Notes: Unit 11 Kinetics and Equilibrium
KEY IDEAS

- Collision theory states that a reaction is most likely to occur if reactant particles collide with the proper energy and orientation. (3.4d)
- The rate of a chemical reaction depends on several factors: temperature, concentration, nature of reactants, surface area, and the presence of a catalyst. (3.4f)
- Some chemical and physical changes can reach equilibrium. (3.4h)
- At equilibrium the rate of the forward reaction equals the rate of the reverse reaction. The measurable quantities of reactants and products remain constant at equilibrium. (3.4i)
- LeChatelier’s principle can be used to predict the effect of stress (change in pressure, volume, concentration, and temperature) on a system at equilibrium. (3.4j)
- Energy released or absorbed by a chemical reaction can be represented by a potential energy diagram. (4.1c)
- Energy released or absorbed during a chemical reaction (heat of reaction) is equal to the difference between the potential energy of the products and the potential energy of the reactants. (4.1d)
- A catalyst provides an alternate reaction pathway, which has a lower activation energy than an uncatalyzed reaction. (3.4g)
- Entropy is a measure of the randomness or disorder of a system. A system with greater disorder has greater entropy. (3.1ll)
- Systems in nature tend to undergo changes toward lower energy and higher entropy. (3.1mm)

PROCESS SKILLS

- Use collision theory to explain how various factors, such as temperature, surface area, and concentration, influence the rate of reaction. (3.4vi)
- Describe the concentration of particles and rates of opposing reactions in an equilibrium system (3.4iv)
- Compare the entropy of phases of matter (3.1xxii)
- Qualitatively describe the effect of stress on equilibrium, using LeChatelier’s principle. (3.4v)
- Identify examples of physical equilibria as solution equilibrium and phase equilibrium, including the concept that a saturated solution is at equilibrium (3.4vii)
- Distinguish between endothermic and exothermic reactions, using energy terms in a reaction equation, ΔH, potential energy diagrams, or experimental data. (4.1)
- Read and interpret potential energy diagrams: PE reactants, PE products, activation energy (with or without a catalyst), heat of reaction (4.1ii)
**Vocabulary:**

<table>
<thead>
<tr>
<th>Word</th>
<th>Definition</th>
</tr>
</thead>
<tbody>
<tr>
<td>Activated complex</td>
<td>An intermediate structure formed in the conversion of reactants to products; the structure at the maximum energy point along the reaction pathway.</td>
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<tr>
<td>Activation energy</td>
<td>The minimum energy required to convert reactants into products; the difference between the energies of the activated complex and the reactants.</td>
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<tr>
<td>Catalyst</td>
<td>A substance that is neither a reactant nor a product but functions to speed up the rate of a chemical reaction by lowering activation energy/providing a shorter or “alternate” pathway.</td>
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<tr>
<td>Collision Theory</td>
<td>In order for a chemical reaction/effective collision to occur, particles must collide with proper energy alignment.</td>
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<tr>
<td>Effective collision</td>
<td>A collision between reactant particles that results in a chemical reaction taking place.</td>
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<tr>
<td>Enthalpy</td>
<td>The total amount of potential energy stored in a substance.</td>
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<tr>
<td>Endothermic</td>
<td>A reaction that absorbs and stores energy from its surrounding environment.</td>
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<tr>
<td>Entropy</td>
<td>A measure of the randomness or chaos associated with a chemical reaction or physical change. A system’s state of disorder. Entropy increases as temperature increases. Entropy increases as a substance is heated from a solid to a liquid to a gas.</td>
</tr>
<tr>
<td>Equilibrium</td>
<td>A system where the rate of forward change is equal to the rate of the reverse change.</td>
</tr>
<tr>
<td>Exothermic</td>
<td>A reaction that releases stored energy into its surrounding environment.</td>
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<tr>
<td>Favored</td>
<td>A change in a thermodynamic property that contributes towards the reaction being spontaneous.</td>
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<tr>
<td>Heat of reaction</td>
<td>The heat energy absorbed or released during a chemical reaction; heat of products minus heat of reactants.</td>
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<tr>
<td>Kinetics</td>
<td>The study of reaction mechanism and reaction rate.</td>
</tr>
<tr>
<td>Mechanism</td>
<td>The pathway or series of steps a reaction takes on its way to forming product.</td>
</tr>
<tr>
<td>Nonspontaneous</td>
<td>A change that requires a constant input of energy to occur, or the change will stop or reverse itself.</td>
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<tr>
<td><strong>Vocabulary:</strong></td>
<td></td>
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<tr>
<td><strong>Phase Equilibrium</strong></td>
<td>When the processes of freezing and melting or evaporating and condensing are occurring at equal rates. Can occur if a substance is at its melting point or boiling point.</td>
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<tr>
<td><strong>Potential Energy Diagram</strong></td>
<td>Used to illustrate the energy lost or gained (the reaction pathway) for a given chemical reaction.</td>
</tr>
<tr>
<td><strong>Reaction rate</strong></td>
<td>The amount of reactant consumed per unit of time.</td>
</tr>
<tr>
<td><strong>Solution Equilibrium</strong></td>
<td>When the processes of dissolving and precipitating are occurring at equal rates. Can occur for a solution at its saturation point.</td>
</tr>
<tr>
<td><strong>Spontaneous</strong></td>
<td>A change that, once started, continues the rest of the way on its own with no addition of energy required.</td>
</tr>
<tr>
<td><strong>Thermodynamics</strong></td>
<td>The study of heat flow during physical and chemical changes.</td>
</tr>
<tr>
<td><strong>Unfavored</strong></td>
<td>A change in a thermodynamic property that contributes towards the reaction being nonspontaneous.</td>
</tr>
</tbody>
</table>

