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.

2 NaF	+	MgS	$\rightarrow$	Na_2S	+	MgF_2
2(-569)	+	(-347)		-373	+	(-1102)
	-1485				-1475	

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

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

$\Delta H =$ +10 kJ ANSWER

b. endothermic (because ΔH is positive)

c. $2\ NaF + MgS + 10\ kJ \rightarrow Na_2S + MgF_2$

2. a. $Fe + 2\ S \rightarrow FeS_2 + 178.8\ kJ$ (note: Sulfur is NOT diatomic)

b. $\frac{1}{2}\ H_2 + \frac{1}{2}\ F_2 \rightarrow HF + 271.1\ kJ$ (note: both hydrogen & fluorine are diatomic)

c. $K + \frac{1}{2}\ O_2 + \frac{1}{2}\ H_2 \rightarrow KOH + 428.8\ kJ$

3. $Q = m\ c\ \Delta\ T$

$Q = (96.7)\ (0.874)\ (37.5) =$ **3169 J or 3.169 kJ**

(to significant figures the answer should be 3.17 kJ)

4. a. $H_2O + C \rightarrow CO + H_2$ $\Delta H = +134$ 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. $\Delta G = \Sigma\ \Delta G_{products} - \Sigma\ \Delta G_{reactants}$

H_2O	+	C	$\rightarrow$	CO	+	H_2
-228.6	+	0		-137.2	+	0
	-228.6				-137.2	

$\Delta G = \Sigma\ \Delta G_{products} - \Sigma\ \Delta G_{reactants}$

$\Delta G =$ -137.2 - (-228.6)

$\Delta G =$ +91.4 kJ ANSWER

g. Because ΔG is positive we know that the reaction will NOT be spontaneous.

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

Find ΔH for:

$$Fe_2O_3(s) + 2\ Al(s) \rightarrow Al_2O_3(s) + 2\ Fe$$

Given: $2\ Al(s) + \frac{3}{2} O_2\ (g) \rightarrow Al_2O_3(s)$ $\Delta H = -1670$ kJ

$2\ Fe(s) + \frac{3}{2} O_2\ (g) \rightarrow Fe_2O_3(s)$ $\Delta H = -\ 822$ kJ

Method 1: Rearrange Equations (there was an error with a value of ΔH given in the question; it has been corrected here)

You will need to reverse the bottom equation in order to get Fe_2O_3 on the reactant side. This will change the value of ΔH from negative to positive (the error on the test gave ΔH as +822 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 ΔH values to get ΔH for the desired equation:

$2\ Al + \frac{3}{2} O_2 \rightarrow Al_2O_3$ $\Delta H = -1670$ kJ

$\underline{Fe_2O_3 \rightarrow 2\ Fe + \frac{3}{2} O_2}$ $\underline{\Delta H = +822 \text{ kJ}}$

$Fe_2O_3(s) + 2\ Al(s) \rightarrow Al_2O_3(s) + 2\ Fe$ $\Delta H = -848$ kJ

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

Fe_2O_3	+	2 Al	$\rightarrow$	Al_2O_3	+	2 Fe
-822	+	2 (0)		-1670	+	2 (0)
	-822				-1670	

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

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

$\Delta H =$ -848 kJ 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 $\frac{55.8g}{mol}$

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 ($\frac{848kJ}{2mol\_Fe}$)

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

6. a. Calculate ΔG for: $2Al + 3\ Cl_2 \rightarrow 2\ AlCl_3$

Also given:
ΔH = -704 kJ
ΔS = 1,100 kJ/K
T = 200°C

To solve, you must use the equation: ΔG = ΔH – T(ΔS)

You must first remember to convert temperature from °C into K, and ΔS from J/K into kJ/K:

K = 200°C + 273 = **473 K** **ΔS = 1.110 kJ/K**

Solve: ΔG = ΔH – T(ΔS) ΔG = -704 – 473 (1.110)
= -704 – 525.03
= -1229 kJ ANSWER

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

The negative value for ΔG tells us that the reaction WILL BE spontaneous.

Question 4c.