Chemistry 12
Unit 2- Equilibrium
Notes

It's important to know that many chemical reactions are reversible. That is:

\[
\text{Reactants} \rightarrow \text{Products} \quad \text{or} \quad \text{Reactants} \leftarrow \text{Products}
\]

\[
\text{Reactants form Products} \quad \text{or} \quad \text{Products form Reactants}
\]

For example, under certain conditions, one mole of the colourless gas N\textsubscript{2}O\textsubscript{4} will decompose to form two moles of brown NO\textsubscript{2} gas:

\[
\text{N}_2\text{O}_4 \rightarrow 2 \text{NO}_2
\]

colourless \quad \text{brown}

Under other conditions, you can take 2 moles of brown NO\textsubscript{2} gas and change it into one mole of N\textsubscript{2}O\textsubscript{4} gas:

\[
\text{N}_2\text{O}_4 \leftarrow 2 \text{NO}_2
\]

colourless \quad \text{brown}

In other words, this reaction, as written may go forward or in reverse, depending on the conditions.

If we were to put some N\textsubscript{2}O\textsubscript{4} in a flask, the N\textsubscript{2}O\textsubscript{4} molecules would collide with each other and some of them would break apart to form NO\textsubscript{2}.

This process is indicated by the forward reaction:

\[
\text{N}_2\text{O}_4 \rightarrow 2 \text{NO}_2
\]
Once this has happened for awhile, there is a build up of NO₂ molecules in the same flask.

Once in awhile, two NO₂ molecules will collide with each other and **join** to form a molecule of N₂O₄!

![Diagram of NO₂ molecules colliding to form N₂O₄](image)

This process, as you might have guessed is indicated by the **reverse reaction**:

\[ \text{N}_2\text{O}_4 \rightleftharpoons 2 \text{NO}_2 \]

Two things you'll have to realize is that *as long as there is N₂O₄ present, the forward reaction will keep on happening* and *as long as there is NO₂ present, the reverse reaction will keep on happening!*

Also, you must keep in mind that all these molecules are mixed **in the same container!**

At one particular time a molecule of N₂O₄ might be breaking up, and **at the same time** two molecules of NO₂ might be joining to form another molecule of N₂O₄! So here's an important thing to understand:

**In any reversible reaction, the forward reaction and the reverse reaction are going on at the same time!**

This is sometimes shown with a **double arrow**:

\[ \text{N}_2\text{O}_4 \rightleftharpoons 2 \text{NO}_2 \]

The double arrow means that both the forward and reverse reaction are happening at the same time.

Just a little comment here. The word "happening" has a similar meaning to the word "dynamic ". Just remember that!

*******************************************************************************
Now what we're going to do is look at how the rate of the forward reaction changes if we put some pure N\textsubscript{2}O\textsubscript{4} in a flask:

If we put some pure N\textsubscript{2}O\textsubscript{4} in a flask (No NO\textsubscript{2} yet!), there will be a high concentration of N\textsubscript{2}O\textsubscript{4}. That is, there will be lots of N\textsubscript{2}O\textsubscript{4} molecules to collide with each other. So at the beginning of our little experiment, (which we will call "time 0") the rate of the forward reaction is quite fast.

\[ \text{N}_2\text{O}_4 \rightarrow 2 \text{NO}_2 \]

So if we were to make a graph of the rate of the forward reaction vs. time, the graph might start out something like this:

![Graph](image)

OK, now you might ask: "Why does the rate of the forward reaction go down?"

Well, if you recall Unit 1, as the forward reaction proceeds:

\[ \text{N}_2\text{O}_4 \rightarrow 2 \text{NO}_2 \]

the N\textsubscript{2}O\textsubscript{4} is used up and so it's concentration goes down. Also, you must remember that if the concentration of a reactant goes down, there is less chances of collisions and the rate of the reaction decreases.

As the reaction continues, the slower rate will use up N\textsubscript{2}O\textsubscript{4} more slowly, so the [N\textsubscript{2}O\textsubscript{4}] will not decrease so quickly and therefore the rate will not decrease quite as quickly. (Read the last sentence over a couple of times and make sure it makes sense to you!) For those "graph wise" people, you will probably guess that this means that the slope of the line on the graph gets more gradual. The rest of the graph might look something like this:
NOW, it's time we consider the rate of the reverse reaction.

