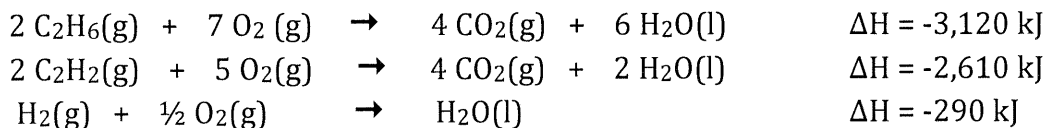
- a. at constant pressure **b.** at constant volume c. when the pressure is large

↳ as in a bomb calorimeter

27. A 3.0 gram sample of an organic compound was burned in a bomb calorimeter. The temperature of the calorimeter increased from 25.0°C to 35.0°C. The heat capacity of the calorimeter was 9.0 kJ/°C. What is the amount of heat released per gram (kJ/g) of the organic substance?
- a. 3.0 b. 6.0 c. 9.0 **d. 30.** e. 90.

Target 11. I can state Hess's law and use Hess's Law to solve problems.

28. State Hess's Law. *ΔH for an overall reaction is equal to the sum of the ΔH 's of a series of reactions ... assuming the series of reactions "adds up" to the overall reaction.*
29. What is the heat of hydrogenation of acetylene, at 25°C and 1 atm,
- $$\text{C}_2\text{H}_2(\text{g}) + 2 \text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_6(\text{g})$$
- given the following thermochemical equations:



- a.** -325 kJ b. +325 kJ c. +1,610 kJ d. -1,610 kJ

Target 12. I can define the terms *standard state* and *standard heat of formation*. I will also be to write a chemical reaction associated with the standard heat of formation.

30. Which of the following conditions represent standard states?
- a. 1 atm & 0°C b. 2 atm & 0°C c. 2 atm & 273°C **d. 1 atm & 25°C**
31. Which equation represents ΔH_f° for NO_2 ?
- a. $\text{NO}(\text{g}) + \text{O}(\text{g}) \rightarrow \text{NO}_2(\text{g})$
 b. $\text{NO}(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{NO}_2(\text{g})$
 c. $2 \text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{NO}_2(\text{g})$
 d. $\text{N}_2(\text{g}) + 2 \text{O}_2(\text{g}) \rightarrow 2 \text{NO}_2(\text{g})$
e. $\frac{1}{2} \text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{NO}_2(\text{g})$

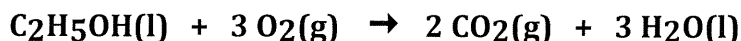
31. Choose the proper ΔH for each equation below. Place the correct ΔH on the provided blanks.

CHOICES: $\Delta H^\circ_{\text{fusion}}$ $\Delta H^\circ_{\text{vap}}$ $\Delta H^\circ_{\text{f}}$ $\Delta H^\circ_{\text{comb}}$



Target 13. I can calculate the enthalpy change in a reaction when given the standard enthalpies of formations of each reactant and product.

32. What is ΔH_f° for one mole of $\text{C}_2\text{H}_5\text{OH}(\text{l})$ given the following data at 25°C :



$$\Delta H_{\text{reaction}}^\circ = -1,366 \text{ kJ}$$

$$\Delta H_f^\circ \text{ for } \text{CO}_2(\text{g}) = -393.5 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{ for } \text{H}_2\text{O}(\text{l}) = -285.8 \text{ kJ/mol}$$

- a. +462.6 kJ b. +278.4 kJ c. -462.6 kJ d. -278.4 kJ

33. What is the heat of combustion for one mole of benzene at 25°C and 1 atm,



given the following data:

$$\Delta H_f^\circ \text{ for } \text{CO}_2(\text{g}) = -393.5 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{ for } \text{H}_2\text{O}(\text{l}) = -285.8 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{ for } \text{C}_6\text{H}_6(\text{l}) = 49.0 \text{ kJ/mol}$$

- a. +3,271 kJ b. -3,271 kJ c. -636 kJ d. +636 kJ

Part II: Answer each of the following questions.

- An experiment is to be performed to determine the standard molar enthalpy of neutralization of a strong acid by a strong base. Standard school laboratory equipment and a supply of standardized 1.00 molar HCl and standard 1.00 molar NaOH are available.
 - What equipment would be needed?
 - What measurements should be taken?
 - Without performing calculations, describe how the resulting data should be used to obtain the standard molar enthalpy of neutralization.
 - When a class of students performed this experiment, the average of the results was -55.0 kilojoules per mole. The accepted value for the standard molar enthalpy of neutralization of a strong acid by a strong base -57.7 kilojoules per mole. Propose two likely sources of experimental error that could account for the result obtained by the class.

2. When iron is oxidized at standard conditions it forms iron (III) oxide.
 - a) Write the reaction representing this reaction.
 - b) What is oxidized and what is reduced?
 - c) Assume that 16.0 grams of iron is oxidized. Calculate the amount of heat (kJ) produced. (ΔH_f° for iron III oxide solid is -824 kJ/mol).