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Lesson 1: Energy Changes in Chemical Reactions

**Objective:**
- Distinguish between endothermic and exothermic reactions.
- Use table I (ΔH) to determine the type of reaction

**Chemical change** (reaction) involves the ___________ and ___________ of bonds to make NEW SUBSTANCES.

BARF:

THEREFORE, chemical change involves changes in ____________.

What are the three things conserved in ALL chemical and/or physical changes:

____________________, ____________________, ____________________

Potential Energy (stored in bonds) can transfer to **Kinetic Energy** (heat) (and vice versa).

If Kinetic energy (heat) goes in, it is __________thermic (energy is ____________); potential energy of products is _______________ potential energy of reactants.

\[ A + B + \text{HEAT} \rightarrow C + D \]

If Kinetic energy (heat) comes out, it is __________thermic (energy is ____________); potential energy of products is _______________ potential energy of the reactants.

\[ A + B \rightarrow C + D + \text{HEAT} \]

The NUMERICAL difference between the PE of the products and the PE of the reactants is called the _______________ (ΔH).

\[ \Delta H = PE_{Products} - PE_{Reactants} \]
Lesson 1: Energy Changes in Chemical Reactions

USING TABLE I (Heats of Reaction aka ΔH)
Remember Table I Footnote:
“A minus sign indicates an exothermic reaction”.

Minus sign means that PE_{Reactants} must be ___________________ than PE_{Products}

If PE went DOWN, energy was _______________ in the form of KE (heat).

*Table I gives numerical values.*

**Example:**
How much energy is released when 2 moles of H\(_2\)O(g) is formed from its elements?

Reverse reaction, reverse the sign...

**Example:**
What is the heat of reaction (ΔH) for the decomposition of 2 moles of Aluminum Oxide into its elements?

IN CLASS NOTES:
Using Table I to write reaction equations.
Write the balanced reaction equation for the following:

Formation of H\(_2\)O(g) from its elements:

Formation of Nitrogen Monoxide gas from its elements:

Formation of C\(_2\)H\(_6\)(g) from its elements:
Objective:
- Read and interpret potential energy diagrams.
- Determine if a reaction is endo- or exothermic based on its Potential Energy diagram.