As you might recall, when we have a container full of pure $\text{N}_2\text{O}_4$, initially there is no $\text{NO}_2$ in the container. Since there is no $\text{NO}_2$, there are no $\text{NO}_2$ molecules to collide with each other, and the rate of the reverse reaction is zero.

But of course, as time goes on, $\text{NO}_2$ is formed from the forward reaction ($\text{N}_2\text{O}_4 \rightarrow 2 \text{NO}_2$) so in a short time, some $\text{NO}_2$ molecules can start colliding and the reverse reaction will begin:

\begin{equation}
\text{N}_2\text{O}_4 \leftrightarrow 2 \text{NO}_2
\end{equation}

As MORE $\text{NO}_2$ is formed by the forward reaction, the rate of the reverse reaction gradually increases. Now, for you "graph buffs", the graph of the Rate of the reverse reaction vs. Time might look like this:
OK. Now, let’s look at the graph for the forward rate and the reverse rate together:

![Graph showing forward and reverse rates over time.]

NOW, focus your attention on the graph at "Time = 4 minutes". You will notice that at this point:

\[
\text{the rate of the forward reaction} = \text{the rate of the reverse reaction}
\]

At this point, NO\(_2\) is being used up \textit{at the same rate} that it is being formed:

\[
\text{N}_2\text{O}_4 \rightleftharpoons 2\text{NO}_2
\]

Because this is so, you should be able to convince yourself that the [NO\(_2\)] is no longer changing!

Because the reverse rate is equal to the forward rate, \text{N}_2\text{O}_4 is being formed \textit{at the same rate} it is being used up. So, also the [N\(_2\)O\(_4\)] is no longer changing either!

Can you predict what will happen to the graph after 4.0 minutes?

YOU GUESSED IT! The rates of the forward reaction and the reverse reaction, no longer change because the [N\(_2\)O\(_4\)] is constant and the[NO\(_2\)] is also constant. The graph will look like this:
The situation happening from 4.0 minutes on in this graph has a special name and a special significance. At this point, the system (meaning the container, the N₂O₄ and the NO₂) is said to be at equilibrium. To describe it even more precisely, we can say that we have reached a state of dynamic equilibrium.

Here are some things that you must understand about dynamic equilibrium:

1. The reaction has not stopped!

2. The forward and the reverse reaction continue to take place, but their rates are equal so there are no changes in concentrations of reactants or products. (The forward and reverse reactions are said to be "balanced") eg. for the reaction:

   \[ \text{N}_2\text{O}_4 \rightleftharpoons 2 \text{NO}_2 \]

   for each N₂O₄ molecule that breaks up to form two NO₂ molecules, two other NO₂ molecules combine to form another N₂O₄, so we don't see individual molecules reacting.

3. As far as we can see from the "outside", there appears to be nothing happening. All observable properties are constant. These include the concentrations of all reactants and products, the total pressure, colour, temperature etc.

4. If no changes were made in conditions and nothing is added or taken away, a system at equilibrium would remain that way forever, the forward and reverse reactions "ticking away", but balanced so that no observable changes happen.
Here are a couple of other things to consider before we summarize everything:

1. Changing the temperature can alter the rates of the reactions at equilibrium. This could "throw off" the balance. So, for a system at equilibrium, the temperature must remain constant and uniform throughout the system.

2. Letting material into or out of the system will affect rates so a system at equilibrium is a closed system.

3. Again, consider the equilibrium reaction: $N_2O_4 \rightleftharpoons 2 NO_2$

   In the example that we did to construct the graphs, we had started with pure $N_2O_4$ and no $NO_2$. The forward reaction rate was high at the start, but the reverse reaction rate eventually "caught up", the rates became equal and equilibrium was established. Can you guess what would happen if we had started with pure $NO_2$ instead (no $N_2O_4$)? The reverse rate would start out high and the forward rate, zero. In time, the forward rate would "catch up". When the rates became equal, again equilibrium would be established.