3. How many grams of hot aluminum at a temperature of 180.0°C would have to be placed in 400.0 g of water to raise the temperature of the water from 18.0°C to 25.0°C? (C_p for Al = 0.941 J/g•°C and C_p for H₂O = 4.18 J/g•°C)

4. When a 1.000 gram sample of the rocket fuel hydrazine, N₂H₄, is burned in a bomb calorimeter which contains 1,200. grams of water, the temperature rises from 24.62°C to 28.16°C. If the heat capacity of the dry calorimeter is 840. J/°C, calculate the heat produced by the combustion of the one gram sample. (C_p for water is 4.18 J/g•°C)

CHAPTER 5 PRACTICE TEST WORK

17) $x \text{ kJ} = \left| \frac{200 \text{ g SO}_3}{80 \text{ g SO}_3} \right| \left| \frac{1 \text{ mol SO}_3}{2 \text{ mol SO}_3} \right| \left| \frac{-192.8 \text{ kJ}}{1} \right| =$

18) Calculate the number of moles of Zn & HCl. Use the balanced equation to determine which choice has the most chemicals reacting ... this would create the biggest temperature rise.

(MM of zinc is $\approx 65 \text{ g/mol}$)

	amounts you have		amounts reacting	
	# mol Zn	# mol HCl	# mol Zn	# mol HCl
(a)	0.01	0.02	0.01	0.02
* (b)	0.02	0.04	0.02	0.04 *
(c)	0.02	0.03	0.015	0.03
(d)	0.04	0.03	0.015	0.03
(e)	0.08	0.02	0.01	0.02

* Choice "b" has the most chemicals reacting!

21) $q = (m)(c_p)(\Delta T)$; $540. \text{ J} = (30. \text{ g})(0.90 \text{ J/g}\cdot\text{C})(\Delta T)$; $\Delta T = 20^\circ\text{C}$

$\Delta T = T_f - T_i$; $20^\circ\text{C} = T_f - 25^\circ\text{C}$; $T_f = 45^\circ\text{C}$

22) - q lost by hot H₂O = q gain by cold H₂O

$(-m)(c_p)(\Delta T) = (m)(c_p)(\Delta T)$

$-(100)(T_f - 30^\circ\text{C}) = (50)(T_f - 0)$

$-100 T_f + 3000 = 50 T_f$

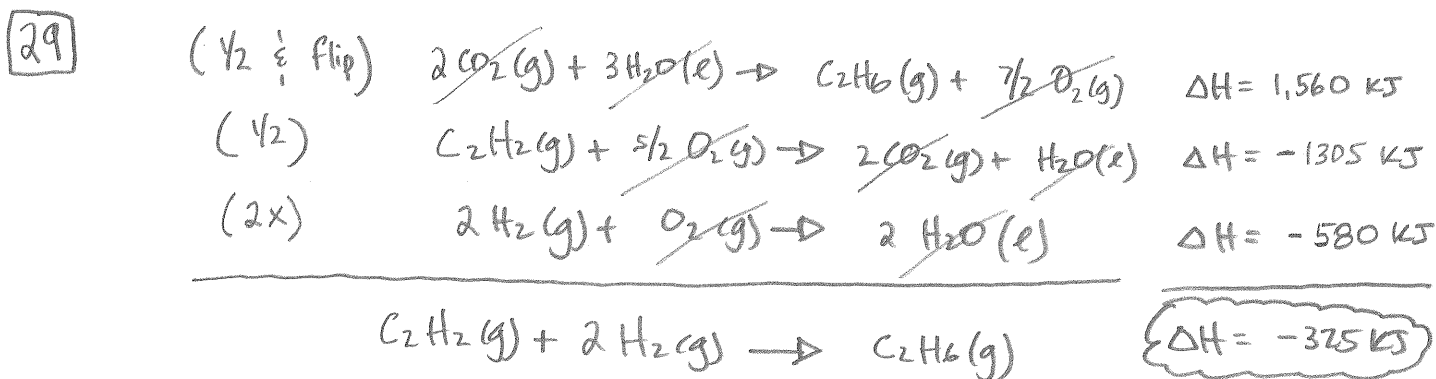
$3000 = 150 T_f$

$T_f = 20^\circ\text{C}$

$$\boxed{23} \quad q = (m)(C_p)(\Delta T) = (10.0 \text{ g})(0.450 \text{ J/g}\cdot\text{C})(50.0^\circ\text{C}) = (4.5)(50) = \boxed{225 \text{ J lost}}$$

This is the ΔT .