Potential Energy Diagrams:
Shows the change in potential (stored) energy during a chemical reaction

Types of Potential Energy Diagrams:
- Recall $\Delta H = PE_{\text{products}} - PE_{\text{reactants}}$

First Possibility: PE of products is lower than reactants, $\Delta H$ is negative, ____________reaction (PE is converted to heat).

Exothermic Potential Energy Diagram:
- Product side ____________than reactant side
- WHY? Energy is released as a product, so the net amount of potential energy decreases.
- (-) $\Delta H$

Second Possibility: PE of products is ____________than reactants, $\Delta H$ is positive, ____________reaction (heat was converted to PE, now stored in the substance created by the reaction):

Endothermic Potential Energy Diagram
- Product side ____________than reactant side
- WHY? Energy is absorbed by the reactants, so the net amount of potential energy increases.
- (+) $\Delta H$
Lesson 2: Potential Energy Diagrams

**HOW A REACTION HAPPENS:**

Example: $H_2 + I_2 \rightarrow 2HI$

First, the bonds within the $H_2$ and $I_2$ molecules need to _______________.
Remember BARF – energy must be ____________ and increases the PE of the system.

THEN, the new bonds between the H and I atoms can _____________.
Remember BARF – this ______________ energy and decreases the PE of the system.

Per Table I, this reaction is ________-thermic. So PE of products is ____________ than PE of reactants.

The “peak” is called *the Activated Complex* – the highest energy, least bonded part of the reaction mechanism.
Lesson 2: Potential Energy Diagrams

IN CLASS NOTES

Parts of the Potential Energy Diagram

***You must be able to label these so label the diagrams in your notes as we go through this!!!

ACTIVATION ENERGY:

__________________ energy required for a reaction to occur (energy needed to get over the hill)

ACTIVATED COMPLEX:
Highest energy point of reaction
- Temporary
- Where bonds are broken and reformed
HEAT OF REACTANTS/PRODUCTS:
Amount of ________________ possessed by the reactants or products

HEAT OF REACTION $\Delta H$
- Amount of energy lost/gained in a reaction
- SUBTRACT heat of products minus reactants

\[ \Delta H = \frac{____________________}{\text{heat of products}} - \frac{____________________________}{\text{heat of reactants}} \]
A Catalyst lowers the __________________ by providing an __________________.

EFFECTS OF ADDING A CATALYST
- LOWERS the __________________ (energy needed to start the reaction)
- Reaction occurs faster
Putting it together... Constructing a PE Diagram

**EXOTHERMIC**

**ENDOTHERMIC**
Objective:
- Describe the factors that influence the rate of a reaction.
- Use collision theory to explain how various factors, such as temperature, surface area, and concentration influence the rate of a reaction.

Effective Collisions:
- In order for a reaction to occur, reactant PARTICLES MUST COLLIDE (effectively) with the following:
  1. Proper amount of ______________
  2. Proper ______________________

Example: $\text{H}_2 + \text{I}_2 \rightarrow 2\text{HI}$

Factors Affecting Reaction Rate
- SIX FACTORS that affect the rate of reaction by changing the number of effective collisions that take place between particles.
- The MORE ____________________________, THE FASTER THE REACTION!

1. TYPE OF SUBSTANCE:
   - ___________ substances react ________________
     - □ Easily break into IONS when you dissolve them. NO BONDS BROKEN (no electrons change hands)
     - Example: $\text{AgNO}_3 (s) \rightarrow \text{Ag}^+ + \text{NO}_3^-$
   - ___________ substances react ________________
     - □ Requires more energy/time to break bonds
     - Example: $\text{H}_2 (g) + \text{I}_2 (g) \rightarrow 2\text{HI} (g)$

Note: You can’t change the nature of the reactants.
Lesson 3: Effective Collisions and Reaction Rates

2. **PRESSURE (GASES ONLY)**
   
   ______________ pressure, ______________ reaction rate (affects GASES ONLY!)
   
   Due to an increase in density – more particles in a given area so more chance for collisions.

