Name ___________________________________**Chem 161, Section**: _______**Group Number:** ______

ALE 27. Hess's Law

(Reference: Chapter 6 - *Silberberg 5th edition*)

> **Important!!** For answers that involve a calculation you must show your work neatly using dimensional analysis with correct significant figures and units to receive full credit. No work, no credit. Report numerical answers to the correct number of significant figures. **CIRCLE ALL NUMERICAL RESPONSES**.

What are thermochemical equations and why can they be added algebraically?

The Model: Heats of Reaction

H is the symbol for **enthalpy** (*i.e.*, "heat content") of a chemical substance. When ΔH is the difference in enthalpies between products and reactants ($\boldsymbol{\Delta H = H_{products} - H_{reactants}}$), ΔH is called the "heat of reaction".

Atomization reactions:

written as a **Thermochemical Equation**	written with the heat of reaction separate and to the side	
$C(s, \text{graphite}) + 715\ kJ \rightarrow C(g)$	$C(s, \text{graphite}) \rightarrow C(g)$	$\Delta H = 715\ kJ$
$H_2(g) + 436\ kJ \rightarrow 2\ H(g)$	$H_2(g) \rightarrow 2\ H(g)$	$\Delta H = 436\ kJ$
$CH_4(g) + 1662\ kJ \rightarrow C(g) + 4\ H(g)$	$CH_4(g) \rightarrow C(g) + 4\ H(g)$	$\Delta H = 1662\ kJ$

Key Questions

1. What is a thermochemical equation?

2. What is an atomization reaction?

3. When a reaction is endothermic (circle the correct answer):
   i. $\Delta H < 0$ and heat is a reactant
   ii. $\Delta H > 0$ and heat is a reactant
   iii. $\Delta H < 0$ and heat is a product
   iv. $\Delta H > 0$ and heat is a product

4. The symbol "kJ" stands for heat just as "$CH_4(g)$" stands for gaseous methane. Thermochemical equations are best interpreted on a "macroscopic" scale. Complete the macroscopic interpretation of the following thermochemical equation.

$$CH_4(g) + 1662\ kJ \rightarrow C(g) + 4\ H(g)$$

One mole of gaseous methane reacts with ____

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5. When 1 mole of gaseous carbon is allowed to combine with 4 moles of gaseous atomic hydrogen, 1 mole of gaseous methane is produced and (circle the correct answer)::

   i. 1662 kJ of heat is consumed   ii. 1662 kJ of heat is liberated

6. When 2 moles of gaseous molecular hydrogen are atomized, 4 moles of gaseous atomic hydrogen are produced and (circle the correct answer)::

   i. 218 kJ of heat are consumed   iv. 218 kJ of heat are liberated

   ii. 436 kJ of heat are consumed   v. 436 kJ of heat are liberated

   iii. 872 kJ of heat are consumed   vi. 872 kJ of heat are liberated

The Model: Hess's Law

Just like algebraic equations can be added together, chemical equations can be added together. A species that is on the reactant side of one reaction algebraically cancels with the <u>*same species in the same physical state*</u> on the product side of another reaction.

<u>A set of chemical equations that we will add together in an attempt to get **equation 6**, below</u>:

eqn 1 $C(s, \text{graphite}) \rightarrow C(g)$ $\Delta H = 715$ kJ

eqn 2 $H_2(g) \rightarrow 2\,H(g)$ $\Delta H = 436$ kJ

eqn 3 $CH_4(g) \rightarrow C(g) + 4\,H(g)$ $\Delta H = 1662$ kJ

We'll start out with **eqn 1** "as is". But we'll have to get rid of the 1 mole of gaseous carbon atoms on the product side. **Equation 3** has 1 mole of gaseous carbon as a product. So we'll reverse **eqn 3** (*i.e.*, we'll make the reactants products, and the products reactants) and label the reverse of **eqn 3** as eqn **4**:

eqn 4 $C(g) + 4\,H(g) \rightarrow CH_4(g)$ $\Delta H = -1662$ kJ

Now if we add eq 1 and eq 4 we'll be able to cancel the gaseous carbon atoms, but we'll still have 4 moles of gaseous atomic hydrogen on the reactant side. If we define eq **5** as two times eq **2**,

eqn 5 $2\,H_2(g) \rightarrow 4\,H(g)$ $\Delta H = 872$ kJ

we have an equation which we can add to eq **1** and eq **4** to eliminate the gaseous atomic hydrogen.

