Chapter 4.4

Expressing and Measuring Chemical Change

Energy in Chemical Reactions

- Different **molecules** have different amounts of potential energy
- **Potential** energy (PE)
  - **Stored** energy a result of intermolecular bonds (between atoms/ions)

We will learn 2 ways to express energy in a balanced equation

1. As **enthalpy,** ΔH, to the right of the equation (not part of the equation)
2. **Thermochemical** equations - Directly as part of the equation

1. Enthalpy or ΔH

- A measure of the energy **produced** or used
- Its symbol is ΔH and means **change** in energy
  - The difference between the PE of products and PE of reactants
- An **extensive** property (depends on the amount of material reacting)
- Measured in **joules** or kilojoules per mole (kJ/mol)

- Methane reacting with oxygen (combustion)
- The ΔH is -2221 kJ/mol
  - This means the **products** have - 2221 kJ/mol LESS potential energy than the reactants
  - During the reaction 2221 kJ/mol of energy was released (to environment)
  - This reaction is **EXOTHERMIC**

$H_2O(s) \rightarrow H_2O(l)$ $\Delta H = 6.0kJ/mol$

- The ΔH is 6.0kJ/mol
  - The products have 6.0kJ/mol MORE potential energy than the reactants
  - During the reaction 6.0kJ/mol of energy was USED/ NEEDED
- This reaction is **ENDOTHERMIC**

Workbook

- Compete the chart on p. 198 top of page

Answers

- Energy is often released during (exothermic)
  - Neutralization
  - Combustion
  - Synthesis
- Energy is often absorbed during (endothermic)
  - Decomposition

Generally

- **Breaking** bonds requires energy input while bond forming results in energy release.

If ΔH is POSITIVE (+ΔH)

- **Products** have more energy
- Reactants have less energy
- Reaction is ENDOTHERMIC

- Example
- $Ba(OH)_2 \cdot 8H_2O\ (s) + 2\ NH_4SCN\ (s) \rightarrow Ba(SCN)_2\ (s) + 10\ H_2O\ (l) + 2\ NH_3\ (g)$
- Energy is needed to break bonds in reactant cmpds (adding energy for reaction to occur= products more energy)