3. **CONCENTRATION:**
   
   ______________ Concentration ______________ reaction rate (speed)
   
   More particles increases chance of effective collisions

4. **TEMPERATURE:**
   
   ______________ temperature ______________ Reaction Rate:
   
   a. Increases # of collisions
   
   b. Reactants have more energy when colliding (so more effective collisions)

5. **SURFACE AREA:**
   
   ______________ in surface area ______________ the reaction rate.
   
   a. Due to more exposed particles that can react (more effective collisions)

6. **CATALYST:**

   Substance that ______________ reaction rate without being consumed in the reaction.
   
   Lower the ______________, so more collisions are effective (have enough energy to get over the hill).
Lesson 3: Effective Collisions and Reaction Rates

IN CLASS VIDEO & DEMONSTRATIONS:
Consider the five ways to increase the rate of a reaction. What was the “dating” analogy from the video? Which in-class demonstration(s) applied?

<table>
<thead>
<tr>
<th>Reaction Rate affected by</th>
<th>To increase rate....</th>
<th>School Dance Analogy</th>
<th>Demonstration Used</th>
</tr>
</thead>
<tbody>
<tr>
<td>Nature of Reactants</td>
<td></td>
<td>No analogy (can’t change this!)</td>
<td></td>
</tr>
<tr>
<td>Pressure (gas only)</td>
<td></td>
<td></td>
<td><em>No Demonstration</em></td>
</tr>
<tr>
<td>Concentration</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Temperature</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Surface Area</td>
<td></td>
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<tr>
<td>Catalyst</td>
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</tbody>
</table>
Lesson 4: Spontaneous Reactions

**Objective:**
- Describe the conditions that favor a reaction (the reaction will be spontaneous).
- Define Entropy
- Determine if entropy increases or decreases in a phase change or reaction

NATURE IS LAZY... favors _______________ energy (_____________ thermic reactions)

NATURE IS MESSY... favors _______________ entropy (randomness or disorder)

**Entropy:**

Entropy and phase changes:

![Solid](image1.png) ![Liquid](image2.png) ![Gas](image3.png)

What about aqueous?

**IF PROCESS IS EXOTHERMIC AND ENTROPY INCREASES, IT WILL ALWAYS BE SPONTANEOUS!**
REVERSIBLE vs IRREVERSIBLE REACTIONS

All physical changes are reversible...
   Examples:

Chemical reactions CAN be under certain circumstances.
However, many are not...
   Examples:

If a process is reversible (under same conditions), it can reach equilibrium...

Example, a substance at its melting point can be in equilibrium between solid and liquid phases.

EQUILIBRIUM REQUIRES A ______________ SYSTEM.

WHEN AT EQUILIBRIUM:

■ The RATE of the forward reaction is __________ to the rate of the reverse reaction
■ The amounts (concentrations) of reactants and products remain ________________

■ Equilibrium is Dynamic (in constant motion)
■ Equilibrium is represented by double arrow

A + B ⇌ C + D
Lesson 5: Equilibrium

Types of Equilibrium (all occur in closed systems)

1. Phase Equilibrium:
Rate of forward phase change equals rate of reverse phase change, e.g.,
\[ \text{Rate}_{\text{vaporizing}} = \text{Rate}_{\text{condensing}} \]

Ex. \[ \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_2\text{O}(g) \]

Water is vaporizing at the same rate it is condensing
SYSTEM MUST BE AT THE _________________or___________________ POINT TO ACHIEVE CHANGE EQUILIBRIUM

If the system is closed during a phase change (no more energy in or out) it will achieve equilibrium

2. Solution Equilibrium:
The rate of__________________equals the rate of__________________.

Ex.

CAN ONLY OCCUR WHEN SOLUTION IS ____________________ with excess solute.
3. **Chemical Equilibrium**: The rate of the forward reaction is equal to the rate of the reverse reaction

*Ex.* The Haber Process
Lesson 6: Changing Equilibrium

Objective:
- Determine the shift in equilibrium when a stress is placed on a system
- Determine the change in concentration when a stress is placed on a system

A SYSTEM AT EQUILIBRIUM IS IN BALANCE – number of ___________________ hence the ___________________ is ______________ for forward and reverse reactions.

