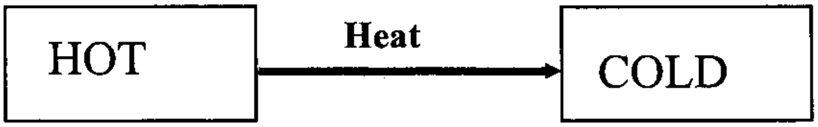
Kinetics

All chemical reactions involve a change of energy. During any chemical reaction, energy will either be released or absorbed.
Remember, **heat always flows from hot to cold.**



Exothermic: If a chemical reaction releases energy it is called "**ex**othermic". Think of it this way. If energy (heat) **ex**its, then the reaction is **ex**othermic. An **ex**othermic reaction causes its surroundings to get warm as stored (potential) energy **ex**its and is converted to kinetic. A burning match is an excellent example of an **ex**othermic reaction.
Endothermic: If a chemical reaction absorbs energy it is called "**en**dothermic". Think of it this way. If energy (heat) goes **in** (**en**ters), then the reaction is **en**dothermic. An **en**dothermic reaction causes its surroundings to get cool as kinetic heat energy is absorbed by the process and stored away as potential energy. A chemical cold pack from a first aid kit is a good example of an **en**dothermic reaction.

Describe: Exothermic ____

____

Describe: Endothermic ____

____

Chemical Reactions and Energy: Look at the word **"Kinetics"**. It looks similar to **"Kinetic energy"**. Remember! Temperature measures kinetic energy of particles. We can't see them with our eyes but all chemical reactions involve collisions of tiny particles. The particles may be molecules, ions, or atoms. They may be solid, liquid, gas or aqueous, but they are still tiny moving particles. The reactants must collide with enough kinetic energy to start a reaction. If particles are merely physically together they will not react. Only when particles collide with enough energy the reaction will occur. Different reactants require different amounts of energy to initiate (start) a reaction. The energy required to activate (start) a reaction is called **"activation energy"**.

Describe: Activation energy____

____

Once a reaction has started, it will either absorb or release energy. Remember, if the reaction absorbs energy it is **en**dothermic. If the reaction releases energy the reaction is **ex**othermic.

Collision Theory: Collision Theory states that for particles to chemically react they must have an **effective collision**. If an effective collision occurs a chemical reaction will occur. An effective collision occurs when colliding particles contact each other at the **proper angle** (orientation) and with the **right amount of energy**. The rate (speed) of reaction is controlled by the rate of effective collisions. The greater the rate of effective collisions, the greater the rate of reaction.

An effective collision occurs when colliding particles contact each other at the

______________________________ and with the ________________________________

What are the factors that can affect the rate of effective collisions thereby controlling the **rate of reaction**?

Factors affecting the rate of reaction:

1. **Temperature**: Temperature measures the average kinetic energy of particles. The greater the kinetic energy, the more effective collisions occur. This results in a greater the rate of reaction.
2. **Concentration**: High concentration results in more effective collisions. This results in a greater the rate of reaction.
3. **Surface Area**: Surface area describes how much of a reactant's surface is available to react. High surface area produces more effective collisions. This results in a greater the rate of reaction. A solid reactant that is powdered will react faster than a solid reactant that is a solid cube because a powder has greater surface area.
4. **Nature of Reactants**: This describes what the reactants are. Are they molecular (covalent)? or are they aqueous ionic? The general rule is: solutions of ionic substances react faster than molecular (covalent) substances. **Why?** Aqueous ions are free to form bonds without the need to first break bonds. Therefore, aqueous ionic reactions are quicker. Molecular (covalent) compounds must first break existing bonds before new ones can form. Therefore, molecular (covalent) reactions are slower.
5. **Catalyst**: A catalyst speeds up the rate of reaction by providing an alternate reaction pathway created by lowering the activation energy. By lowering the activation energy there are more effective collisions. This results in a greater the rate of reaction.
   A catalyst increases the rates of both the forward and reverse reactions.
   A catalyst is not consumed (used up) in the reaction. You may recognize the term "catalytic converter". An automotive catalytic converter contains catalysts such as platinum and rhodium.
   The opposite of a catalyst is an **inhibitor**. An inhibitor slows down the rate of reaction by increasing the activation energy.

