

HEAT, TEMPERATURE, & THERMAL ENERGY

Energy - A property of matter describing the ability to do **work**.

Work - is done when an object is moved through a distance by a force acting on the object.

Kinetic Energy - Associated with the **motion** of an object.

$$E_k = \frac{1}{2}mv^2$$

Potential Energy - Stored energy due to an object's **position**.

BONDS

Internal Energy - **Sum** of the kinetic and potential energies of the particles in a substance.

1st Law of Thermodynamics: The law of conservation of energy. Energy can be **transformed** from one form to another but it cannot be created or destroyed.

Units of Energy: Joule ($\text{kg/m}^2/\text{s}^2$) or 1 calorie = **4.184 J**

Q

→ 1g by 1°C
H₂O

Heat energy (q) is the actual energy **transferred** between substances. i.e an object cannot possess heat

physical

Why is heat energy essential in modern society?

- i) Over 80% of the world's electricity is generated using heat.
- ii) Heat energy is required to refine metals.
- iii) Heat is required to both shape and weld pieces of metals.
- iv) Heat produced by burning fuels operates many different kinds of engines.

The Kinetic Molecular Theory of Heat

Many theories have been developed in order to explain the concept of heat. The theory which appears to best explain heat is called the **kinetic molecular theory**. This theory relates heat to the motion of particles. It is a combination of the **particle theory** and the **kinetic theory of matter**.

**The four main points of this theory are as follows: **

1. All matter is composed of many tiny particles called **molecules**.
2. The molecules are separated from one another by empty space. The **distance between** the molecules is large **compared** to their **size**.
3. All molecules are **constantly moving** in some manner, and therefore possess **kinetic energy** (energy of motion).
4. When heat is added to matter, the molecules **absorb** the energy and move **faster** (their kinetic energy increases). When heat is **removed**, the molecules **slow down** (their kinetic energy decreases).

Temperature: A measure of the **average kinetic energy** of the molecules in a substance.

Thermal energy - is the total of the **kinetic energy** of the molecules (energy of motion) and the **potential energy**(energy of the bonds holding the molecules together)

Systems: Substances in which a change (physical or chemical) occurs. i.e. reactants and products, represented by a chemical equation)

Surroundings: The rest of the universe! Usually the surroundings are considered to be anything nearby capable of absorbing or releasing thermal energy.

Endothermic: Process in which heat is **absorbed** by the system from the surroundings (+q).

Exothermic: Process in which heat is **released** from the system into the surroundings (-q).

Open System: Both matter and energy can flow freely

Isolated System: Neither matter nor energy can flow freely (ideal but impossible)

Closed System: Energy can flow freely but not **matter**

RELATING HEAT CHANGE & TEMPERATURE CHANGE

Heat Capacity: Amount of heat energy required to raise the temperature of a substance (of any mass) by 1°C or 1K . $C = \text{J}/^{\circ}\text{C}$

Specific Heat Capacity: Amount of heat energy required to raise the temperature of 1 g of a substance by 1°C or 1K . $c = \text{J}/\text{g}^{\circ}\text{C}$

You are provided a glass of milk and a swimming pool full of milk.

Which will have the higher heat capacity?

Which will have the higher specific heat capacity?

Molar Heat Capacity: Amount of heat energy per **mole** required to raise the temperature of 6.02×10^{23} molecules of a substance by 1°C or 1K . $c = \text{J}/\text{mol}^{\circ}\text{C}$

Factors which affect heat capacity:

1. **Mass** - the greater the number of molecules which need their average kinetic energy increased, the **more heat** required.
2. **Temperature Change** - the **greater the temperature change**, ie. from 10°C to 30°C , compared to 10°C to 15°C , the greater the amount of heat is required.
3. **Type of Substance** - each substance has a different density and a different ability to absorb heat.

CALCULATING QUANTITIES OF HEAT

It is difficult to measure the quantity of heat transferred during a temperature change. Instead, we can calculate the quantity of heat using a simple formula, $Q = m c \Delta T$, where Q is the quantity of heat transferred, m is the mass of the substance, c is the specific heat capacity (page 743 of text), and $\Delta T = T_2 - T_1$, the change in temperature of the substance.

$$+Q_{\text{gained}} = -Q_{\text{lost}}$$

1. Calculate the amount of energy needed to heat 100 g of H_2O from 20°C to 45°C .

$$\begin{aligned} \checkmark_1 Q_{\text{gained}} &= m c \Delta T \\ &= (100)(4.184)(45 - 20) \\ \checkmark_2 &= 10460 \text{ J} \\ \checkmark_3 & \text{ } 10460 \text{ J is required} \end{aligned}$$

2. A 24.6g sample of nickel is heated to 110°C and then placed in a coffee cup calorimeter containing 125g of water at a temperature of 23°C . After the nickel cools the final temperature of the metal and water is 24.83°C . Assuming that no heat has escaped to the surroundings or has been absorbed by the calorimeter, calculate the specific heat of nickel.

$$\begin{aligned} \checkmark_1 +Q_{\text{gained}}_{\text{H}_2\text{O}} &= -Q_{\text{lost}}_{\text{Ni}} \\ \checkmark_2 +m c \Delta T_{\text{H}_2\text{O}} &= -m c_{\text{Ni}} \Delta T_{\text{Ni}} \\ \checkmark_3 c_{\text{Ni}} &= \frac{+m c \Delta T_{\text{H}_2\text{O}}}{-m \Delta T_{\text{Ni}}} \\ &= \frac{(125)(4.184)(24.83 - 23)}{-(24.6)(24.83 - 110)} \end{aligned}$$

$4 = +0.45 \rightarrow J/g \cdot C$

CHEMICAL ENERGY AND ENTHALPY CHANGE

Chemical systems include both kinetic and potential energy.

Kinetic Energies: Involved with the motion of particles.

- Electron movement within atoms
- Translation in gas and liquids, the movement of particles in a linear path
- Rotation about a bond axis
- Vibration, the oscillation of atoms connected by chemical bonds

Potential Energies: Involved with particles' positions within an attractive or repulsive force field.

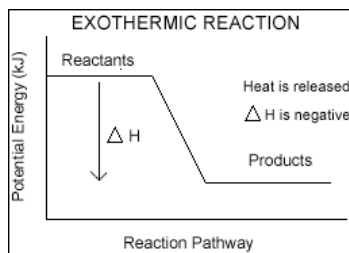
- Van der Waals Forces
- Bond energy
- Nuclear energy
- It is extremely difficult to measure the sum of all these kinetic and potential energies. Instead we study **enthalpy change, ΔH**

Enthalpy Change (ΔH): The difference in enthalpies of reactants and products during a change. AKA heat of reaction $\Delta H_{rxn} = H_{products} - H_{reactants}$

$H_{products} > H_{reactants} = + \Delta H = \text{endothermic}$

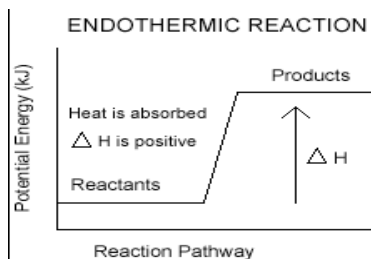
$H_{reactants} > H_{products} = - \Delta H = \text{exothermic}$

Exothermic Reactions: Less energy is required to **break** bonds in the reactants than is **released** by **formation** of new bonds in the products.