We can summarize all this by saying that the equilibrium can be approached from the left (starting with reactants) or from the right (starting with products).

Just a little term before we summarize: The word macroscopic means large scale or visible or observable. (The opposite is microscopic, which means too small to see eg. molecular level). Some macroscopic properties are total pressure, colour, concentrations, temperature, density etc.

Alright, let's summarize:

<table>
<thead>
<tr>
<th>Characteristics of a System at Dynamic Equilibrium</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. The rate of the forward reaction = The rate of the reverse reaction</td>
</tr>
<tr>
<td>2. Microscopic processes (the forward and reverse reaction) continue in a balance which yields no macroscopic changes. (so nothing appears to be happening.)</td>
</tr>
<tr>
<td>3. The system is closed and the temperature is constant and uniform throughout.</td>
</tr>
<tr>
<td>4. The equilibrium can be approached from the left (starting with reactants) or from the right (starting with products).</td>
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</table>
**Enthalpy**

Enthalpy is "The heat content of a system." Another way to think of enthalpy is as "Chemical Potential Energy".

Any change in the Potential Energy of a system means the same thing as the "Enthalpy Change". The symbol for Enthalpy is "\( H \)". Therefore the "change in Enthalpy" of a chemical reaction is called "\( \Delta H \)". In Chemistry 12, a Potential Energy Diagram is the same thing as an "Enthalpy Diagram".

In an **Exothermic** Reaction (\( \Delta H \) is negative), the **Enthalpy** is decreasing.

In an **Endothermic** Reaction (\( \Delta H \) is positive), the **Enthalpy** is increasing.
If the "Heat Term" is written right in the equation. (a "thermochemical equation").

If the heat term is on the left side, it means heat is being used up and it's endothermic.

If the heat term is on the right side, heat is being released and it's exothermic.

Look at the following examples:

1. A + B \rightleftharpoons C + D \quad \Delta H = -24 \text{ kJ} \quad \text{is exothermic so enthalpy is decreasing.}

2. X + Y \rightleftharpoons Z \quad \Delta H = 87 \text{ kJ} \quad \text{is endothermic so enthalpy is increasing.}

3. E + D \rightleftharpoons F + 45 \text{ kJ} \quad \text{is exothermic so enthalpy is decreasing.}

4. G + J + 36 \text{ kJ} \rightleftharpoons L + M \quad \text{is endothermic so enthalpy is increasing.}

Systems will tend toward a state of lower potential energy if nothing else is acting upon them.

In Chemistry, we are interested in is chemical potential energy, otherwise known as enthalpy!

Chemical systems will tend toward a state of minimum enthalpy if sufficient activation energy is available and no other factors are considered.

Another way of stating this might be:

A chemical reaction will favour the side (reactants or products) with minimum enthalpy if no other factors are considered.
Thus for an *exothermic reaction*, if no other factors are considered:

![Diagram of Enthalpy vs Progress of Reaction]

Here, the Products have lower enthalpy than the Reactants so the reaction tends to "favour the products". In other words, if the reactants are mixed, they will tend to form products *spontaneously* (without outside assistance) rather than remain as reactants.

The *products will be favoured* because the products have *minimum enthalpy*. In other words, there is a natural tendency here for reactants to *spontaneously* form products.

In an *endothermic reaction*, the __________________________ have minimum enthalpy, so the __________________________ will be favoured. In other words, if the reactants are mixed they will (tend to remain as reactants / spontaneously form products) ____________________
Let's look at a diagram for an *endothermic reaction*:

![Diagram of an endothermic reaction](image)

In the case of an *endothermic* reaction, the enthalpy of the *Reactants* is *lower* than the enthalpy of the *Products*. Since chemical systems favour a state of *minimum enthalpy*, the *Reactants are favoured* in this case. In other words if the reactants are mixed, they will tend to remain as reactants rather than forming products.

