ENTHALPY CHANGE

- **Thermochemistry** is the study of energy changes (energy produced or absorbed) by a chemical system during a chemical reaction
  - Energy is always conserved, just converted from one form to another

- To study energy changes, an **isolated system** is required (neither matter nor energy can enter or leave)
  - However it is impossible to create an isolated system
  - Usually takes place in a **closed system** (matter cannot enter of leave, but energy can enter or leave)

- Chemical systems have different forms of energy
  - **Kinetic energy:** energy of motion from moving electrons and also moving/vibrating atoms "heat energy"
  - **Potential energy:** stored energy in the form of intermolecular bonds (between molecules) and intramolecular bonds (between atoms within the molecule) "bond energy"

- The total kinetic and potential energy of a chemical system is called its **enthalpy (*H*)** => measured in joules (J or kJ)
  - If the reactants and products are at the same initial and final temperatures, then the kinetic energy does not change. This is usually the case and in this course we will assume enthalpy is potential energy only. *
- Enthalpy is usually communicated as a difference, therefore **enthalpy change (ΔH)** communicates the difference between the enthalpy of the products and the enthalpy of the reactants
  - Also referred to as the **net energy** for a reaction

$$\Delta H = H_{products} - H_{reactants} \qquad \text{or} \qquad \Delta H = H_p - H_r$$

- If a chemical reaction *produces or releases* energy to the surroundings, ΔH will be negative because the chemical system is losing energy. This type of reaction is called an **exothermic reaction**.
- If a chemical reaction *gains or absorbs* energy from the surroundings (ΔH will be positive because the chemical system is gaining energy. This type of reaction is called an **endothermic reaction**.

*****Activity: Making an Ice Pack*****

- * All reactions involve both the breaking and forming of bonds, but exothermic and endothermic reactions can be explained in terms of bond energy
  - When bonds break, energy is required
  - When bonds form, energy is released } also defined as bond energy

Type of reaction	Breaking chemical bonds (reactants)	Forming chemical bonds (products)	Overall energy change
exothermic	Energy Absorbed small	ENERGY RELEASED large	energy released ΔH = released (large) − absorbed (small) ΔH = released
endothermic	ENERGY ABSORBED large	Energy Released small	energy absorbed ΔH = released (small) − absorbed (large) ΔH = absorbed

- All chemical reactions/systems are accompanied by a change in energy. A common, natural occurring example is the photosynthesis and cellular respiration reactions (↳ combustion rxn)
  - o All combustion reactions in an open system will produce gaseous water
  - o All combustion reactions in a closed system will produced liquid water