LE CHATELIER’S PRINCIPLE:
- If a system at equilibrium is subjected to a ______________, the equilibrium will __________ in the direction that relieves that stress
- For example, if one changes the _____________________, it will change the ________ of the forward or reverse ______________. When no longer __________, equilibrium will adjust until it is again __________.
- Causes a change in concentration of both the reactants and products until the equilibrium is re-established.

TYPES OF STRESS:
- Change in Concentration
- Change in temperature
- Change in pressure (gases only)

CHANGE IN CONCENTRATION:
Affects the rate of the reaction

$$N_2(g) + 3H_2(g) \leftrightarrow 2NH_3(g)$$

Ex 1: Add $N_2(g)$  
SO  Fwd reaction rate ____________:
consumes more $N_2$ and $H_2$ (concentrations go down);
produces more $NH_3$ (concentration goes up)

$$N_2(g) + 3H_2(g) \leftrightarrow 2NH_3(g)$$
Lesson 6: Changing Equilibrium

\[ \text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \leftrightarrow 2\text{NH}_3(\text{g}) \]

Ex 2: Remove \( \text{H}_2(\text{g}) \) SO Fwd reaction rate _________________:
Less \( \text{N}_2 \) and \( \text{H}_2 \) are reacting (concentrations increase)
Lesss \( \text{NH}_3 \) produced (concentration goes down)

\[ \text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \leftrightarrow 2\text{NH}_3(\text{g}) \]

CHANGE IN TEMPERATURE:
Increasing temperature favors _________________ reaction.

Ex: \[ \text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \leftrightarrow 2\text{NH}_3(\text{g}) + 91.8 \text{kJ} \]

Decreasing temperature favors _________________ reaction.

Ex: \[ \text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \leftrightarrow 2\text{NH}_3(\text{g}) + 91.8 \text{kJ} \]

CHANGE IN PRESSURE (GASES ONLY):
PICTURE BOTH Reactants & Products IN SAME CONTAINER....

Increasing pressure favors the production of _______ moles of gas

Ex: \[ \text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \leftrightarrow 2\text{NH}_3(\text{g}) \]

___ + ___ VS ___ moles of gas

More products means less reactants (2 particles created from every 4) so fewer particles of gas in the container....
Lesson 6: Changing Equilibrium

Decreasing pressure favors _________ moles of gas

EX: N$_2$(g) + 3H$_2$(g) $\leftrightarrow$ 2NH$_3$(g)

More reactants means less products (4 particles created from every 2) so more particles of gas in the container....

ADDING A CATALYST:

- Both the forward and reverse reactions will be affected equally (both will speed up) so _________________ to the equilibrium.

APPLYING L’Chatelier’s Principle:

Why does my soda go flat...

CO$_2$(aq)$\leftrightarrow$ CO$_2$(g)

In the closed bottle it is at _______________________.

When I open it and reseal it, I both allow CO$_2$(g) to escape AND reduce the pressure of the system

What is the effect of removing the CO$_2$(g)?

What is the effect of reducing the pressure?

What happens to the concentration of CO$_2$(aq)?
Lesson 6: Changing Equilibrium

\[ \text{CO}_2(\text{aq}) \rightleftharpoons \text{CO}_2(\text{g}) \]

What happens if I heat it up?
What happens to solubility of a gas with increasing temperature?

Which reaction is favored (forward or reverse)?

Based on your observations, is dissolving CO$_2$ (g) in water exothermic or endothermic?

\[ \text{CO}_2(\text{aq}) + \text{_________} \rightleftharpoons \text{CO}_2(\text{g}) \]

**EXAMPLE:**

\[ 2\text{CO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 566 \text{ kJ} \]

If concentration of CO is increased:

1. What direction will the equilibrium shift?

2. What happens to the concentration of CO$_2$?

3. What happens to the concentration of O$_2$?