Hess's Law says that when we add chemical equations together, the heat of the resultant reaction is a sum of the heats of the individual reactions. So . . .

eqn 1 $C(s, \text{graphite}) \rightarrow \cancel{C(g)}$ $\Delta H = 715$ kJ

eqn 4 $\cancel{C(g)} + \cancel{4\,H(g)} \rightarrow CH_4(g)$ $\Delta H = -1662$ kJ

eqn 5 $2\,H_2(g) \rightarrow \cancel{4\,H(g)}$ $\Delta H = 872$ kJ

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eqn 6 $C(s, \text{graphite}) + 2\,H_2(g) \rightarrow CH_4(g)$ $\boldsymbol{\Delta H} = (715 - 1662 + 872)$ kJ
= **-75 kJ**

Equation 6, above, is an example of a ***formation reaction*** because it shows the production of 1 mole of a compound of interest (*e.g.*, gaseous methane) from elements in their naturally-occurring states (*e.g.*, solid carbon graphite and gaseous molecular hydrogen).

Key Questions

7. What is a formation reaction?

8. What is the heat of formation of *any* element in its naturally-occurring state? *Explain why.*

Exercise

A. A **combustion reaction** is one in which a substance reacts with gaseous molecular oxygen. The **complete** combustion reactions of graphite, hydrogen, and methane are:

$$C(s, \text{graphite}) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H = -394 \text{ kJ}$$

$$H_2(g) + \tfrac{1}{2} O_2(g) \rightarrow H_2O(l) \qquad \Delta H = -286 \text{ kJ}$$

$$CH_4(g) + 2\, O_2(g) \rightarrow CO_2(g) + 2\, H_2O(l) \qquad \Delta H = -891 \text{ kJ}$$

Use the above combustion reactions and Hess's Law to determine the heat of the following reaction:

$$C(s, \text{graphite}) + 2\, H_2(g) \rightarrow CH_4(g) \qquad \Delta H = ?$$

Key Question

9. Compare the heat of the formation reaction of methane you determined in Exercise A (*i.e.*, when combustion reactions were added together) with that which was presented in the Model (*i.e.*, when atomization reaction were added together). Explain the relevance of this discovery.

Exercise

B.① Write the balanced *formation reaction* for $Fe_3O_4(s)$. *Hint*: What are the naturally-occurring states of iron and oxygen?

② Calculate the heat of formation (ΔH_f) for $Fe_3O_4(s)$ given the following "bank" of thermo-chemical equations:

$$Fe_2O_3(s) + 3\ CO(g) \rightarrow 2\ Fe(s) + 3\ CO_2(g) \qquad \Delta H = -28\ kJ$$

$$3\ Fe_2O_3(s) + CO(g) \rightarrow 2\ Fe_3O_4(s) + CO_2(g) \qquad \Delta H = -59\ kJ$$

$$2\ CO(g) + O_2(g) \rightarrow 2\ CO_2(g) \qquad \Delta H = -568\ kJ$$

C. Would you expect $O_{2(g)} \rightarrow 2\ O_{(g)}$ to have a positive or negative ΔH_{rxn}? Explain your reasoning.

D. Consider the following balanced thermochemical equation for the decomposition of the mineral magnesite:

$$MgCO_{3(g)} \rightarrow MgO_{(s)} + CO_{2(g)} \qquad \Delta H_{rxn} = 117.3\ kJ$$

a.) Is heat absorbed or released in the reaction? Explain your reasoning.

b.) What is ΔH_{rxn} of the reverse reaction?

c.) What is ΔH when 5.35 mol of CO_2 reacts with excess MgO? *Show work and circle your answer*.

d.) What is ΔH when 35.5 g of CO_2 reacts with excess MgO? *Show work and circle your answer*.

E. When 1 mol of solid potassium bromide, KBr, decomposes to its elements, 394 kJ of heat is absorbed.

a.) Write the balanced thermochemical equation for this reaction.

b.) How much heat is released when 10.0 kg of KBr forms from its elements? *Show work and circle your answer*.

F. Sucrose (table sugar, $C_{12}H_{22}O_{11}$) is oxidized in the body by O_2 via a complex set of reactions (glycolysis, Krebs cycle and electron transport chain) that ultimately produces $CO_{2(g)}$ and $H_2O_{(l)}$ and releases 5.64 x 10^3 kJ/mol sucrose.

a.) Write the balanced thermochemical equation for this reaction.

b.) How much heat is released per gram of sucrose oxidized? *Show work and circle your answer*.

G. Given the thermochemical reactions 1 and 2, below, calculate ΔH_{rxn} for reaction #3. *Show work and circle your answer*.

Reaction #1:	$\frac{1}{2} N_{2(g)} + \frac{1}{2} O_{2(g)} \rightarrow NO_{(g)}$	ΔH_{rxn} = 90.3 kJ
Reaction #2:	$NO_{(g)} + \frac{1}{2} Cl_{2(g)} \rightarrow NOCl_{(g)}$	ΔH_{rxn} = -38.6 kJ
Reaction #3:	$2\ NOCl_{(g)} \rightarrow N_{2(g)} + O_{2(g)} + Cl_{2(g)}$	ΔH_{rxn} = ???