- The products of **EXOTHERMIC** reactions have **less** stored potential energy than the reactants had (more energetically stable)

Endothermic Reaction: More energy is required to **break** bonds in the reactants than is released by formation of new bonds in the products.



- The products of **ENDOTHERMIC** reactions have **more** stored potential energy than the reactants had (less energetically stable)

REPRESENTING ΔH

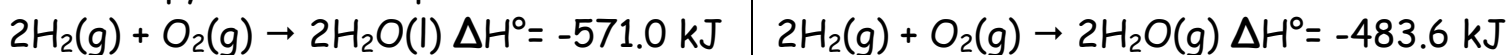
1. Thermochemical Equations with ΔH values:

E.g. $2\text{Na}(s) + 2\text{H}_2\text{O}(l) \rightarrow 2\text{NaOH}(aq) + \text{H}_2(g)$ $\Delta H = -368.6 \text{ kJ}$

Remember:

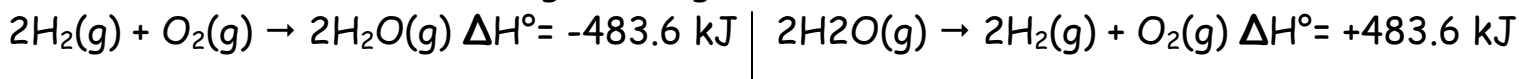
a) ΔH is a "state dependent" property, it is affected by **temperature** and **pressure**.

I.e. If a reaction produces water in the form of a gas or a liquid, the enthalpy change will be different since the enthalpy of liquid water is lower than the enthalpy of water vapour.

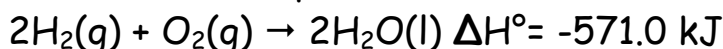


Scientists will often report results at **SATP** and use the symbol ΔH° to indicate that the value is the "standard enthalpy of reaction".

b) Exothermic reactions in one direction, become endothermic reactions in the reverse direction. (I.e. change the sign of ΔH)



c) The value of ΔH depends on the **molar** amounts of reactants and products involved.

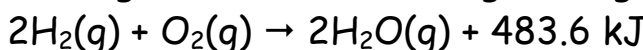


2. Thermochemical Equations with Energy Terms:

Endothermic (energy is absorbed)



Exothermic (energy is released)

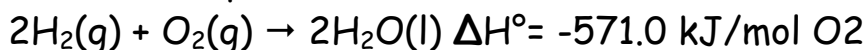


NO MINUS

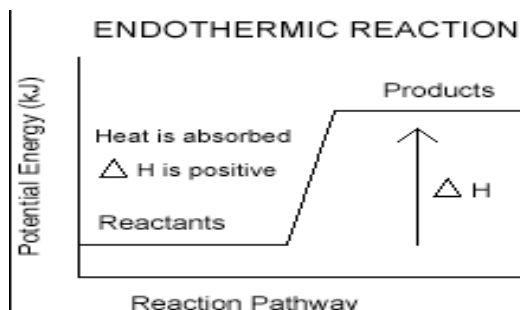
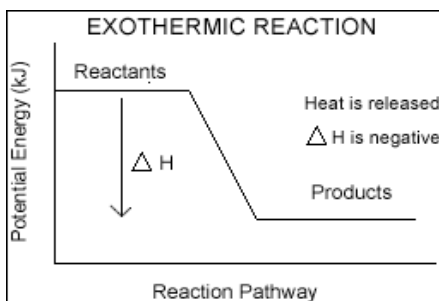
SIGNS!

3. Molar Enthalpies of Reaction:

The enthalpy change associated with 1 mole of a substance. The particular reactant or product must be specified.



4. Potential Energy Diagrams:



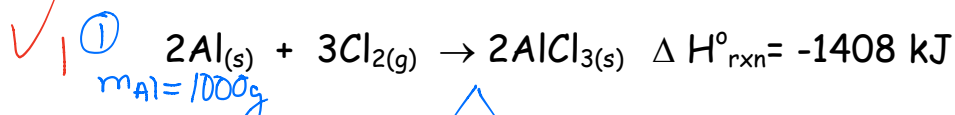
Stoichiometry and Thermochemical Equations

The enthalpy of reaction is **linearly** dependent on the quantity of products.

$$\frac{\Delta H_2}{\Delta H_1} = \frac{n_2}{n_1}$$

EXAMPLE:

1. What is the enthalpy change when 1.0 kg of Al reacts completely with excess Cl_2 according to the following equation?



✓ 2 convert to moles $\frac{m}{n/M}$

✓ 2 $n_{\text{Al}} = \frac{m}{M}$

$= \frac{1000 \text{ g}}{26.982 \text{ g/mol}}$

✓ 3 $= 37.06 \text{ mol}$

✓ 4 $\Delta H_2 = \frac{n_2 \Delta H_1}{n_1}$

$= \frac{(37.06)(-1408)}{(2)}$

✓ 5 $= -26091 \text{ kJ}$

✓ 6 -26091 kJ is produced

Calorimetry

The science of **measuring** the change in heat of **chemical reactions** or **physical changes**.

A **calorimeter** is an insulated reaction vessel in which a reaction can occur and where the change in temperature of the system can be measured

How does a Calorimeter work?

Calorimetry measures changes in temperature of a system being studied in order to "track" heat change

The calorimeter **isolates** the system from its surroundings

The examples we will look at in class will involve **constant pressure calorimetry**:

When pressure is kept constant heat measured represents the enthalpy change:

$$\Delta H = Q$$

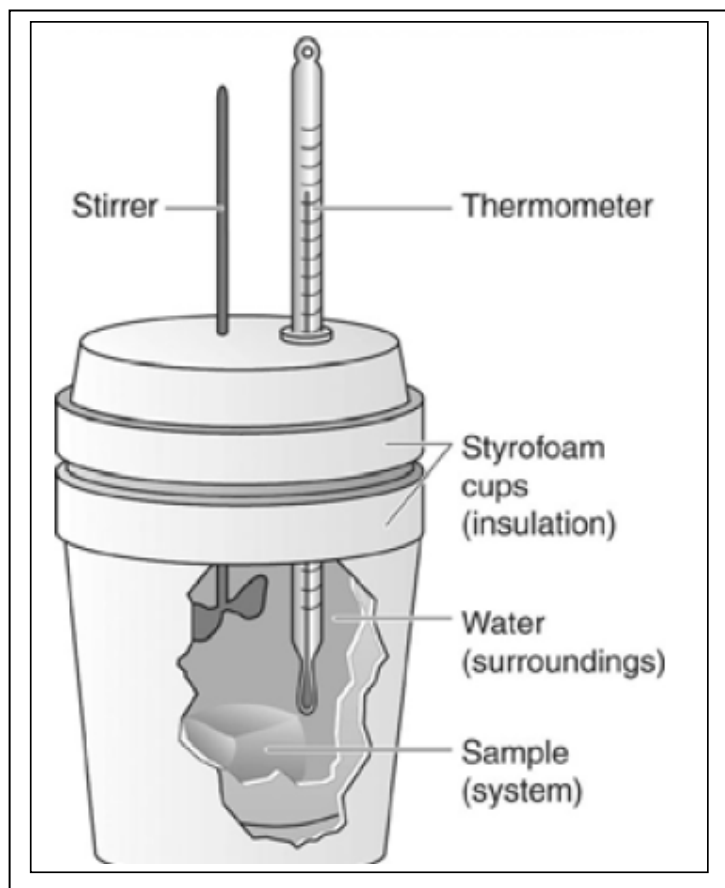


Assumptions:

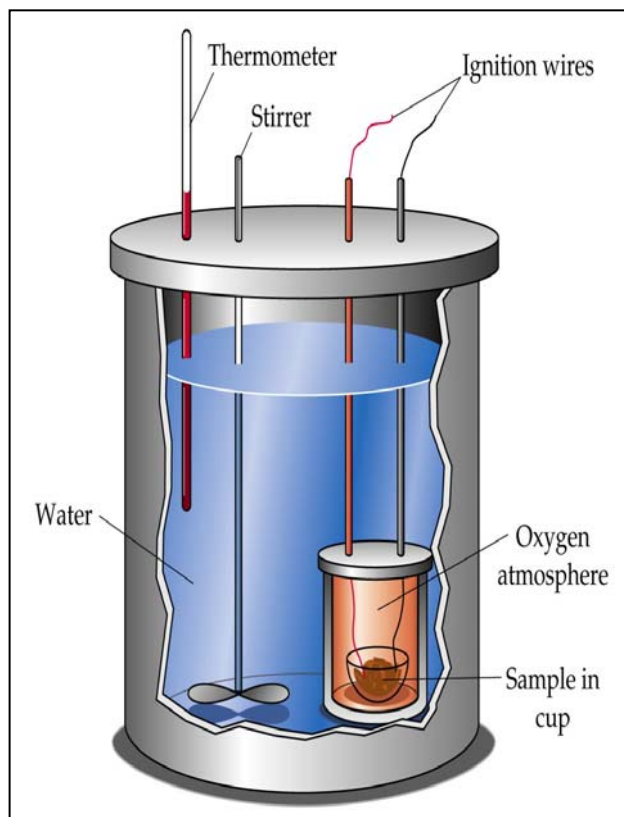
- **No heat** is transferred between the calorimeter and the outside environment
- Any heat absorbed or released by the calorimeter itself is **negligible**
- A dilute aqueous solution is assumed to have a density and specific heat capacity equal to that of **water** (i.e. $D_{\text{solution}} = 1.0 \text{ g/ml}$, $c_{\text{solution}} = 4.18 \text{ J/g}^\circ\text{C}$)

Types of Calorimeters

Coffee-cup Calorimeter



Bomb Calorimeter



Limitations:

- Cannot be used for reactions involving **gases**
- Cannot be used for **high** temperature reactions

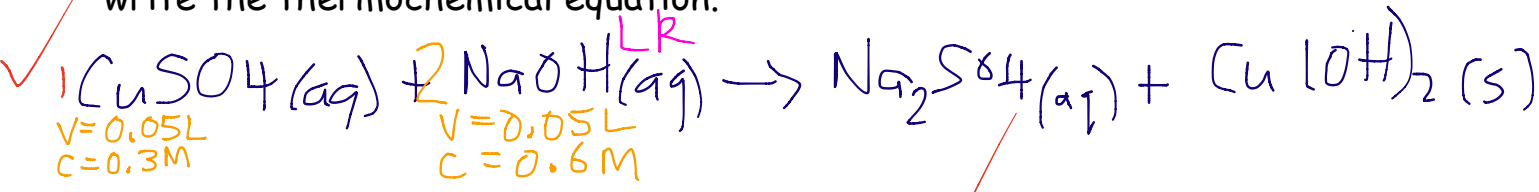
Points:

- Reaction takes place in a sealed metal container
- Temperature difference of the **water** is measured
- Calculations are more complex because they must take into account heat flow through the **metal** container

Determining the Enthalpy of a Chemical Reaction

EXOTHERMIC: $\text{temp} \uparrow$ $+\Delta H_{\text{reaction}} = -\Delta H_{\text{solution}}$
 ENDOTHERMIC: $\text{temp} \downarrow$ $-\Delta H_{\text{reaction}} = +\Delta H_{\text{solution}}$

1. 50.0 ml of 0.300 M $\text{CuSO}_4(\text{aq})$ solution is mixed with an equal volume of 0.600 M $\text{NaOH}(\text{aq})$ solution. The initial temperature of both solutions is 21.4°C . After mixing the solutions in the coffee-cup calorimeter, the highest temperature that is reached is 24.6°C . Determine the enthalpy change, ΔH , of the reaction and then write the thermochemical equation.



$n_{\text{CuSO}_4} = CV = (0.3)(0.05) = 0.015\text{mol}$
 $n_{\text{NaOH}} = CV = (0.6)(0.05) = 0.03\text{mol}$

$\frac{1 \text{ mol CuSO}_4}{0.015} = \frac{1 \text{ mol Na}_2\text{SO}_4}{x}$
 $x = 0.015\text{mol}$

$\frac{2 \text{ mol NaOH}}{0.03 \text{ mol}} = \frac{1 \text{ mol Na}_2\text{SO}_4}{x}$
 $x = 0.015\text{mol}$

✓ $\Delta H_{\text{reaction}} = -\Delta H_{\text{solution}} = -mc\Delta T$
 $= -(100)(4.184)(24.6 - 21.4)$
 $= -1339\text{J}$

$m = \rho V = (1)(100) = 100\text{g}$

$\Delta H_1 = \frac{n_2}{n_1} \Delta H_2$
 $= \frac{2}{0.03} (-1339)$
 $\Delta H_1 = -89266.7\text{J}$



Hess's Law of Heat Summation

Recall: Calorimetry is an accurate technique for determining enthalpy changes of a system.

- How do chemists deal with chemical systems that **cannot** be analyzed using this technique:

- **Slow reactions**

- **Small changes in enthalpy**

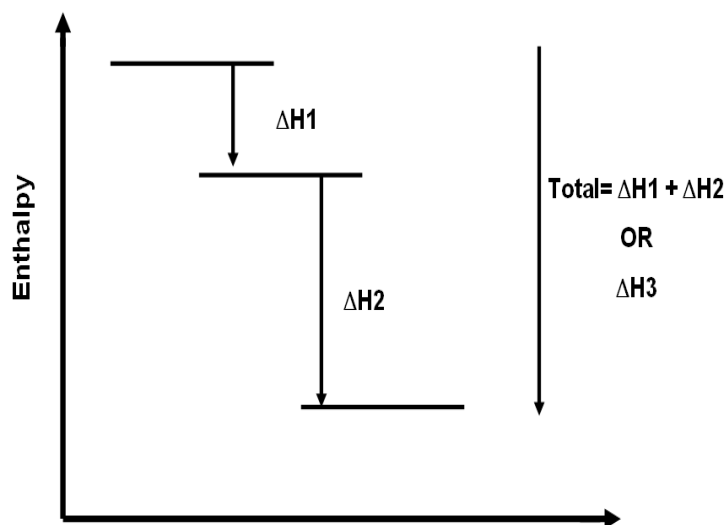
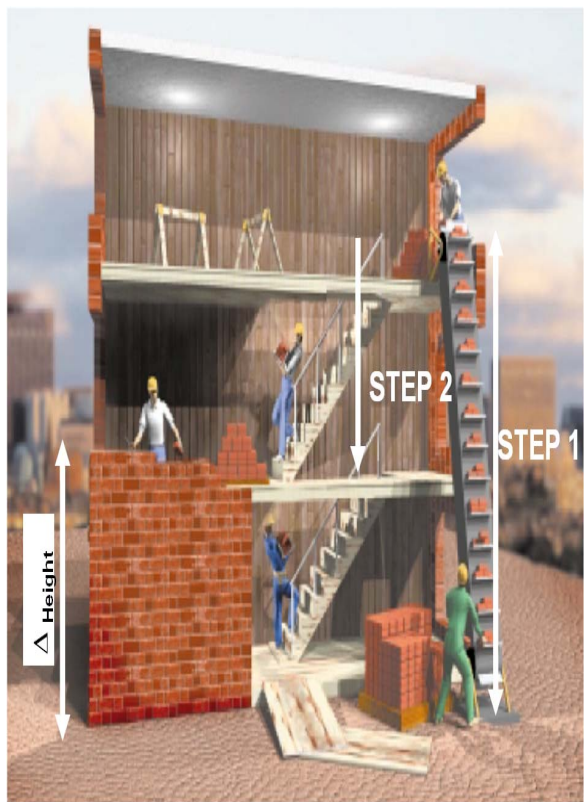
- The net changes in some properties of a system are independent of the way the system changes from the initial state to the final state.

What does this have to do with **ENTHALPY** changes???



- If a set of reactions occurs in different steps but the initial reactants and final products are the same, the overall enthalpy change is the same.

Hess's Law of Summation

For any reaction that can be written in a **series of steps**, the standard heat of reaction is the same as the sum of the **STANDARD HEATS** of reaction for the steps

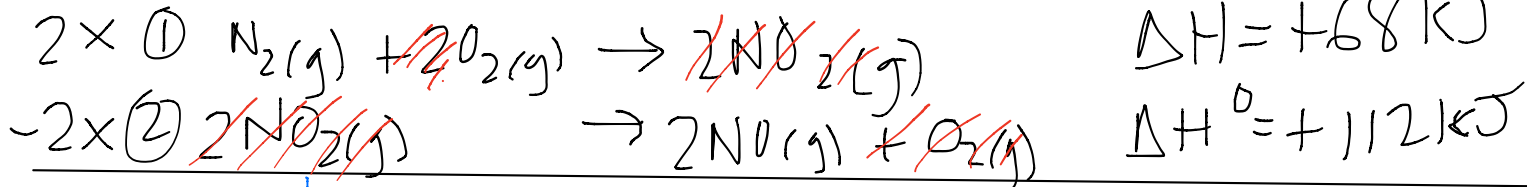
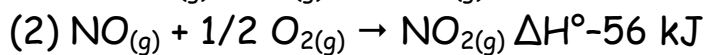
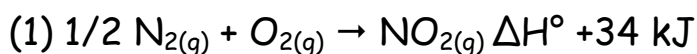


Combining and Manipulating Chemical Equations: A 7-Step process

1. Ensure ALL chemical equations are correctly balanced
 2. Examine the given equations to see how they compare with the **TARGET** 
 3. "FLIP" equations to obtain reactants and products on the correct sides 
 - ANY time you Flip (or reverse) an equation you **MUST** multiple the associated enthalpy change by [-1]
 4. Multiply coefficients in an equation by an integer or fraction if required
 - **MULTIPLY** the enthalpy value for this equation by the same factor
 5. Write the manipulated equations so that their **ARROWS** line up
 6. Add reactants and products on each side, **cancel substances** that appear on **both sides**
 7. **Add** the enthalpy changes for the combined reactions
- ** ALL equations need to add together to arrive at the TARGET equation**

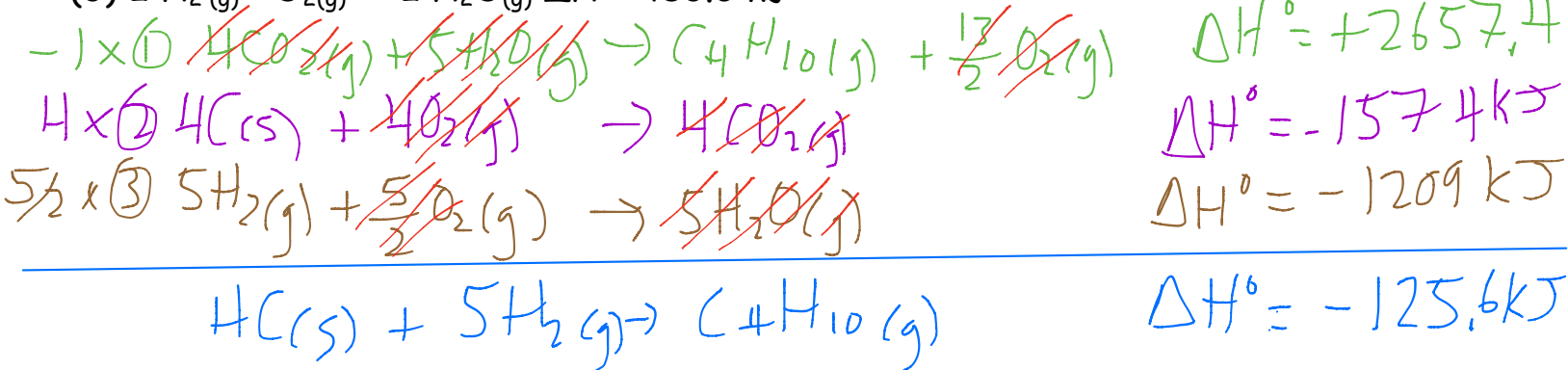
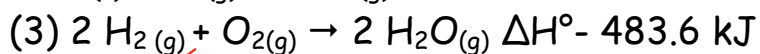
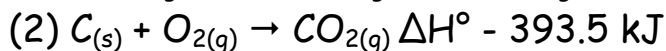
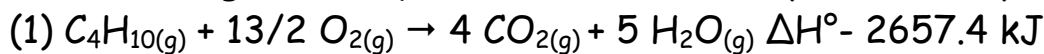
Example 1

What is the enthalpy change for the formation of two moles of nitrogen monoxide from its elements? $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}(\text{g}) \Delta H^\circ = ?$



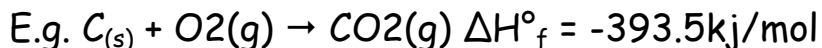
What is the enthalpy change for the formation of one mole of butane (C_4H_{10}) gas from its elements? The reaction is: $4\text{C}(\text{s}) + 5\text{H}_2(\text{g}) \rightarrow \text{C}_4\text{H}_{10}(\text{g}) \Delta H^\circ = ?$

The following known equations, determined by calorimetry, are provided:

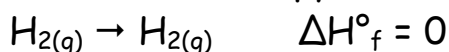


STANDARD ENTHALPIES OF FORMATION

- Reactions in which compounds are formed from their elements (in their standard states) are called **formation reactions**.