List 5 factors that determine reaction rate:

1.
2.
3.
4.
5.

Potential Energy Diagrams: If you are taking the Chemistry regents (and you are) you will have to interpret and label potential energy diagrams. Energy released or absorbed by a chemical reaction can be represented by a potential energy diagram. In this topic, potential energy is represented as **H** (stored **Heat**). The energy released or absorbed by a chemical reaction is called **Heat of Reaction.** Heat of reaction is represented as **ΔH**. Heat of reaction (ΔH) equals the difference between the potential energy of the products and the potential energy of the reactants. The energy **stored** in chemical bonds is **potential** energy. Potential energy diagrams plot potential energy vs. reaction coordinate. On a potential energy diagram, potential energy may be expressed as **H**.

You will see **ΔH** represented as $\Delta H = PE_{products} - PE_{reactants}$ or $\Delta H = H_{products} - H_{reactants}$.

To help you remember! **P** comes before **R** in the alphabet song. **P** for **P**roducts comes before **R** for **R**eactants in the formula.

Potential energy diagrams plot *potential energy vs. reaction coordinate.* The reaction coordinate can be thought of as the time it takes to complete a chemical reaction.

You must know how to interpret potential energy diagrams.

Key potential energy diagram terms:

Activation Energy: Is the amount of energy required to start (initiate) a chemical reaction.

Activated Complex: Is the highest point of any potential energy diagram (top of the hill). It has the greatest potential energy on the diagram. It consists of a temporary combination of reactant particles as they begin to react.

Forward Reaction: The forward reaction is the normal "left to right" reaction path. ⟶

Reverse Reaction: The reverse reaction is the reaction proceeding from "right to left". ⟵

Heat of Reaction: Is represented as **Δ H** (called delta H). **Δ** is the symbol for change or difference. Heat of reaction is the difference in the potential energy of the products and the potential energy of the reactants. Its formula is $\Delta H = H_{products} - H_{reactants}$.

Describe: Heat of reaction ____

Describe: Activated complex ____

____

Table I has many heats of reactions.

The units for Heat of Reaction (Δ H) are **kJ** (kilojoules). It is right on the table. A **- Δ H** indicates an exothermic reaction. A reaction is exothermic if the potential energy of the products is less than the potential energy of the reactants. Think about this. The only way a reaction can be exothermic is if the energy of its products are lower than the energy of the reactants.

A **+ ΔH** indicates an endothermic reaction. A reaction is endothermic if the potential energy of the products is greater than the potential energy of the reactants.

Understanding + or - Δ H

$$CH_{4(g)} + 2\ O_{2(g)} \rightarrow CO_{2(g)} + 2\ H_2O_{(l)} + \mathbf{890.4\ kJ}$$

Methane **oxygen** **carbon dioxide** **water** **heat**

Reactants **Products**

The equation above shows 1 mole of methane **reacting** with 2 moles of oxygen to **produce** 1 mole of carbon dioxide, 2 moles of water, and 890.4kJ of heat.
Heat is on the product side. Heat is produced by the reaction.
This is an **exothermic** reaction in the normal forward direction.

Do you find it confusing that a **- Δ H** is shown as + heat? Think of the + signs in an equation as saying "and".

$$CH_{4(g)} \text{ and } 2\ O_{2(g)} \rightarrow CO_{2(g)} \text{ and } 2\ H_2O_{(l)} \text{ and } \mathbf{890.4\ kJ}$$