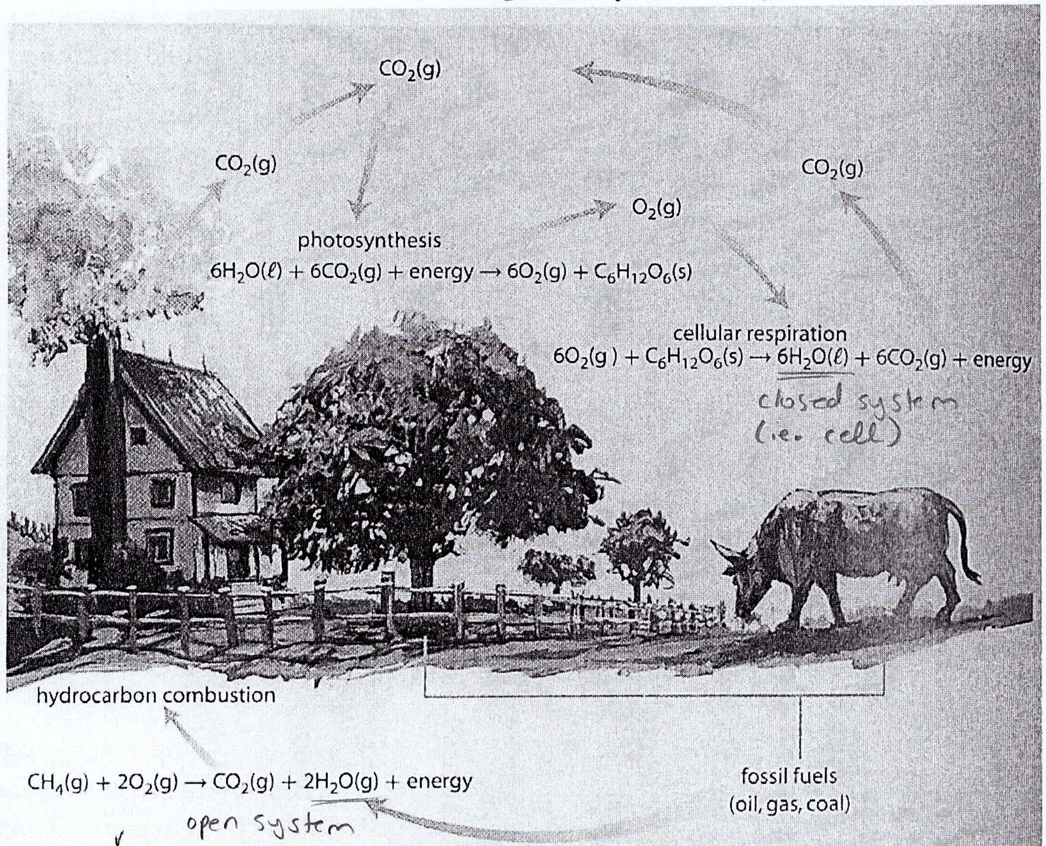


Figure 9.1 Sunlight is the original source of energy for the formation of fossil fuels.

*need to memorize photosynthesis & cellular respiration rxns!

- Sunlight is the original source of energy for many biological chemical reactions and ultimately the energy source for the formation of fossil fuels (specific outcome!)

*****Gummy Bear Demo*****

COMMUNICATING ENTHALPY CHANGE FOR AN EXOTHERMIC & ENDOTHERMIC REACTIONS

1. **Thermochemical Equation**: A balanced chemical equation for a chemical reaction the includes the change in enthalpy (energy term) as a reactant or product
   - Indicates the amount of energy that will be released/absorbed for a *specific* amount/moles of reactants and products
   - * *Exothermic:* The enthalpy change/energy term is always written on the product side because energy is being released
   - Example:

$$CH_{4(g)} + 2O_{2(g)} \longrightarrow CO_{2(g)} + 2H_2O_{(g)} + 802.5kJ$$

802.5kJ is released when 1mol of CH_4 & 2mols of O_2 combust

   - * *Endothermic:* The enthalpy change/energy term is always written on the reactant side because energy is being absorbed
   - Example:

$$6H_2O_{(l)} + 6CO_{2(g)} + 2802.5kJ \longrightarrow 6O_{2(g)} + C_6H_{12}O_{6(s)}$$

2802.5kJ is required to react 6 moles of $H_2O_{(l)}$ with 6 moles of $CO_{2(g)}$

   - Enthalpy change is directly proportional to the amount of substances that are being consumed and produced in the reaction (ie. if you double the amount of all reactants, the enthalpy doubles as well).

$$12H_2O_{(l)} + 12CO_{2(g)} + 5605kJ \longrightarrow 12O_{2(g)} + C_6H_{12}O_{6(s)}$$

2. **ΔH Notation**: The enthalpy change/energy term is written as a separate expression beside/at the end of a chemical equation
   - * ΔH is always *negative* for an *exothermic* reaction
   - Example:

$$CH_{4(g)} + 2CO_{2(g)} \longrightarrow CO_{2(g)} + 2H_2O_{(g)} \quad \Delta H = -802.5kJ$$