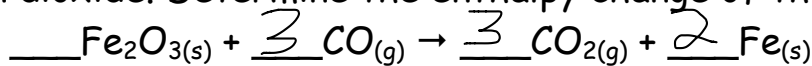
- ΔH_f° = standard enthalpy of formation ** Table of values on p. 743 **
- Always written for one mole of product.
- The product may be in any state but the reactant elements must be in their standard states.
- The standard enthalpy of formation of an element in its standard state is **zero**.



$$\Delta H = \sum n \Delta H_f^\circ(\text{products}) - \sum n \Delta H_f^\circ(\text{reactants})$$

EXAMPLE:

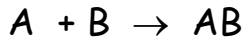
1. Iron(III)oxide reacts with carbon monoxide to produce elemental iron and carbon dioxide. Determine the enthalpy change of this reaction,



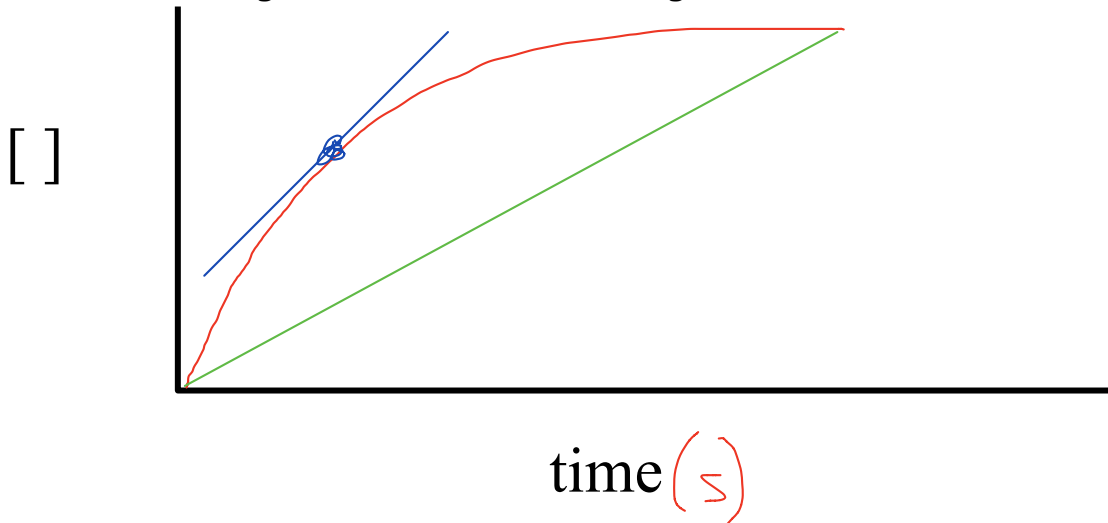
$$\begin{aligned} \Delta H_{rxn}^\circ &= \sum n \Delta H_f^\circ(\text{products}) - \sum n \Delta H_f^\circ(\text{reactants}) \\ &= \sum (3 \Delta H_f^\circ \text{ CO}_{2(g)} + 2 \Delta H_f^\circ \text{ Fe}_{(s)}) - \sum (1 \Delta H_f^\circ \text{ Fe}_2\text{O}_{3(s)} + 3 \Delta H_f^\circ \text{ CO}_{(g)}) \\ &= (3(-393.5) + 2(0)) - ((-824.2) + 3(-110.5)) \\ &= -1180.5 + 1155.7 \\ &= -24.8 \text{ kJ/mol} \end{aligned}$$

RATES OF REACTIONS

The change in **concentration** of a reactant or product per unit **time**. For example:



Rate = $\frac{\text{decrease in } [A]}{\text{change in time}}$ or $\frac{\text{increase in } [AB]}{\text{change in time}}$



Rates are usually determined at the **beginning** of the reaction due to the **maximum** amount of reactant present (max. collisions). The average or instantaneous reaction rate can be determined from a graph using the slope formula. e.g.

- i) Average rate = use first and last points
- ii) Instantaneous rate = slope of tangent

Determining Average and Instantaneous Rates of Reactions Graphically

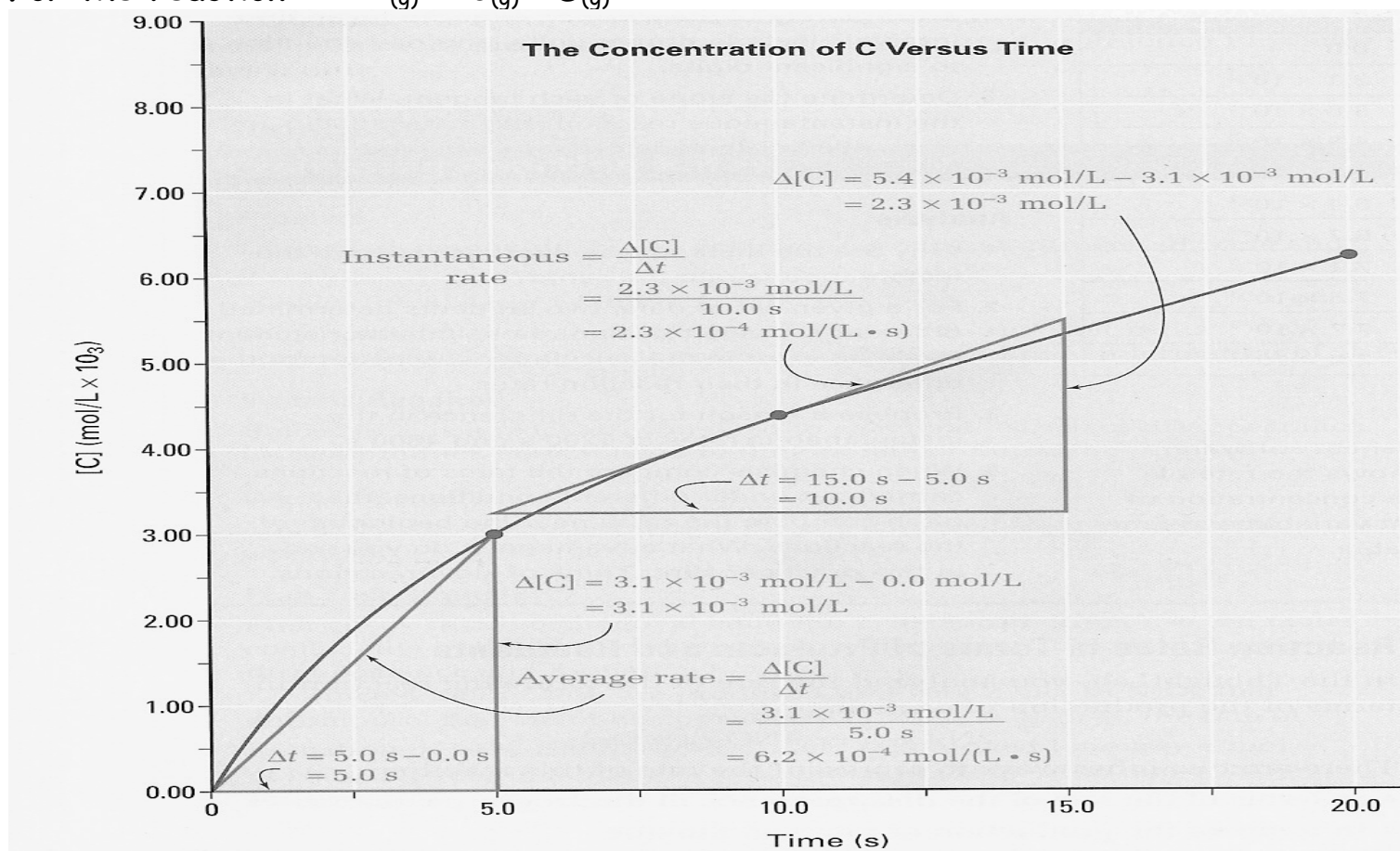
Instantaneous Rate of Reaction

- the rate of reaction at a particular time
- found using slope of **tangent** (the best straight-line approximation to the curve at a particular point. Only touches curve at one point)