Methane **oxygen** **carbon dioxide** **water** **heat**

Reactants **Products**

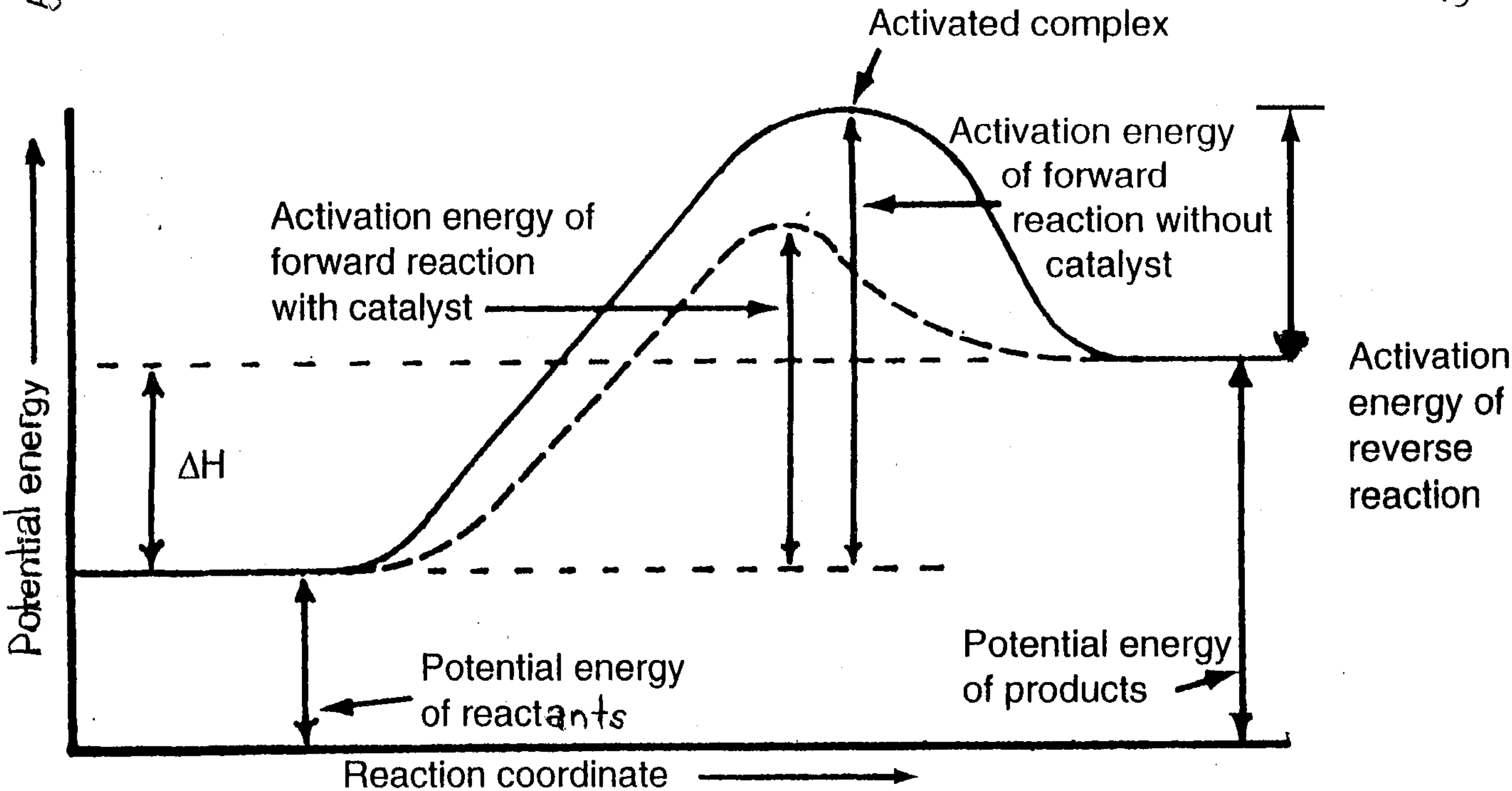


Figure 6-2. Potential Energy Diagram For An Endothermic Reaction

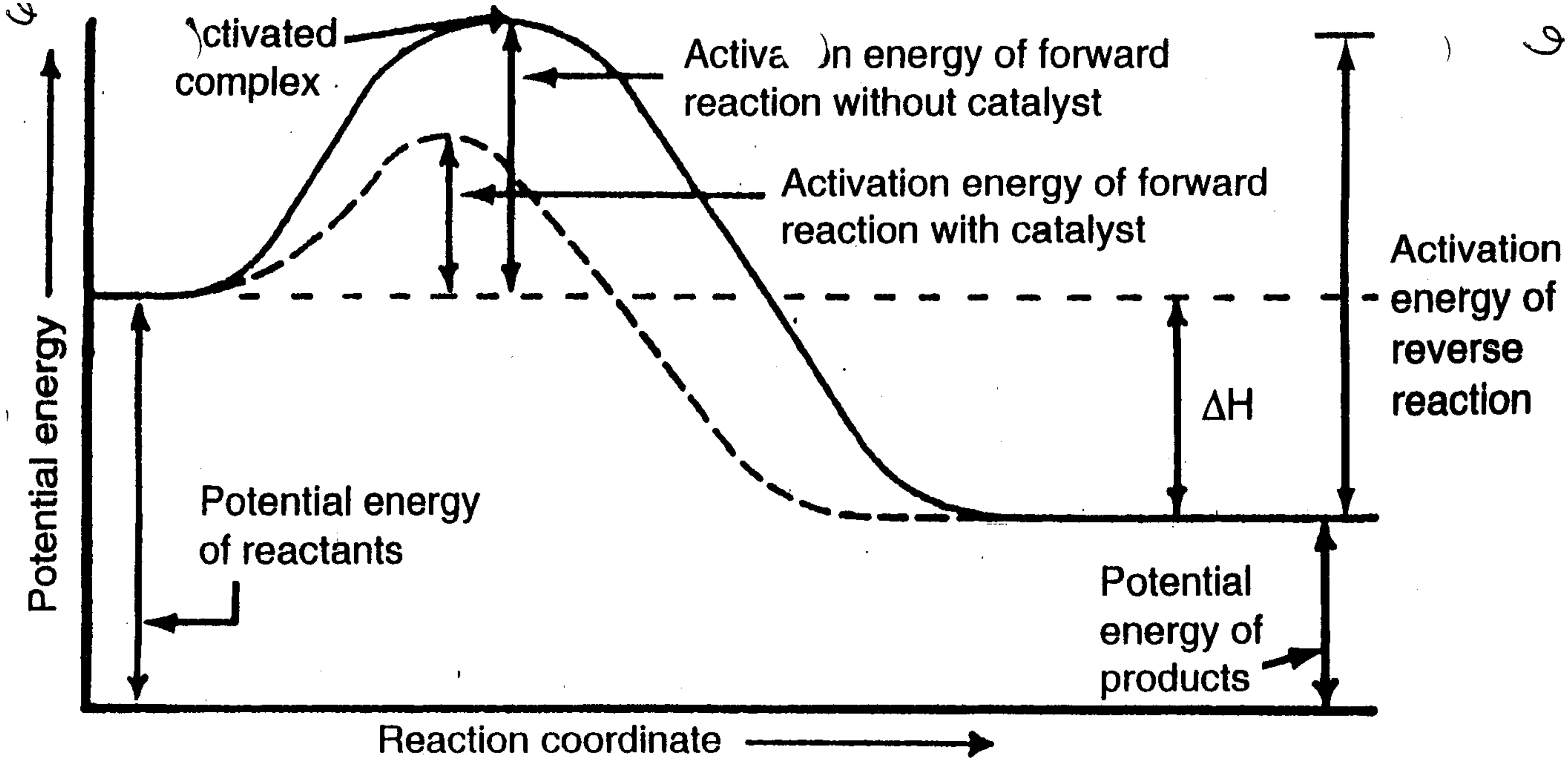


Figure 6-1. Potential Energy Diagram For An Exothermic Reaction

For the following potential energy diagram in the **FORWARD DIRECTION** determine the following: **Include appropriate units.**

