Thermochemical Equations

Introduction

Chemical reactions always involve some sort of energy change. These energy changes often involve either heat gain or heat loss. A little terminology might help here. The *system* refers to the substance being observed or studied, such as the chemicals involved in a reaction. The *surroundings* are everything outside of the system. *Heat* is often described as the energy that flows into or out of a system as a result of a temperature difference. Heat always flows from the "hotter" substance, that being the substance with a higher temperature, to the substance with a lower temperature - the "colder" one. In an *exothermic* reaction, heat flows out of the system into the surroundings. It could be said the system has lost some heat, and the surroundings have gained some. In an *endothermic* reaction, the system gains energy, or heat, from the surroundings. A flame is a good example of an exothermic reaction. It feels hot to the touch, because the system - the reaction involved in the combustion - is releasing heat to the environment, made up of your hand, among other things. Oppositely, an instant cold pack represents an endothermic change; it feels cold because heat is leaving your hand and flowing into the system - the cold pack.

Units

As heat is a form of energy, it is measured in SI units in joules (J), or, kilojoules (kJ), where 1000 J = 1 kJ. There are other units used for heat as well, such as calories or BTU's, British Thermal Units. We will consistently refer to heat using SI units.

Thermochemical Equations

Since exothermic reactions involve the loss of heat to the surroundings, an energy term may be included on the product side of a chemical equation. For example, suppose octane (C_8H_{18}) burns in the presence of oxygen, producing carbon dioxide and water. The balanced equation for this reaction is:

$$2\ C_8H_{18(l)} + 25\ O_{2(g)} \rightarrow 16\ CO_{2(g)} + 18\ H_2O_{(l)}$$

When two moles of octane burns according to this reaction, 10,148kJ of heat are *produced*. Thus, the corresponding *thermochemical equation* may be written:

$$2\ C_8H_{18(l)} + 25\ O_{2(g)} \rightarrow 16\ CO_{2(g)} + 18\ H_2O_{(l)} + 10{,}148\text{ kJ}$$

Similarly, for an endothermic reaction, an energy term may be included on the reactant side of a chemical equation, indicating energy is being added into the system.

Why Use Thermochemical Equations ?

Thermochemical equations are very valuable tools in chemistry, since they directly relate moles of reactants and products to the heat generated or absorbed by the reaction. For example, consider the above reaction, the combustion of octane. How many kilojoules of heat are released into the surroundings for each gram of octane that reacts? We can easily calculate the molar mass of octane:

C: 8 x 12.01g = 96.08g
H: 18 x 1.01g = 18.18g
114.26g

And using this information, together with the balanced equation, we can calculate:

$$1.000\text{g } C_8H_{18} \times \frac{1\text{ mole}}{114.26\text{g}} \times \frac{10{,}148\text{ kJ}}{2\text{ moles } C_8H_{18}} = 44.41\text{ kJ}$$

Thus, 44.41 kJ of heat is released into the surroundings for each gram of octane that is consumed in the reaction. Or consider the following question: how many grams of carbon dioxide have been produced according to the above reaction if 5000 kJ of heat has been generated?

C: 1 x 12.10g = 14.01g
O: 2 x 16.00g = 32.00g
44.01g

$$5000\text{ kJ} \times \frac{16\text{ mole } CO_2}{10{,}148\text{ kJ}} \times \frac{44.01\text{g}}{1\text{ mole } CO_2} = 346.95\text{g } CO_2$$

Practice with these....

1. Consider the following thermochemical reaction for kerosene:

$$2\ C_{12}H_{26(l)} \quad + \quad 37\ O_{2(g)} \quad \rightarrow \quad 24\ CO_{2(g)} \ + \ 26\ H_2O_{(l)} \quad + \quad 15{,}026\ \text{kJ}$$

   (a) When 21.3 g of CO_2 are made, how much heat is released?

   (b) If 500.00 kJ of heat are released by the reaction, how grams of $C_{12}H_{26}$ must have been consumed ?

   (c) If this reaction were being used to generate heat, how many grams of $C_{12}H_{26}$ would have to be reacted to generate enough heat to raise the temperature of 750g of liquid water from $10^{o}C$ to $90^{o}C$?

2. Consider the reaction: $NaNO_3(s) + H_2SO_4(l) \rightarrow NaHSO_4(s) + HNO_3(g)$ $\Delta H° = 21.2$ kJ

   How much heat must absorbed by the reaction system to convert 100g of $NaNO_3$ into $NaHSO_4(s)$?

3. What is the enthalpy change when 49.4 mL of 0.430 *M* sulfuric acid reacts with 23.3 mL of 0.309 *M* potassium hydroxide?

$$H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l) \qquad \Delta H° = -111.6\ \text{kJ/mol}$$