Chapter 12: Heat in Chemical Reactions

12-1 Chemical Reactions that Involve Heat

- Chemical reactions involve breaking and/or making bonds and rearranging atoms.
- Breaking bonds requires energy and making bonds releases energy.
- Heat- the energy that is transferred from one object to another due to a difference in temperature
- Temperature – average kinetic energy of a substance

Thermochemistry – study of the changes in heat in a chemical reaction.

- Exothermic reactions – reactions that RELEASE heat
- Endothermic reactions – reactions that ABSORB heat
- Surroundings vs. system

Exothermic reactions:

- Burning of a camp stove (propane C_3H_8)

  $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O + 2043\ kJ$
- 2043 kJ of heat are released when 1 mole of C_3H_8 is burned.
- The energy ***released*** during forming new bonds is **greater** than the energy ***required*** to break the old bonds…end result is RELEASE of energy…**EXOTHERMIC**

Endothermic Reactions:

- A process in fuel industry called water gas, when steam (H_2O) is passed over hot coals (C)

$$C + H_2O + 113KJ \rightarrow CO + H_2$$

- The energy ***released*** as new bonds are formed in the products is **less than** the energy ***required*** to break the bonds in the reactants. This energy must be provided for the reaction to take place and is stored in the bonds of the products…end result is ABSORBING of energy…**ENDOTHERMIC**

12-2 Heat and Enthalpy Changes

- Enthalpy – the heat absorbed or gained during a chemical reaction.

- The difference between energy and enthalpy is very subtle. When the pressure remains constant, the **energy** absorbed or released during a chemical reaction **is equal** to the **enthalpy** change for the reaction.

Interpreting data

- Temperature change (ΔT)

$$\Delta T = T_{final} - T_{initial}$$

- Enthalpy change ΔH

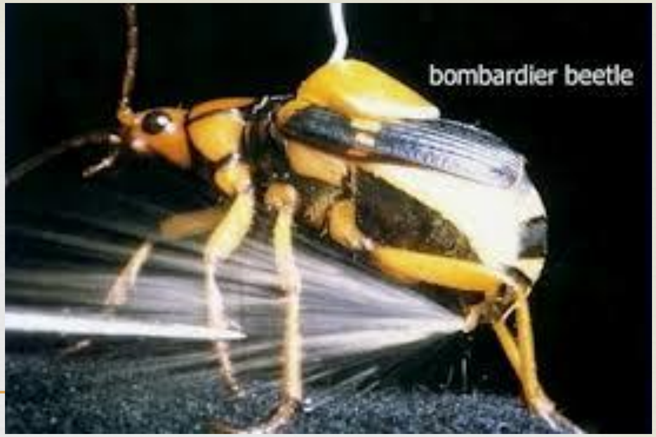
$$\Delta H = H_{products} - H_{reactants}$$

(not used to calculate though)

Sign ΔH	**Process**	**Heat**
+	Endo	Absorbed
-	Exo	Released

Sample problem…

Bombadier beetle →



- How much heat will be released if 1.0g of H_2O_2 decomposes in a beetle to produce the steam spray (see picture above)?

$$\underline{2H_2O_2} \rightarrow 2H_2O + O_2 \qquad \underline{\Delta H = -190KJ}$$

…hint – think stoich with heat ΔH as part of the ratio

- $1.0g\ H_2O_2 \times \frac{1mol\ H_2O_2}{34\ g} \times \frac{-190kJ}{2\ mol\ H_2O_2} = -2.8KJ$

12-3 Hess's Law

- Hess's Law- if a series of reactions are added together, the enthalpy change for the net reaction will be the sum of the enthalpy changes for the individual steps.

- Allows you to find enthalpy changes of reactions that cannot be measured directly

Consider the haze in a large city

$$N_2 + 2O_2 \rightarrow \underline{2NO_2}$$ (net reaction)

$N_2 + O_2 \rightarrow 2NO$ $\Delta H = +181\ kJ$ (equation 1)

$2NO + O_2 \rightarrow 2NO_2$ $\Delta H_2 = -113\ kJ$ (equation 2)

$$N_2 + 2O_2 + \cancel{2NO} \rightarrow \cancel{2NO} + 2NO_2$$ (net reaction)

ΔH net = ΔH equation 1 + ΔH equation 2

$\Delta H = 181 + -113 = 68\ kJ$

Rules for applying Hess's Law

- If you multiply or divide the coefficients by a number do the same to ΔH.

- If the equation is reversed, so is the sign of ΔH.