a) potential energy of products

b) potential energy of reactants

c) activation energy without catalyst

d) activation energy with catalyst

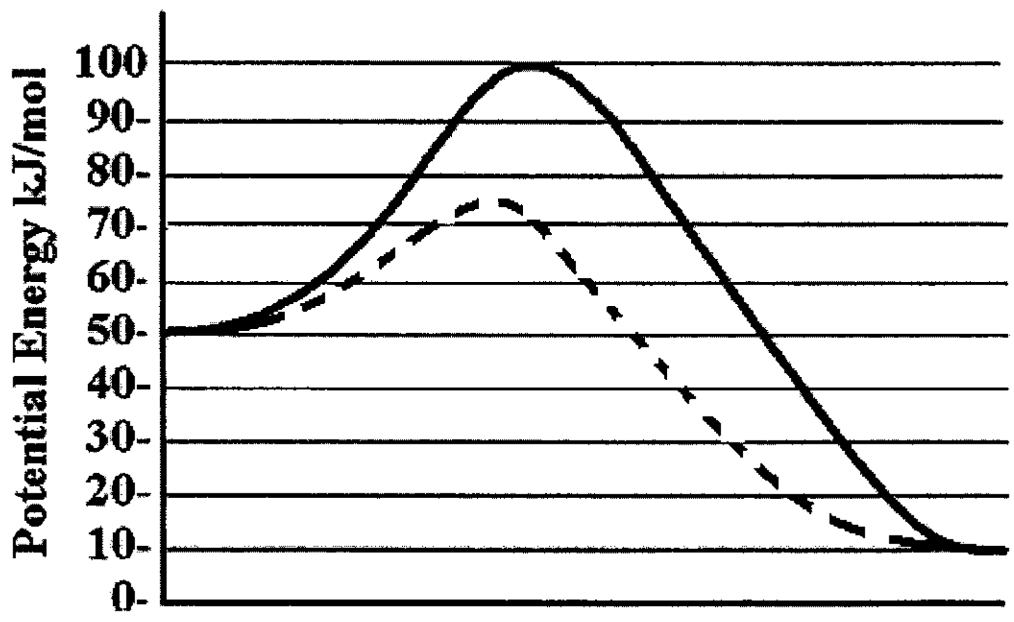
e) heat of reaction "ΔH"

f) Is this an endothermic or exothermic reaction?

g) Which has more potential energy, products, or reactants?

h) potential energy of activated complex without catalyst

i) potential energy of activated complex with catalyst



For the following potential energy diagram in the **REVERSE DIRECTION** determine the following: **Include appropriate units.**

a) potential energy of products

b) potential energy of reactants

c) activation energy without catalyst

d) activation energy with catalyst

e) heat of reaction "ΔH"

f) Is this an endothermic or exothermic reaction?

g) Which has more potential energy, products, or reactants?

h) potential energy of activated complex without catalyst

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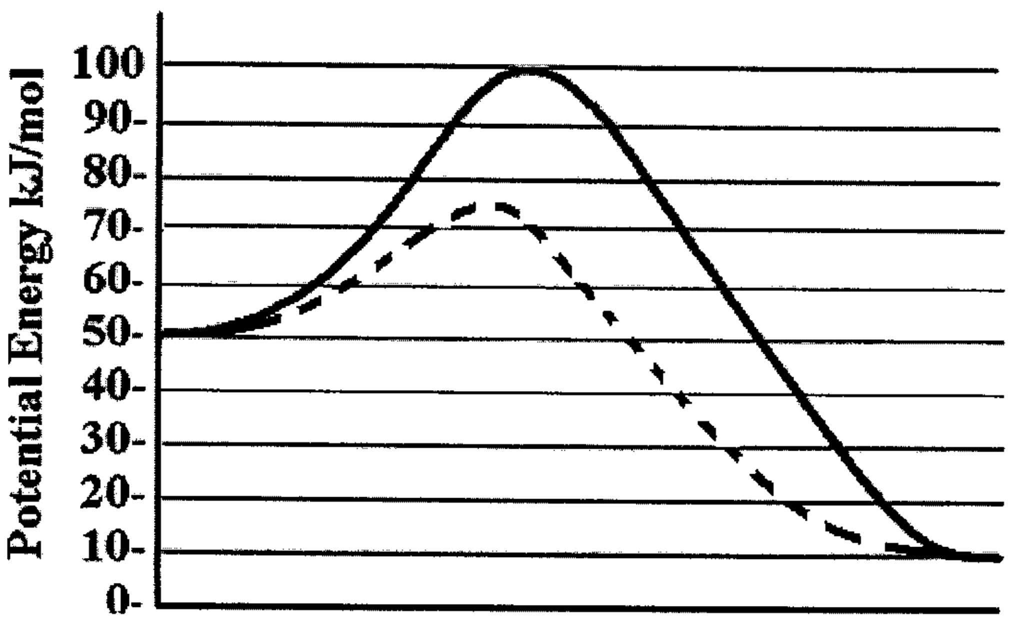