1. Tell whether each of the following is *endothermic* or *exothermic* and state which has *minimum enthalpy*, the *reactants* or the *products*:

   a. $\text{PCl}_5(g) \rightleftharpoons \text{Cl}_2(g) + \text{PCl}_3(g)$ $\Delta H = 92.5 \text{ kJ}$
      
      _______thermic and the ________________ have *minimum enthalpy*.

   b. $2\text{NH}_3(g) + 92.4 \text{ kJ} \rightleftharpoons \text{N}_2(g) + 3\text{H}_2(g)$
      
      _______thermic and the ________________ have *minimum enthalpy*.

   c. $\text{CO}(g) + 3\text{H}_2(g) \rightleftharpoons \text{CH}_4(g) + \text{H}_2\text{O}(g) + 49.3 \text{ kJ}$
      
      _______thermic and the ________________ have *minimum enthalpy*.

   d. $\text{Cl}_2(g) \rightleftharpoons \text{Cl}_2(\text{aq})$ $\Delta H = -25 \text{ kJ}$
      
      _______thermic and the ________________ have *minimum enthalpy*. 
2. When no other factors are considered, a reaction will move in such a way (left or right) in order to achieve a state of \( \text{__________________________________________} \) enthalpy.

3. Given the equation: \( 2\text{NH}_3(g) + 92.4 \text{ kJ} \rightleftharpoons \text{N}_2(g) + 3\text{H}_2(g) \)

   If only the \textit{enthalpy} is considered, the (reactant / products) \( \text{__________________________________________} \) will be favoured at equilibrium.

4. Given the equation: \( \text{Cl}_2(g) \rightleftharpoons \text{Cl}_2(aq) \quad \Delta H = -25 \text{ kJ} \)

   If only the \textit{enthalpy} is considered, the (reactant / products) \( \text{__________________________________________} \) will be favoured at equilibrium.

5. If the reaction: \( \text{CO}(g) + 3\text{H}_2(g) \rightleftharpoons \text{CH}_4(g) + \text{H}_2\text{O}(g) + 49.3 \text{ kJ} \)

   was proceeding to the \textit{right}, the enthalpy would be \( \text{_______________} \)ing. Is this a \textit{favourable} change? \( \text{___________} \).

6. If the reaction: \( \text{PCl}_5(g) \rightleftharpoons \text{Cl}_2(g) + \text{PCl}_3(g) \quad \Delta H = 92.5 \text{ kJ} \)

   was proceeding to the \textit{right}, the enthalpy would be \( \text{_______________} \)ing. Is this a \textit{favourable} change? \( \text{___________} \).

7. If the reaction: \( \text{Cl}_2(g) \rightleftharpoons \text{Cl}_2(aq) \quad \Delta H = -25 \text{ kJ} \)

   was proceeding to the \textit{right}, the enthalpy would be \( \text{_______________} \)ing. Is this a \textit{favourable} change? \( \text{___________} \).

8. If the reaction: \( 2\text{NH}_3(g) + 92.4 \text{ kJ} \rightleftharpoons \text{N}_2(g) + 3\text{H}_2(g) \)

   was proceeding to the \textit{right}, the enthalpy would be \( \text{_______________} \)ing. Is this a \textit{favourable} change? \( \text{___________} \).
As you can see by looking at the exercises above, there are two ways of looking at what happens to the enthalpy:

**If the reaction is exothermic, the products have minimum enthalpy and the formation of products (move toward the right) is favourable.**

**If the reaction is endothermic, the reactants have minimum enthalpy and the formation of products (move toward the right) is unfavourable. In this case the formation of reactants (move toward the left) is favourable.**

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Now, consider the simple melting of water:

\[ \text{H}_2\text{O} (s) + \text{heat} \rightleftharpoons \text{H}_2\text{O} (l) \]  

(The subscript (s) stands for solid) (the subscript (l) stands for liquid)

If we were to look at only the enthalpy in this process, you can see that the reactant (\( \text{H}_2\text{O} (s) \)) would have minimum enthalpy and would be favoured. So all of the water in the universe should exist only as a solid! (It would not be favourable for water to exist as a liquid!) We would all be frozen solid!!!!

The answer to this problem lies in looking at another factor that governs equilibrium. That factor is called **entropy** (or randomness or disorder)

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**Entropy**

**Entropy simply means disorder, or lack of order.**
In Grade 8, you probably learned about the arrangement of molecules in solids, liquids and gases.