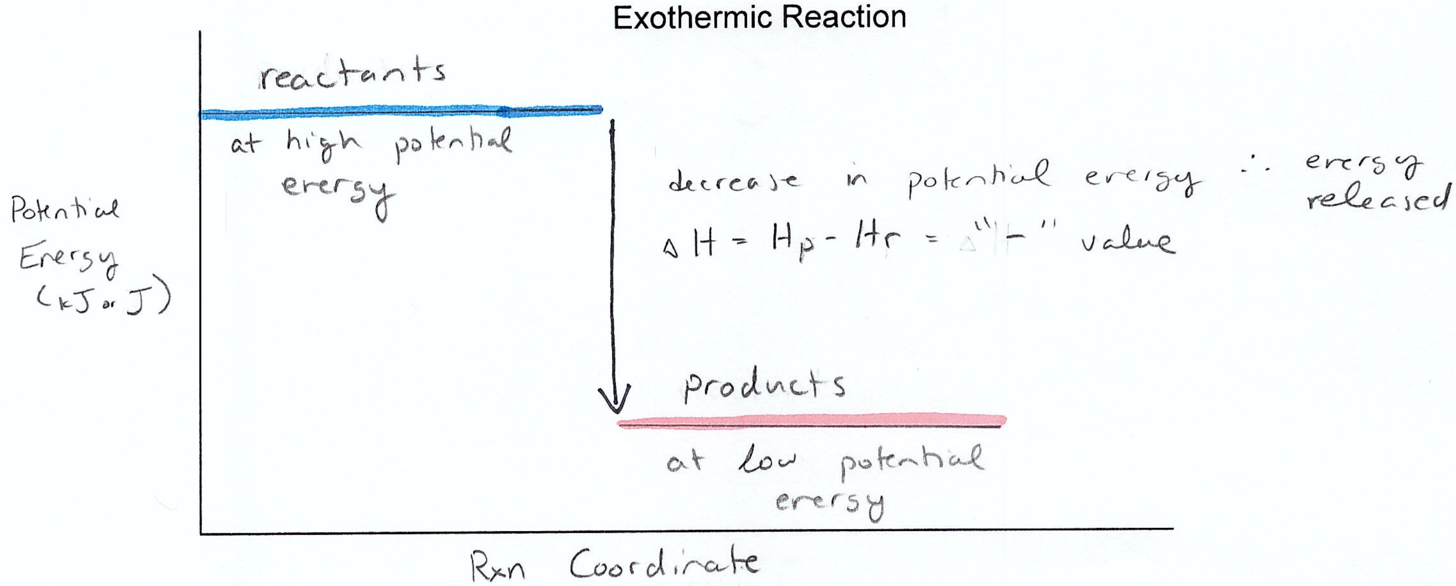
   - * ΔH is always *positive* for an *endothermic* reaction
   - Example:

$$6H_2O_{(l)} + 6CO_{2(g)} \longrightarrow 6O_{2(g)} + C_6H_{12}O_{6(s)} \quad \Delta H = 2802.5kJ$$

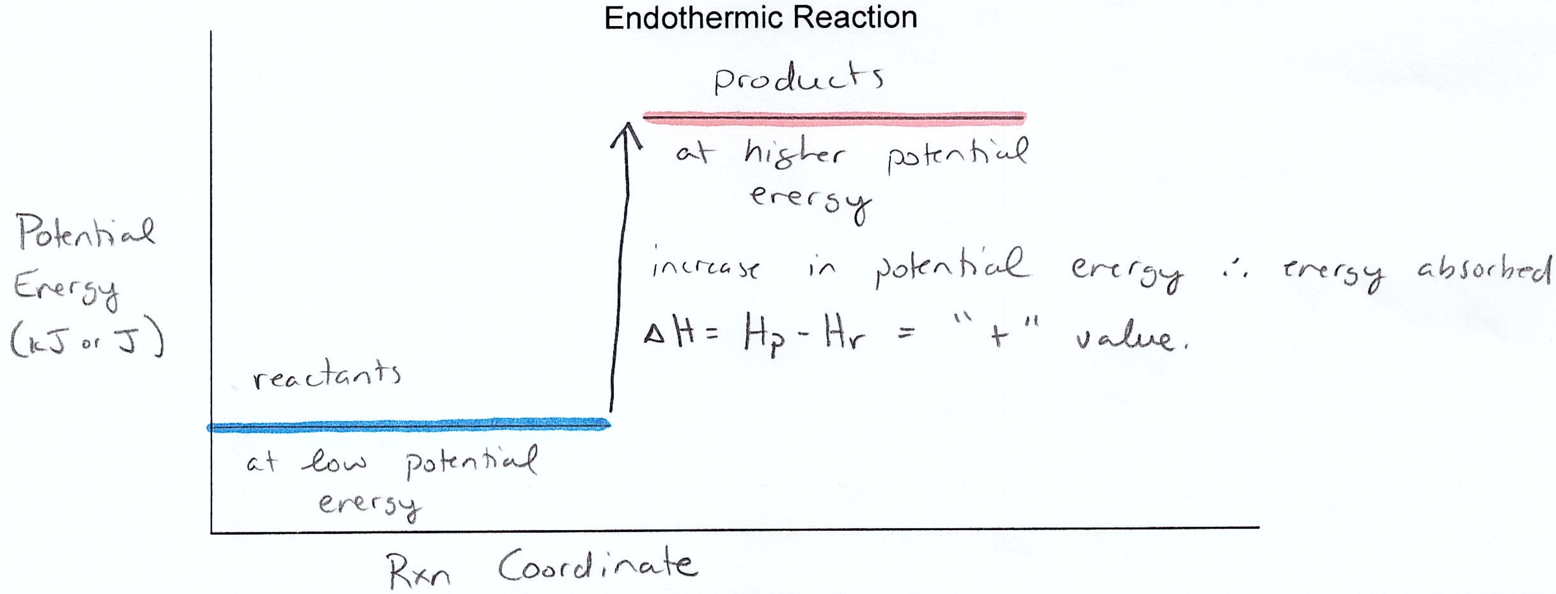
- Again, enthalpy change is directly proportional to the amount of substances that are being consumed and produced in the reaction (ie. if you double the amount of all reactants, the enthalpy doubles as well).