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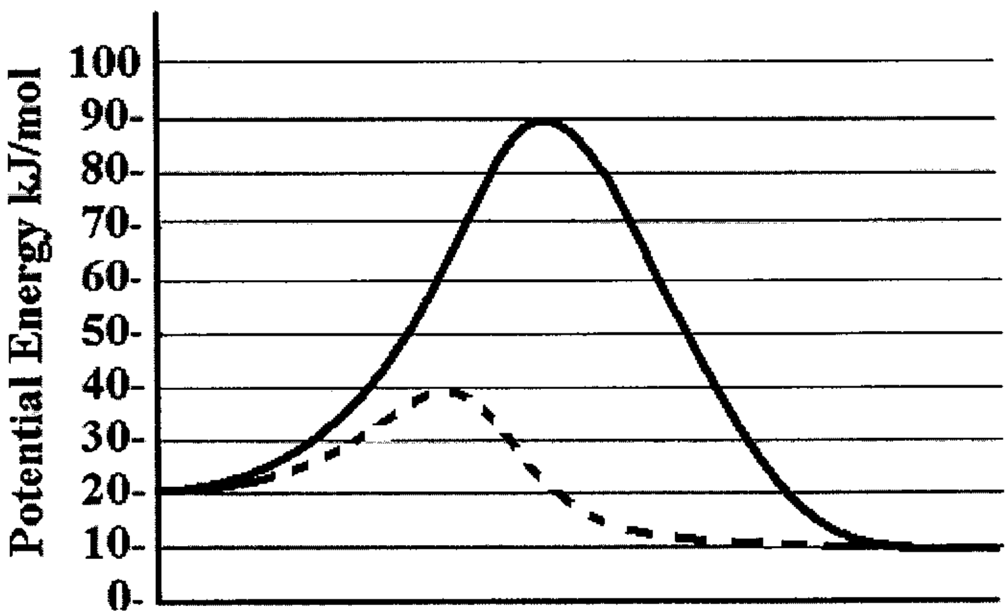

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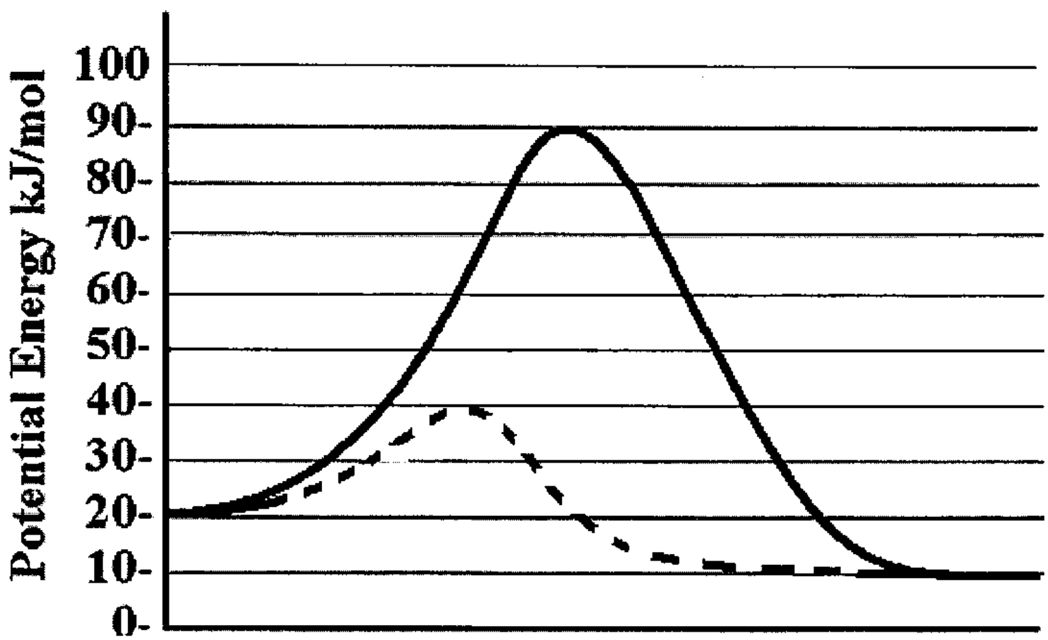