Average Rate of Reaction

- the mean change in concentration of reactants or products per unit of time
- found by determining slope of **secant** (a line that intersects two or more points on the curve)

For the reaction: $A_{(g)} \rightarrow C_{(g)} + D_{(g)}$



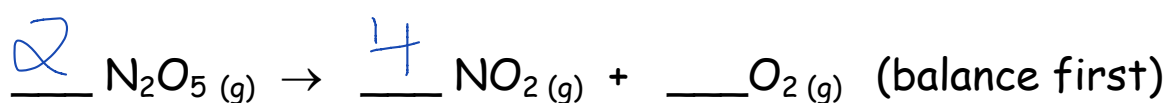
- For each line, a **steeper** slope means a **faster** rate of reaction
- As the reaction proceeds the reactants are used up and the slope of the tangent decreases, therefore rate of reaction decreases.

Reaction Rates in Terms of Reactants and Products

There are two ways to represent the rate of a reaction:

1. Rate of disappearance of reactant
2. Rate of appearance of product

For example;



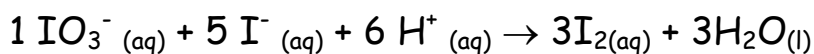
For every 1 mol of O₂ produced 4 mol of NO₂ is produced.

For every 1 mol of O₂ produced 2 mol of N₂O₅ are consumed.

EXAMPLE 2.

Consider the reaction:

$$\text{RLF} = \frac{c \text{ RLF}}{c \text{ KR}}$$



What are the rates of reaction with respect to the various reactants and products?

The rate of reaction with respect to iodate ions (rate of consumption of IO₃⁻) is determined experimentally to be 3.0 × 10⁻⁵ mol/(L·s).

$$\frac{\Delta[\text{I}^-]}{\Delta t} = \frac{5 \Delta[\text{IO}_3^-]}{1 \Delta t} = \frac{5}{1} (-3.0 \times 10^{-5}) = -1.5 \times 10^{-4} \text{ mol/L}\cdot\text{s}$$

$$\frac{\Delta[\text{H}^+]}{\Delta t} = \frac{6 \Delta[\text{IO}_3^-]}{1 \Delta t} = \frac{6}{1} (-3.0 \times 10^{-5}) = -1.8 \times 10^{-4} \text{ mol/L}\cdot\text{s}$$

$$\frac{\Delta[\text{I}_2]}{\Delta t} = \frac{-3 \Delta[\text{IO}_3^-]}{1 \Delta t} = \frac{-3}{1} (-3.0 \times 10^{-5}) = +9.0 \times 10^{-5} \text{ mol/L}\cdot\text{s}$$

$$\frac{\Delta[\text{H}_2\text{O}]}{\Delta t} = \frac{-3 \Delta[\text{IO}_3^-]}{1 \Delta t} = \frac{-3}{1} (-3.0 \times 10^{-5}) = +9.0 \times 10^{-5} \text{ mol/L}\cdot\text{s}$$

Methods for Measuring Rates of Reaction

Table 6.2 p. 362 Chemistry 12 MHR.

Property Measured	Type of Data Collected	Equipment Used	Eq'n to Determine Rate
Volume	Volume of a gas formed	Gas syringe	$rate = \frac{\Delta volume}{\Delta time}$
Mass	Change in mass of reactant or product	Balance	$rate = \frac{\Delta mass}{\Delta time}$
Temperature	Change in temp during an exo/endothermic reaction	Thermometer	$rate = \frac{\Delta temperature}{\Delta time}$
Pressure	Change in pressure of a closed container caused by production or consumption of gas	Pressure sensor	$rate = \frac{\Delta pressure}{\Delta time}$
Colour	Change in amount of light of a specific wavelength absorbed by a chemical compound; changes with [] of compound	Spectrophotometer	$rate = \frac{\Delta absorbance}{\Delta time}$
pH	Change in $[H_3O^+]$ / $[OH^-]$ ions as rx'n proceeds	pH meter	$rate = \frac{\Delta pH}{\Delta time}$
Electrical conductivity	Change in [] of dissolved ions as a reaction proceeds	Electrical conductivity probe	$rate = \frac{\Delta conductivity}{\Delta time}$

Influencing the Rate of Reaction

Collision Theory states that in order for reactions to occur molecules must collide.

- These collisions must be "effective" collisions, that is:

1. **Orientation** of molecules must be correct
2. There must be sufficient **collision energy**

Factors that Effect the Rate of Reaction

a. Temperature

- When two chemicals react, their molecules have to collide with each other with sufficient energy for the reaction to take place. By heating the mixture, you will raise the **kinetic energy** of the molecules involved in the reaction.

b. Concentration of Reactants

- Increasing the concentration of the reactants will increase the **frequency** of collisions between the two reactants.

c. Catalysts

- A catalyst is a substance that lowers the amount of **activation energy**, (EA), necessary to initiate a chemical reaction, but is not consumed in the reaction.

d. Surface Area of a solid reactant

- If a solid particle is broken up, the surface area of the molecule is increased. Increased surface area means an increased number of possible **sites** for reaction to occur.

e. Pressure of gaseous reactants & products

- By increasing the pressure, molecules are forced **closer** together which will increase the frequency of collisions between them.

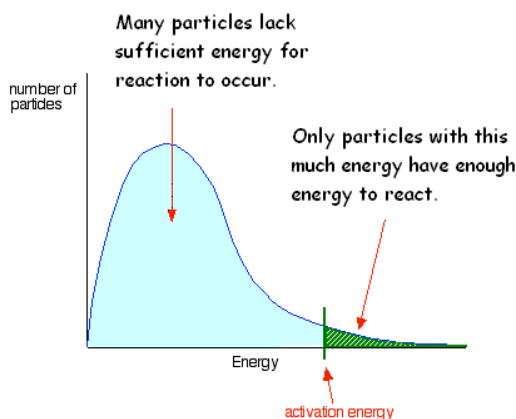
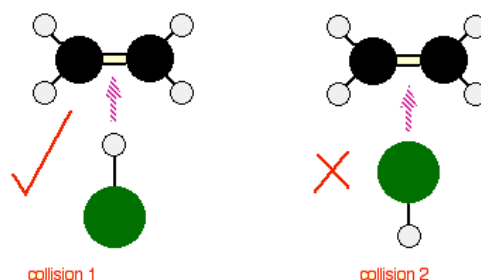
Activation Energy

