

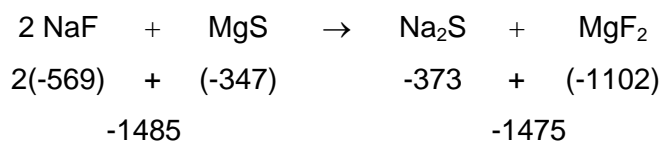
# Unit 1: Thermochemistry Exam Answers

## Multiple Choice

- |     |   |     |   |
|-----|---|-----|---|
| 1.  | A | 11. | A |
| 2.  | D | 12. | C |
| 3.  | B | 13. | B |
| 4.  | C | 14. | D |
| 5.  | C | 15. | B |
| 6.  | A | 16. | C |
| 7.  | B | 17. | D |
| 8.  | C | 18. | B |
| 9.  | B | 19. | D |
| 10. | D | 20. | C |

## Short Answer

1. a.

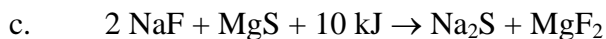


$$\Delta H = \Sigma \Delta H_{\text{products}} - \Sigma \Delta H_{\text{reactants}}$$

$$\Delta H = -1475 - (-1485)$$

$$\Delta H = +10 \text{ kJ} \quad \text{ANSWER}$$

b. endothermic (because  $\Delta H$  is positive)



2. a.  $\text{Fe} + 2 \text{ S} \rightarrow \text{FeS}_2 + 178.8 \text{ kJ}$  (note: Sulfur is NOT diatomic)

b.  $\frac{1}{2} \text{ H}_2 + \frac{1}{2} \text{ F}_2 \rightarrow \text{HF} + 271.1 \text{ kJ}$  (note: both hydrogen & fluorine are diatomic)

c.  $\text{K} + \frac{1}{2} \text{ O}_2 + \frac{1}{2} \text{ H}_2 \rightarrow \text{KOH} + 428.8 \text{ kJ}$

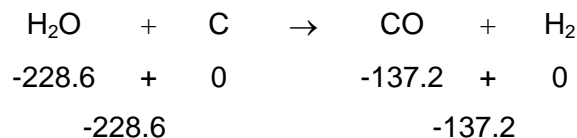
3.  $Q = m c \Delta T$

$$Q = (96.7) (0.874) (37.5) = \mathbf{3169 \text{ J or } 3.169 \text{ kJ}}$$

(to significant figures the answer should be 3.17 kJ)

4. a.  $\text{H}_2\text{O} + \text{C} \rightarrow \text{CO} + \text{H}_2$   $\Delta\text{H} = +134 \text{ kJ}$
- b. endothermic
- c. see separate page
- d. not spontaneous because the reaction is **endothermic**
- e. predict a spontaneous reaction because entropy increases – a solid (carbon) changes into a gas (CO), which is a more random state.

f.  $\Delta\text{G} = \Sigma \Delta\text{G}_{\text{products}} - \Sigma \Delta\text{G}_{\text{reactants}}$



$$\Delta\text{G} = \Sigma \Delta\text{G}_{\text{products}} - \Sigma \Delta\text{G}_{\text{reactants}}$$

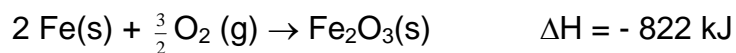
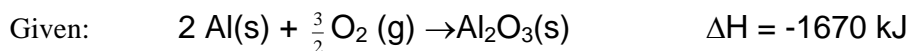
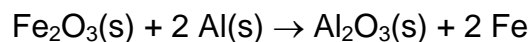
$$\Delta\text{G} = -137.2 - (-228.6)$$

$$\Delta\text{G} = +91.4 \text{ kJ} \quad \text{ANSWER}$$

- g. Because  $\Delta\text{G}$  is positive we know that the reaction will NOT be spontaneous.

5. a. May solve using either of two methods:

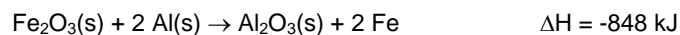
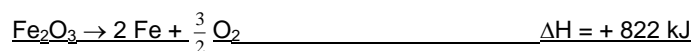
Find  $\Delta\text{H}$  for:



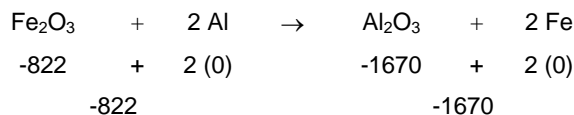
Method 1: Rearrange Equations (there was an error with a value of  $\Delta\text{H}$  given in the question; it has been corrected here)

You will need to reverse the bottom equation in order to get  $\text{Fe}_2\text{O}_3$  on the reactant side. This will change the value of  $\Delta\text{H}$  from negative to positive (the error on the test gave  $\Delta\text{H}$  as  $+822 \text{ kJ}$ )

Once you've reversed the bottom equation, add the two equations together to get the desired equation given by the question. Then add the two  $\Delta\text{H}$  values to get  $\Delta\text{H}$  for the desired equation:



Method 2: Use  $\Delta H = \Sigma \Delta H_{\text{products}} - \Sigma \Delta H_{\text{reactants}}$



$$\Delta H = \Sigma \Delta H_{\text{products}} - \Sigma \Delta H_{\text{reactants}}$$

$$\Delta H = -1670 - (-822)$$

$$\Delta H = -848 \text{ kJ} \quad \text{ANSWER}$$

- b. How much heat will be released when 1000.0 g of iron is produced?

This can be set up in a number of different ways. The following way uses unit analysis – all units will cancel out except for the desired unit of kJ, the unit for heat.

You will need to use the molar mass of iron, Fe, which is  $\frac{55.8 \text{ g}}{\text{mol}}$

Also, note from the balanced equation given in the question and your calculation in part a that 848 kJ are released for every two moles of iron ( $\frac{848 \text{ kJ}}{2 \text{ mol Fe}}$ )

$$\text{Set up and solve: } \text{kJ} = \frac{848 \text{ kJ}}{2 \text{ mol}} \times \frac{1 \text{ mol}}{55.8 \text{ g}} \times \frac{1,000 \text{ g}}{1} = 7599 \text{ kJ} \quad \text{Answer}$$

6. a. Calculate  $\Delta G$  for:  $2 \text{ Al} + 3 \text{ Cl}_2 \rightarrow 2 \text{ AlCl}_3$

$$\begin{array}{l} \text{Also given: } \quad \Delta H = -704 \text{ kJ} \\ \quad \quad \quad \Delta S = 1,100 \text{ kJ/K} \\ \quad \quad \quad T = 200^\circ\text{C} \end{array}$$

$$\text{To solve, you must use the equation: } \quad \Delta G = \Delta H - T(\Delta S)$$

You must first remember to convert temperature from  $^\circ\text{C}$  into K, and  $\Delta S$  from J/K into kJ/K:

$$K = 200^\circ\text{C} + 273 = \mathbf{473 \text{ K}} \quad \Delta S = \mathbf{1.110 \text{ kJ/K}}$$

$$\begin{array}{l} \text{Solve: } \quad \Delta G = \Delta H - T(\Delta S) \quad \Delta G = -704 - 473(1.110) \\ \quad \quad \quad \quad \quad \quad \quad \quad = -704 - 525.03 \\ \quad \quad \quad \quad \quad \quad \quad \quad = -1229 \text{ kJ} \quad \quad \quad \text{ANSWER} \end{array}$$

- b. Is the reaction spontaneous at this temperature? How do you know?

The negative value for  $\Delta G$  tells us that the reaction WILL BE spontaneous.

Question 4c.