How to use Table I

Table I lists the heat gained or lost for chemical reactions and dissolving.
Here is the first chemical reaction in Table I.

$CH_{4(g)} + 2\ O_{2(g)} \rightarrow CO_{2(g)} + 2\ H_2O_{(l)}$

In this chemical reaction 1 mole of methane reacts with 2 moles of oxygen to produce 1 mole of carbon dioxide and 2 moles of water. This reaction has a **- ΔH** which means it is exothermic (heat is released) Heat is on the product side. It produces 890.4 kJ.
Heat energy is produced and appears as a product.

$$CH_{4(g)} + 2\ O_{2(g)} \rightarrow CO_{2(g)} + 2\ H_2O_{(l)} + \mathbf{890.4\ kJ.}$$

All equations from table I are balanced. If the number of moles (coefficients) of reactants are doubled the amount of heat produced would also be doubled.
How much heat would be released if 2 moles of methane ($CH_{4(g)}$) reacted with 4 moles of oxygen ($O_{2(g)}$)?

To solve:

Step 1: Set up a ratio.

$$\frac{\mathbf{2}}{\mathbf{1}}\ CH_{4(g)} + 2\ O_{2(g)} \rightarrow CO_{2(g)} + 2\ H_2O_{(l)} + \frac{\mathbf{X}}{\mathbf{890.4\ kJ}}$$

Step 2: Cross multiply and divide to solve for X. $\frac{2}{1} = \frac{X}{890.4kJ}$ **X = 1780.8 kJ**

1) How much heat is released in the combustion of 10 moles of octane C_8H_{18}?
Show your work

2) How much heat is released when 12 moles of Al react completely with oxygen?
Show your work

Here is the first example of dissolving in Table I.
The arrow with H_2O above it indicates dissolving.
Soluble ionic compounds break apart into + and - ions in water.

$$KNO_{3(s)} \xrightarrow{H_2O} K^+_{(aq)} + NO_3^{-1}{}_{(aq)}$$

The above example process has a + **ΔH**, indicating that it is endothermic. This means that when one mole of potassium nitrate dissolves it absorbs 34.89kJ. This process will cause the water to get cold. **Why?** The endothermic dissolving absorbs heat. The water loses heat to the process.

Entropy: The symbol for entropy is **S**. Entropy is the term used to describe the amount of disorder or randomness of a system. The greater the disorder of a system, the higher its entropy. Think of it this way. A clean, well organized room has low entropy. A messy, disorganized room has high entropy. Any physical or chemical process that increases the amount of disorder or randomness increases the entropy of that system. **Warm particles vibrate faster than cold particles and therefore have higher entropy.**
Liquid particles have more disorder than solid particles. Gas particles have more disorder than liquid or solid particles.

Solid → Liquid → Gas

→ entropy increases →
→ disorder increases →

You can identify an increase in entropy by comparing phases.
If a liquid or gas is produced from a solid, entropy increases.
If a gas is produced from a solid or liquid, entropy increases.