Recall - For a reaction to occur:

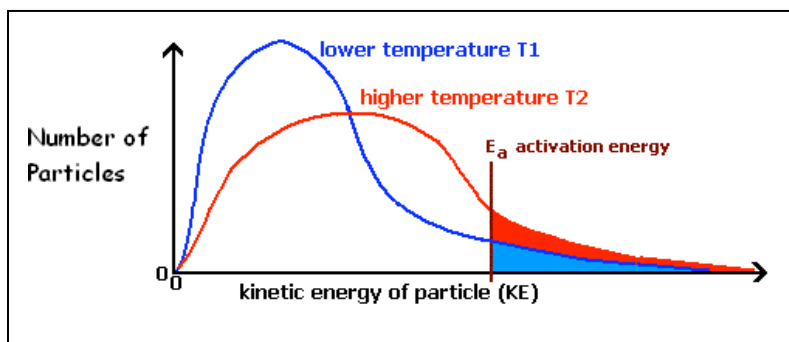
- a. particles must collide in a specific orientation
- b. particles must collide with sufficient kinetic energy

- **Activation energy** is the minimum collision energy required for a successful reaction to occur.

- Diagrams called **Maxwell Boltzmann Distribution Curves** are plots of kinetic energy vs. number of particles.



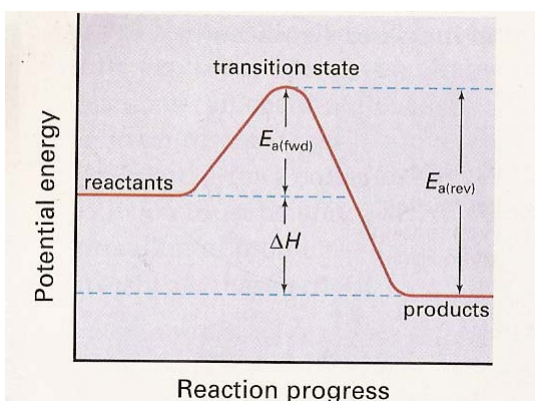
- As temperature increases the kinetic energy of molecules increases. Therefore more particles will have sufficient energy required to react.



Potential Energy Diagrams

- When molecules collide, kinetic energy of the particles is converted to potential energy. Potential energy diagrams are used to illustrate the change in potential energy during a reaction.

Exothermic reaction



Endothermic Reaction

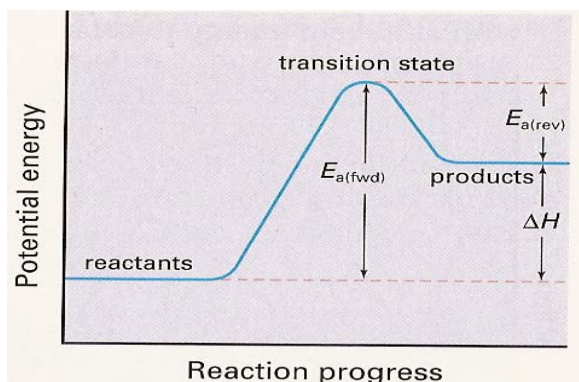
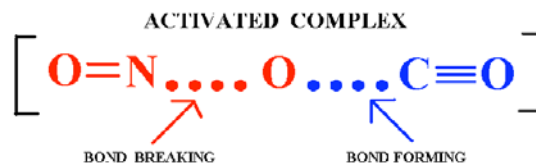
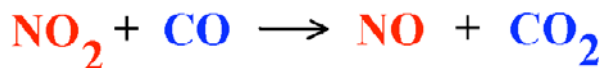


Diagram Terminology

- **Ea** - activation energy
- For an exothermic reaction: $E_{a(rev)} = E_{a(fwd)} - \Delta H$
- For an endothermic reaction: $E_{a(rev)} = E_{a(fwd)} + \Delta H$
- **Transition state** - point when reactants are converted to product.
- **Activated complex** - chemical species that exist at the transition state. For example:

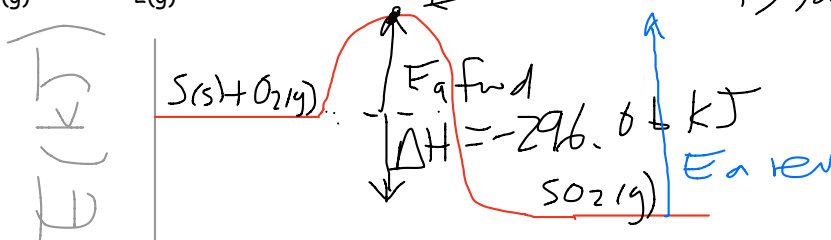


Questions:

1. It is a general rule that with a 10°C temperature increase most reaction rates will double. This is not the result of doubling the number of collisions. Explain.

2. Sketch a PE diagram for the following reactions. Include labels for Ea, ΔH, & transition state.

a. $\text{S}_{(s)} + \text{O}_{2(g)} \rightarrow \text{SO}_{2(g)} \quad \Delta H = -296.06 \text{ kJ}$ transition state