3. **<u>Potential Energy Diagram:</u>** a graphical representation of the change in potential energy as a chemical reaction progresses
    - Looks like a "*step-down*" for *exothermic* reactions
    - Looks like a "*step-up*" for *endothermic* reactions

Exothermic Reaction



Endothermic Reaction



*****Now try pg. 346 #3-4 & pg. 350 #2, 3, 5*****

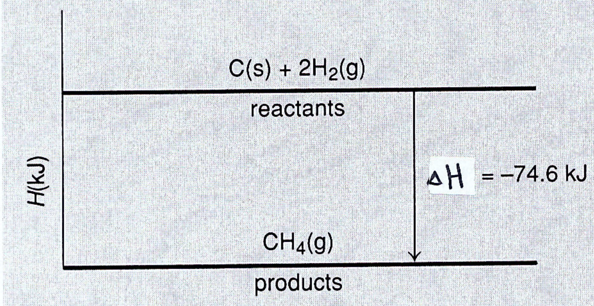
Answers to Questions for Comprehension

Student Textbook page 346

Q3. (a) $C(s) + 2H_2(g) \rightarrow CH_4(g) + 74.6\ kJ$

(b) $C(s) + 2H_2(g) \rightarrow CH_4(g)\ \Delta H = -74.6\ kJ$

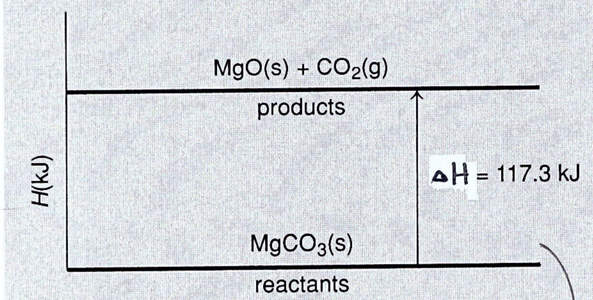
(c)



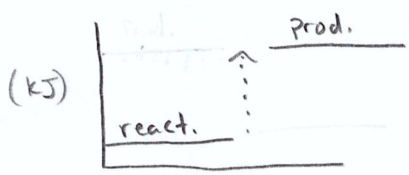
Q4. (a) $MgCO_3(s) + 117.3\ kJ \rightarrow MgO(s) + CO_2(g)$

(b) $MgCO_3(s) \rightarrow MgO(s) + CO_2(g)\ \Delta H = 117.3\ kJ$

(c)



same diagram as



b/c products are still higher in enthalpy than the reactants.

Section 9.1 Review Answers

Student Textbook page 350

2. If a reaction is endothermic, more energy is consumed in breaking bonds than is released when bonds are formed.

3. (a) $Ag(s) + \frac{1}{2}Cl_2(g) \rightarrow AgCl(s)\ \Delta H = -127.0\ kJ$

(b) $C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$
$\Delta H = -1322.9\ kJ$

(c) $NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)\ \Delta H_{sol} = -44.2\ kJ$

5. In exothermic reactions, the potential energy of the products is less than the potential energy of the reactants. If the potential energy of the system decreases, the ΔH is negative and the energy of the surroundings increases. Therefore, the reaction is exothermic.