Entropy also increases as dissolving occurs. When a solid dissolves in a liquid the system becomes more random and disorderly. Entropy increases.
It can be represented as:
$C_6H_{12}O_{6(s)} \rightarrow C_6H_{12}O_{6(aq)}$ Solid glucose is dissolving. Entropy is increasing!
Remember ! (aq) stands for aqueous (mixed or dissolved in water).
$NaCl_{(s)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)}$ Solid NaCl is dissolving. Entropy is increasing!
NaCl is ionic when it dissolves it becomes positive $Na^+_{(aq)}$ ions and negative $Cl^-_{(aq)}$ ions.

State whether entropy increases or decreases for the following, and name the phase change or process (if applicable).

1. $H_2O_{(l)} \rightarrow H_2O_{(g)}$
2. $H_2O_{(l)} \rightarrow H_2O_{(s)}$
3. $H_2O_{(s)} \rightarrow H_2O_{(l)}$
4. $H_2O_{(g)} \rightarrow H_2O_{(l)}$
5. Liquid water being warmed.
6. $CO_{2(g)} \rightarrow CO_{2(s)}$
7. $CO_{2(s)} \rightarrow CO_{2(g)}$
8. $C_6H_{12}O_{6(s)} \rightarrow C_6H_{12}O_{6(aq)}$
9. $NaCl_{(s)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)}$
10. Solid iron being cooled.

Spontaneous Reactions occur by themselves. A spontaneous reaction is driven by the forces of nature. A spontaneous reaction does not need of a constant source of energy. It happens by itself.
There are 2 factors that combine to make a reaction proceed spontaneously.
The **2 factors** are.

1. The **release of energy**. Systems with a high amount of energy tend to release energy to the surroundings (exothermic).
2. An increase of **entropy**. Organized systems tend to become disorganized.

Both of the above factors work together to determine if a reaction will be spontaneous.

Kinetics Questions:
1. List **5 factors** that can affect the rate of a chemical reaction and state the effect of each.

A)

B)

C)

D)

E)

2. Define: Entropy. What is its symbol?

3. Define: Heat of Reaction. What is its symbol?

4. What table lists many Heats of Reaction?
5. What are the units for Heat of Reaction?
6. What does a - ΔH indicate?
7. What does a + ΔH indicate?
8. What is the formula to calculate Heat of Reaction?
9. What is a catalyst?
10. Does a catalyst get consumed during a chemical reaction?
11. Does a catalyst increase the rate of the forward reaction, reverse reaction, or both?
12. What does a catalyst do to activation energy ?
13. Describe "Collision Theory".
14. What is a spontaneous reaction?
15. What are 2 factors that determine if a reaction will be spontaneous?
16. Explain the term: activation energy

17. State whether entropy increases or decreases for the following and name the phase change.

A) $I_{(s)} \rightarrow I_{(g)}$

B) $I_{(g)} \rightarrow I_{(s)}$

18. What does temperature measure?

19. What type of energy is ***stored*** in chemical bonds?

20. Compare the entropy of a new deck of cards (still in the wrapper) and a shuffled deck of cards.

21. What type of reaction is usually faster……….ionic or covalent (molecular)? **Why?**

22. What symbol from Table I indicates dissolving?

Use Table "I" to predict if the following are endothermic or exothermic and state how much energy is absorbed or released. (include proper units)

23. 2 moles of octane reacting with 25 moles of oxygen

24. 1 mole of iodine reacting with 1 mole of hydrogen

25. 1 mole of ammonium nitrate dissolving

26. **4** moles of methane reacting with **8** moles of oxygen (Think about this one!)

27. If an exothermic process occurs in a beaker of water, will the water become *warmer* or *colder*? **Why?**

28. If an endothermic process occurs in a beaker of water, will the water become *warmer* or *colder*? **Why? Oh why?**

29. Which side of the arrow will heat be found for an endothermic reaction?

30. Which side of the arrow will heat be found for an exothermic reaction?