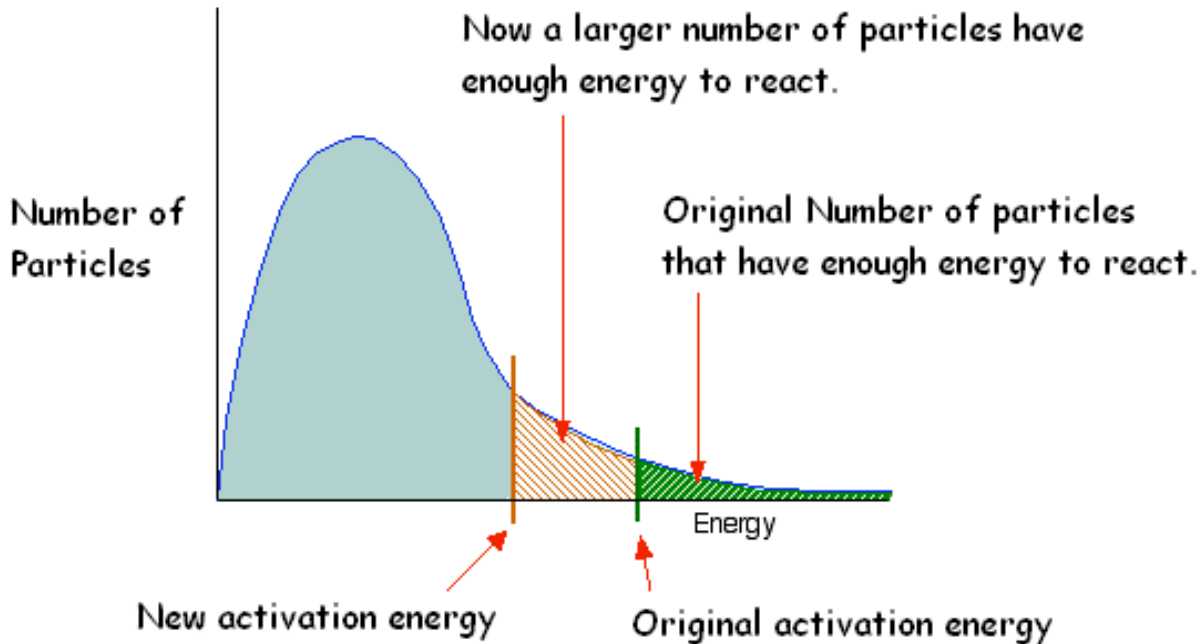


Catalysts

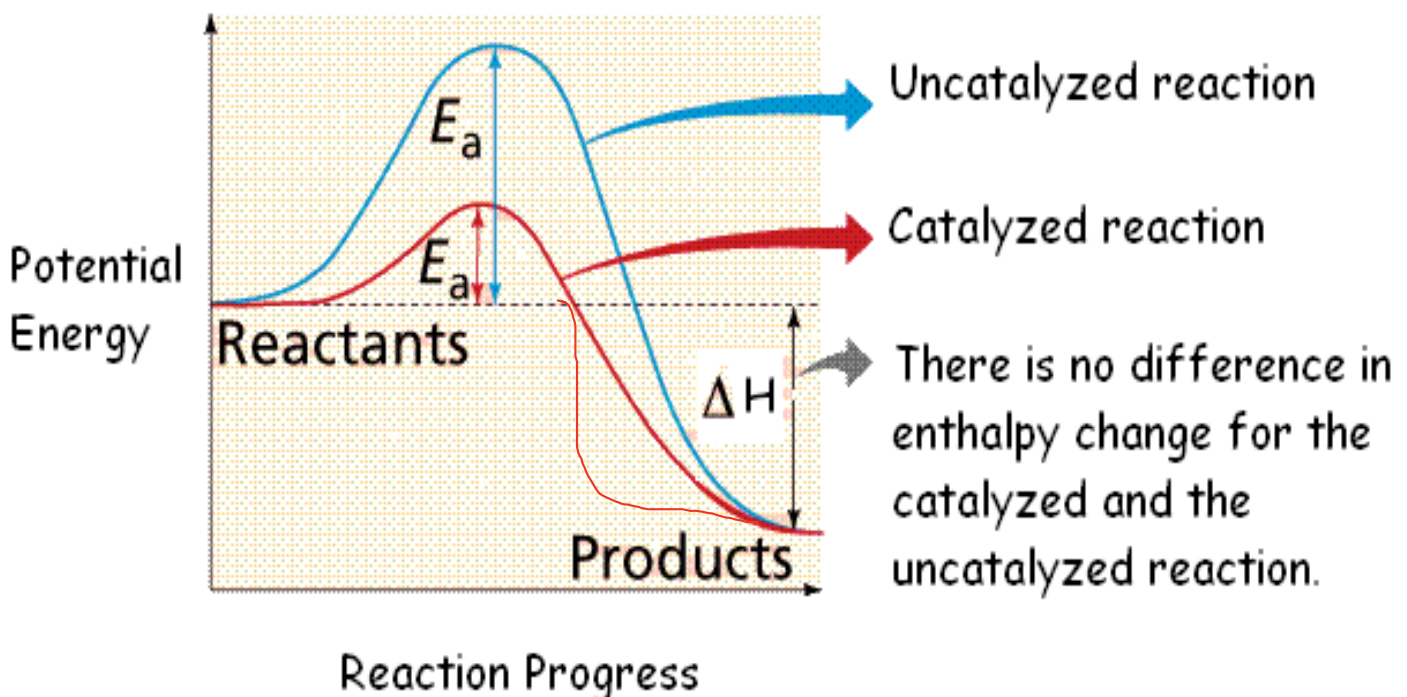
time (s)

Recall: A catalyst is a substance that **increases** the rate of a chemical reaction without being **consumed** in the reaction.

- A catalyst works by **lowering** the activation energy of a reaction so that a **larger** number of reactants have **sufficient** energy to react.

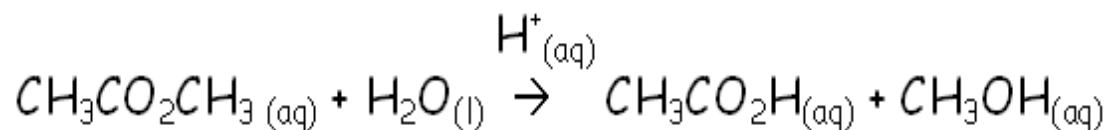


Potential Energy Diagram



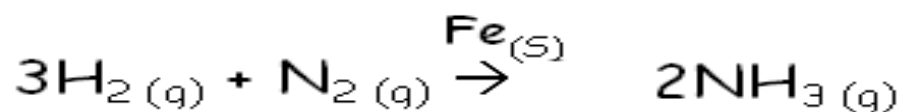
Homogeneous catalysts

- exist in the same phase as the reactants.
- most often catalyze gaseous & aqueous rxns



Heterogeneous catalysts

- exists in a different phase than reactants.
- without a catalyst this type is very slow



Enzymes

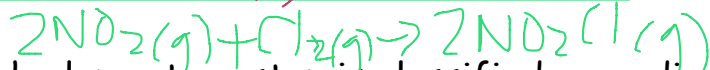
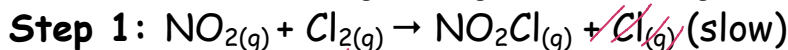
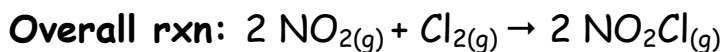
- Enzymes are organic catalyst used in biological reactions. In an enzymatic reaction the reactant molecule(s) are called the **substrate** and the **active site** is the portion of the enzyme where the substrate binds to the enzyme.

Reaction Mechanisms

$$r = k[\text{NO}_2]^1[\text{Cl}_2]^1$$

- Most reactions occur in a **series** of steps.
- These series of steps are together are called a **reaction mechanism**.
- Each individual step is called an **elementary reaction**.
- Molecules that form in one step and are consumed in the next are called **reaction intermediates**.

For example this reaction occurs in two steps:



Each elementary step is classified according to number of reactants.

- **Unimolecular** - elementary reaction with one particle
- **Bimolecular** - elementary reaction with two particles
- **Termolecular** - elementary reaction with three particles (rare)

In an **elementary reaction**, the exponents in the rate law equation are the same as the stoichiometric coefficients.

Elementary Reaction	Rate Law
$A \rightarrow \text{products}$	Rate = $k[A]$
$A + B \rightarrow \text{products}$	Rate = $k[A][B]$
$2A \rightarrow \text{products}$	Rate = $k[A]^2$
$2A + B \rightarrow \text{products}$	Rate = $k[A]^2[B]$

Elementary Reaction Rate Law

A reaction mechanism must:

1. contain equations that **combine** to give **overall** equation
2. contain **reasonable** elementary steps
3. support the experimentally determined **rate law**

- Each elementary reaction has its **own** rate.
- The **slowest** elementary reaction is called the **rate-determining step**.
- It is assumed that this "slow" step by itself **controls** the rate of reaction.
- As a result the rate law for the rate-determining step is the rate law for the overall reaction.
- From the above example the first step is the **rate determining step**. Therefore the rate law for the first step is the rate law for the overall reaction.

$$\text{Rate} = k[\text{NO}_2]^1[\text{Cl}_2]^1$$

✓ SHOW ME!