$$\boxed{27} \quad X \frac{\text{kJ}}{\text{g}} = \left| \frac{10^\circ\text{C}}{3.0 \text{ g}} \right| \left| \frac{9.0 \text{ kJ}}{1^\circ\text{C}} \right| = \boxed{30. \text{ kJ/g}}$$



$$\boxed{32} \quad \Delta H_{\text{rxn}}^\circ = \sum \Delta H_{\text{products}}^\circ - \sum \Delta H_{\text{reactants}}^\circ$$

$$-1366 \text{ kJ} = \left[(2 \text{ mol})(-393.5 \frac{\text{kJ}}{\text{mol}}) + (3 \text{ mol})(-285.8 \frac{\text{kJ}}{\text{mol}}) \right] - \Delta H_f^\circ \text{C}_2\text{H}_5\text{OH}$$

$$-1366 \approx \left[(2)(-395) + (3)(-285) \right] - X$$

$$-1366 \approx (-790 + -855) - X$$

$$-1366 \approx -1645 - X$$

$$279 = -X$$

$$\boxed{-279 = X \dots \text{choice "D"}}$$

$$\boxed{33} \quad \Delta H^\circ_{\text{rxn}} = \sum \Delta H^\circ_{\text{products}} - \sum \Delta H^\circ_{\text{reactants}}$$

$$\Delta H^\circ_{\text{rxn}} = [12(-393.5) + 6(-285.8)] - [2(49)]$$

$$\Delta H^\circ_{\text{rxn}} \approx [12(-400) + 6(-300)] - 100$$

$$\Delta H^\circ_{\text{rxn}} \approx -4800 + -1800 - 100$$

$$\Delta H^\circ_{\text{rxn}} \approx -6,700 \text{ kJ (this is for 2 moles of } C_6H_6\text{!)}$$

$$\Delta H^\circ \text{ for 1 mole of } C_6H_6 = \frac{-6700}{2} = -3,350 \text{ kJ}$$

choice "B" is closest!

PART 2:

$\boxed{1}$ (a) thermometer, calorimeter, graduated cylinder

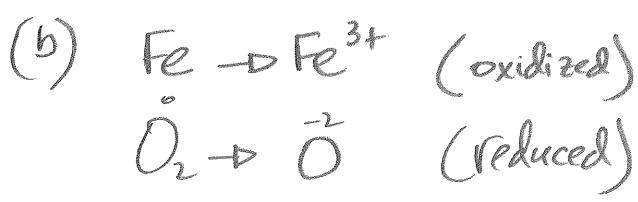
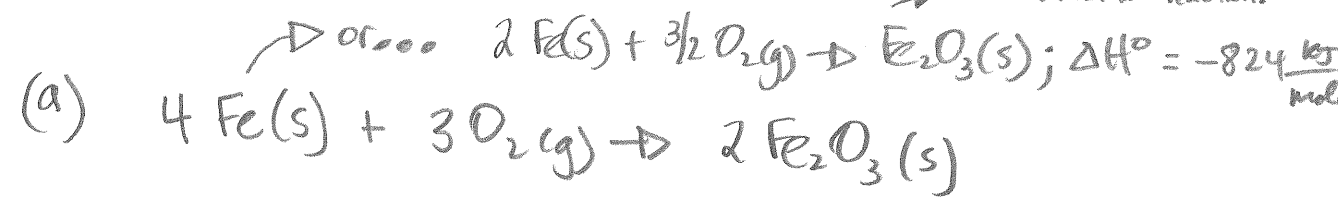
- (b)
- initial temp's of acids & bases
 - final temp. of the solution
 - volumes of the acid & base

(c) $q = (m)(C_p)(\Delta T)$; Multiply the mass of the solution, the specific heat of the solution and the change in temp. of the solution. Divide the product of this by the moles of either the acid or the base.

(d) POSSIBLE ERRORS: 1 - loss of heat to the calorimeter, air, or to the thermometer 2 - incomplete transfer of either the acid or base (not all poured out of grad. cylinder), or 3 - splashing of acid or base onto walls of calorimeter.

2

→ this is a formation reaction.



(c) $X \text{ kJ} = \frac{16.0 \text{ g Fe}}{55.85 \text{ g Fe}} \times \frac{1 \text{ mol Fe}}{2 \text{ mol Fe}} \times \frac{1 \text{ mol Fe}_2\text{O}_3}{1 \text{ mol Fe}_2\text{O}_3} \times \frac{-824 \text{ kJ}}{1 \text{ mol Fe}_2\text{O}_3} = 118 \text{ kJ evolved}$

3

$-q_{\text{lost H}_2\text{O}} = q_{\text{gain by Al}}$
 $-(m)(c_p)(\Delta T) = (m)(c_p)(\Delta T)$
 $-(400.0 \text{ g})(4.18 \text{ J/g}\cdot^\circ\text{C})(25.0^\circ\text{C} - 18.0^\circ\text{C}) = (\text{grams})(0.941 \text{ J/g}\cdot^\circ\text{C})(25^\circ\text{C} - 18.0^\circ\text{C})$
 80.24 grams Al
 $80. \text{ grams of Al}$
 2 sig. fig's

4

$q_{\text{reaction}} = \text{heat absorbed by water} + \text{heat absorbed by calorimeter}$
 $q_{\text{reaction}} = (m)(c_p)(\Delta T) + (C_{\text{cal}})(\Delta T)$
 $q_{\text{reaction}} = (1,200. \text{ g})(4.18 \text{ J/g}\cdot^\circ\text{C})(3.54^\circ\text{C}) + (840. \text{ J/}^\circ\text{C})(3.54^\circ\text{C})$
 $q_{\text{reaction}} = -20,730 \text{ J}$ or $-20,700 \text{ J}$
 3 sig. fig's

NOTE: negative sign indicates heat is RELEASED by the combustion process.