Endothermic + ΔH

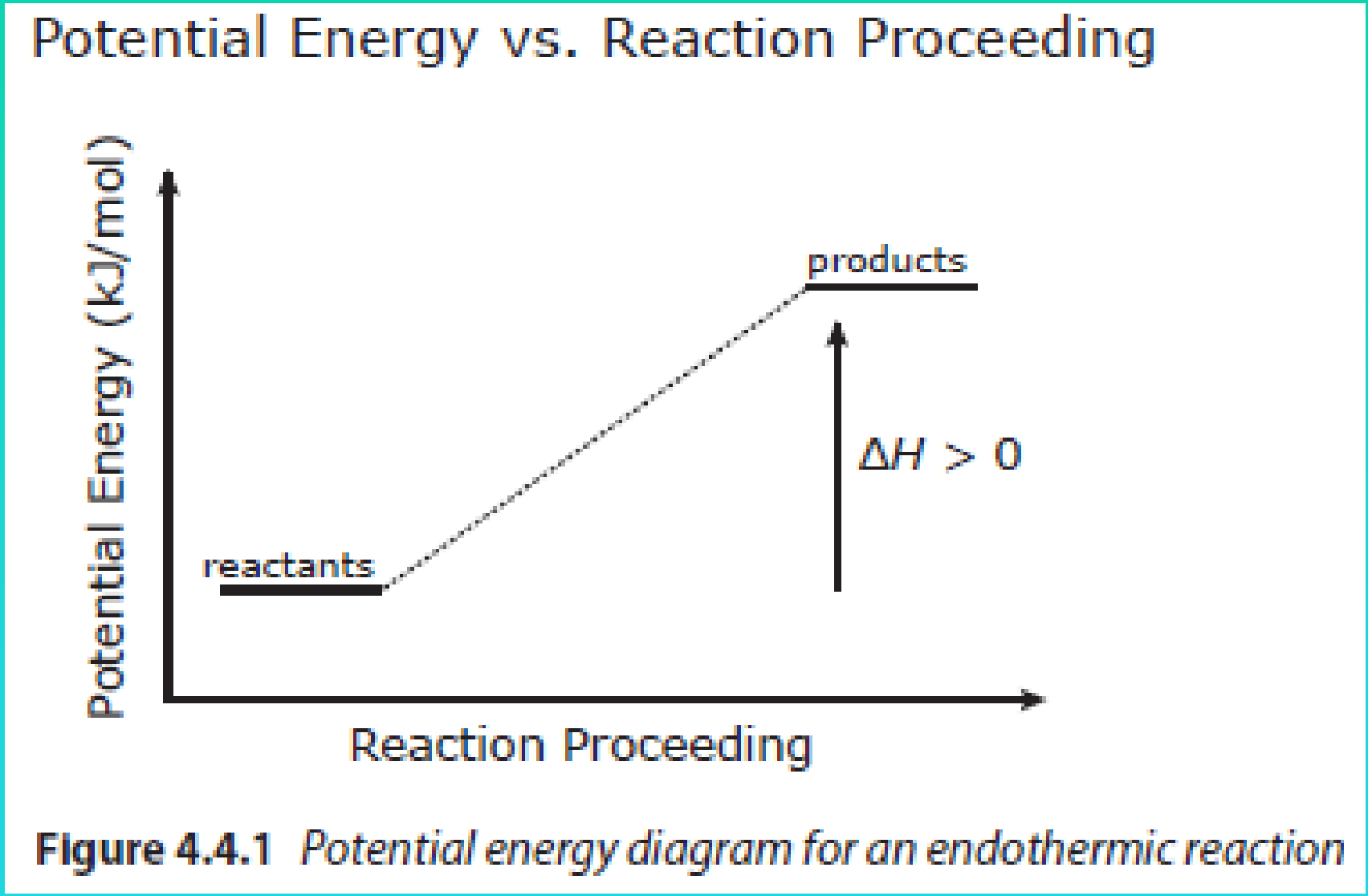


Figure 4.4.1 *Potential energy diagram for an endothermic reaction*

If ΔH is NEGATIVE (-ΔH)

- Products have less **energy** **(b/c some energy lost to environment)**
- **Reactants** have more energy
- Reaction is EXOTHERMIC

- Example:

$2\ Li(s) + 2\ H_2O \rightarrow 2\ LiOH\ (aq) + H_2(g)$

- It is the hydrogen gas that is combusting

Exothermic -ΔH

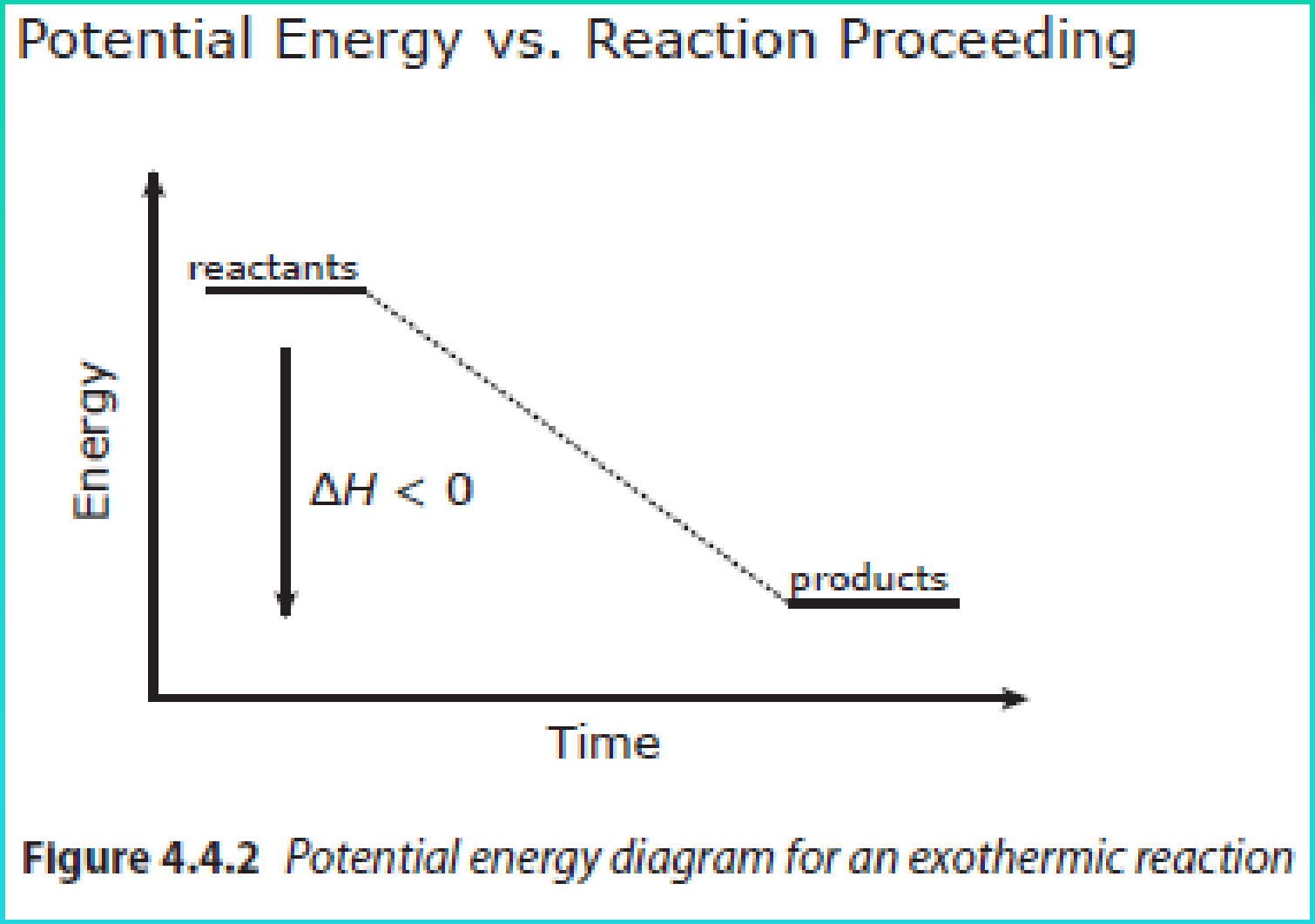


Figure 4.4.2 *Potential energy diagram for an exothermic reaction*

Workbook

- Complete the table on page 201

- Answers:

Reaction	Endothermic or Exothermic
Dissolving potassium hydroxide	exothermic
Combustion of propane	exothermic
Melting ice	endothermic
Replacement of iron by aluminum	exothermic
Formation of calcium hydroxide	exothermic

2. The second way to express energy in an equation

- Thermochemical Equations
  - Written as **part** of the equation
  - kJ/mol

Endothermic

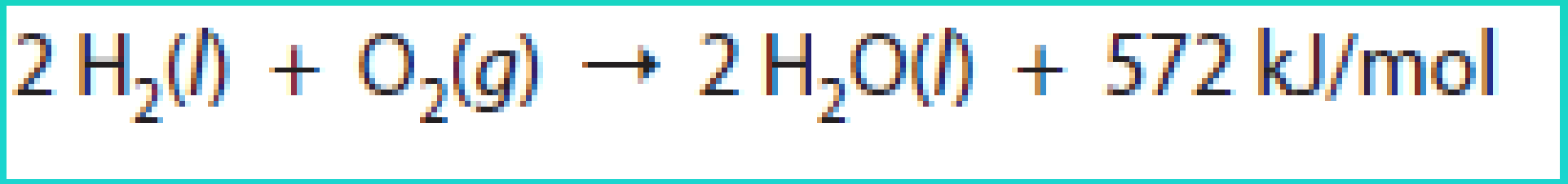
- Energy is **absorbed** by reactants to form products thus write the energy with the reactants! (energy used)

$$572\ \text{kJ/mol} + 2\,H_2O(l) \rightarrow 2\,H_2(l) + O_2(g)$$

- What would ΔH be??
- ΔH = 572kJ/mol

Exothermic

- Energy is **released** by reactants as products are formed thus write the energy on the products side (energy produced).



- What would the ΔH be?
- ΔH = -572 kJ/mol

Practice Problems p.203 #1-3

- Review some of the previous pages BEFORE asking me for help

Answers

Practice Problems — Representing Exothermic and Endothermic Changes (p. 203)

Given the following ΔH values, write a balanced thermochemical equation and an equation using ΔH notation with the smallest possible whole number coefficients for each of the following chemical changes:

1. $\Delta H_{combustion}$ of $C_2H_6(g) = -1428.5$ kJ/mol $2C_2H_{6(g)} + 7O_{2(g)} \rightarrow 6H_2O_{(\ell)} + 4CO_{2(g)} + 2857\frac{kJ}{mol}$

   $2C_2H_{6(g)} + 7O_{2(g)} \rightarrow 6H_2O_{(\ell)} + 4CO_{2(g)} \quad \Delta H = -2857\frac{kJ}{mol}$

2. $\Delta H_{decomposition}$ of $NH_3(g) = 46.1$ kJ/mol $2NH_{3(g)} + 92.2\frac{kJ}{mol} \rightarrow N_{2(g)} + 3H_{2(g)}$

   $2NH_{3(g)} \rightarrow N_{2(g)} + 3H_{2(g)} \quad \Delta H = 92.2\frac{kJ}{mol}$

3. $\Delta H_{formation}$ of $HBr(g) = -36.1$ kJ/mol $H_{2(g)} + Br_{2(g)} \rightarrow 2HBr_{(g)} + 72.2\frac{kJ}{mol}$

   $H_{2(g)} + Br_{2(g)} \rightarrow 2HBr_{(g)} \quad \Delta H = -72.2\frac{kJ}{mol}$

Homework

- Read Chapter 4.4
- Review Questions:
  - **1-6 all**
- Quiz on 4.1, 4.2, 4.4: Monday, Mar5th