Let's try one…

- Calculate ΔH for the reaction that produces SO_2…

$$S + O_2 \rightarrow SO_2$$

- $2SO_2 + O_2 \rightarrow 2SO_3$ ΔH = -196KJ (equation 1)
- $2S + 3O_2 \rightarrow 2SO_3$ ΔH = -790 KJ (equation 2)

- 1st look at the products and reactants in the net equation and see if they are on the same side in one of the numbered equations…if not FLIP the equations to get them on the correct side (don't forget that the ΔH must change accordingly)
  - $2SO_3 \rightarrow 2SO_2 + O_2$ $\Delta H = 196$ kJ (equation 1)
- 2nd look at the coefficients for the net equation and see if they will match the numbered equations…if not X or / all by the correct number to get the coefficients you need (don't forget that the ΔH must change accordingly)
  - $SO_3 \rightarrow SO_2 + \frac{1}{2} O_2$ $\Delta H = 98$ kJ (equation 1)
  - $S + 3/2\ O_2 \rightarrow SO_3$ $\Delta H = -395$ kJ (equation 2)

- $SO_3 \rightarrow SO_2 + \frac{1}{2} O_2$ $\Delta H = 98$ kJ (equation 1)
- $\underline{S + 3/2\, O_2 \rightarrow SO_3}$ $\Delta H = -395$ kJ (equation 2)

$$\cancel{SO_3} + S + \cancel{3/2}\, O_2 \longrightarrow SO_2 + \cancel{\tfrac{1}{2} O_2} + \cancel{SO_3}$$

$$S + O_2 \rightarrow SO_2 \text{ (net reaction)}$$

$$\Delta H = 98 \text{ kJ} + -395 \text{ kJ}$$

$$\Delta H = -297 \text{ kJ}$$

You try it now…

- WS 12-3 PP try # 2
  - Answer: -123 kJ
- HW…do 2 – 4

12-4 Calorimetry

- Calorimetry- the indirect study of heat flow and heat measurement

- Heat capacity – the amount of heat needed to raise the temperature of the object by 1 Celsius degree.
  - For example, the heat capacity of a cup of water at 18^{o} C is the number of joules needed to make it 19.
  - Depends on mass and composition

- Specific heat capacity (c) – the heat capacity of 1 gram of a substance

$$c_{water} = 4.184\ J/g\ ^{o}C$$

- To raise the temp of 1 gram of water 1 º C you need to add 4.184 J of energy

Calorimetry Equations

$$Q = mc\Delta T$$

$$Q_{surroundings} = -Q_{system}$$

- Q = quantity of energy transferred (heat)
- m = mass of substance that gains/loses energy
- ΔT = change in temperature of substance that gains/loses energy
- c = specific heat capacity of substance that gains/loses energy

Let's try one…

- Determine ΔH for the addition of NaOH to water when the calorimeter is filled with 75.0g water. The initial temp is 19.8 ºC. A 2.0 g sample of solid NaOH is added and the temp increases to 26.7 ºC.
  - 4 steps:
    - Calculate $Q_{\text{water (surroundings)}}$
    - Determine Q_{system}
    - Convert g solid added to moles
    - $\Delta H = Q_{\text{system}}$/moles solid added

The work…

- Step 1:

  $Q_{surroundings} = mc\Delta T$

  $Q_{surroundings} = (75.0g)(4.184J/g^oC)(26.7 - 19.8\ ^oC)$

  $Q_{surroundings} = +2165\ J$

- Step 2:

  $Q_{surroundings} = -Q_{system}$…so $Q_{system} =$ -2165 J

- Step 3:

  $2.0\ g\ NaOH \times \frac{1\ mol\ NaOH}{40.0\ g\ NaOH} = 0.050\ mol\ NaOH$

- $\Delta H = \frac{-2165\ J}{0.050\ mole\ NaOH} = -43,304\ J$

- When a 4.25 g sample of solid NH_4OH dissolves in 60.0 g of water, the temperature drops from 21.0 ºC to 16.9 ºC solve ΔH.

- 1st calculate $Q_{surroundings}$ $Q_{surroundings} = mc\Delta T$

$Q_{surroundings} = (60.0g)(4.184 J/g^oC)(16.9\ ^oC - 21.0\ ^oC)$

$Q_{surroundings} = -1029\ J$

- 2nd $Q_{system} = -Q_{surroundings} = +1029\ J$
- 3rd $4.25g\ NH_4OH \times \frac{1\ mole\ NH_4OH}{35\ g\ NH_4OH} = 0.121 m NH_4OH$
- 4th $\Delta H = \frac{1029\ J}{0.121 m NH_4OH} = +8504\ J$

Calculating c…(specific heat capacity)

- We can use $Q=mc\Delta T$ to solve for other variables in the equation…specifically c

$$c = \frac{Q}{m\,\Delta T}$$

- Ex. What is the specific heat of a piece of silver if 30.8 g sample increases 11.2°C when 81 J of heat are added?

$$c = \frac{81\ \text{J}}{(30.8\ \text{g})(11.2\ {}^{o}\text{C})} = 0.23\ \text{J/g}\ {}^{o}\text{C}$$

You try one now…

- WS 12-4 PP WS do number 2
- Answer: 0.90 J/g°C