The particles in a **solid** are very close together and very **ordered**. A solid has very **low entropy**. (very little disorder)

The particles in a **liquid** are fairly close together and they are not as ordered as they are in the solid. A liquid has **higher entropy** than a solid, but **less entropy** than a gas.

The particles of a **gas** are very far apart and they are moving **randomly** (all different directions - any old way!). They have very little order. A **gas** has **very high entropy**.

So we can summarize by saying that:

\[
\text{Entropy of a Solid} < \text{Entropy of a Liquid} < \text{Entropy of a Gas}
\]

We can look at a chemical equation with subscripts showing the phases and tell which has **maximum entropy**, the **reactants** or the **products**.

In other words, they can look at an equation and tell whether **entropy** is **increasing** or **decreasing** as the reaction **proceeds to the right**.

In the following examples, the **entropy is increasing** (or the **products** have **greater entropy**):

1. There is a **gas** (or gases) on the **right**, when there are **no gases** on the **left** of the equation:

   \[
   \text{CaCO}_3(\text{s}) + 2 \text{HCl}(\text{aq}) \rightleftharpoons \text{CaCl}_2(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})
   \]
   a gas is formed on the right.

2. When there are **gases on both sides**, the **products** have **greater entropy** when there are **more moles of gas on the right** (add up coefficients of gases on left and right.):

   \[
   4\text{NH}_3(\text{g}) + 5\text{O}_2(\text{g}) \rightleftharpoons 4\text{NO}(\text{g}) + 6\text{H}_2\text{O}(\text{g})
   \]
   There are \((4 + 5) = 9\) moles of gas on the left
   There are \((4 + 6) = 10\) moles of gas on the right.
Another way to look at the last example is to say that:

"The side with the greater number of moles of gas has the greatest entropy."

3. When a solid dissolves in water, the products (the aqueous solution of ions) have greater entropy. This makes sense because:

The ions in a Solid are very ordered. They have low entropy.

When dissolved in water, the ions are separated and surrounded by water molecules. The ions are much more spread out and disordered. The entropy of an aqueous solution is higher than that of a solid.

So in order of lowest to highest entropy:

**Solids < Liquids < Aqueous solutions < Gases < More moles of Gas**

Here are few exercises for you:

9. For each of the following, decide whether the **reactants** or the **products** have greater entropy:

   a) $\text{I}_2(\text{s}) \rightleftharpoons \text{I}_2(\text{aq})$  The __________________________ have greater entropy.
b) \[ 2\text{NH}_3(g) \rightleftharpoons \text{N}_2(g) + 3\text{H}_2(g) \]
The ___________________________________________have greater entropy.

c) \[ \text{NH}_3(g) \rightleftharpoons \text{NH}_3(aq) \]
The ___________________________________________have greater entropy.

d) \[ \text{CO}(g) + \text{Cl}_2(g) \rightleftharpoons \text{COCl}_2(g) \]
The ___________________________________________have greater entropy.

e) \[ \text{MgCO}_3(s) + 2\text{HCl}(aq) \rightleftharpoons \text{MgCl}_2(aq) + \text{H}_2\text{O}(l) + \text{CO}_2(g) \]
The ___________________________________________have greater entropy.

If you have any questions about these, check with your teacher!

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Remember: \[ \text{H}_2\text{O}(s) + \text{heat} \rightleftharpoons \text{H}_2\text{O}(l) \]

We decided that all the \( \text{H}_2\text{O} \) in the universe should remain as a solid because \( \text{H}_2\text{O}(s) \) has lower enthalpy than \( \text{H}_2\text{O}(l) \) and nature favours a state of minimum enthalpy.

Well, now we can explain why there is some liquid water in the universe (lots of it):

\( \text{H}_2\text{O}(l) \) has **higher entropy** than \( \text{H}_2\text{O}(s) \)

There is a natural tendency in nature toward maximum disorder or maximum entropy!

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**Chemical systems will tend toward a state of maximum entropy if no other factors are considered.**

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Another was of stating this might be:

**A chemical reaction will favour the side (reactants or products) with maximum entropy if no other factors are considered.**
Remember, the other factor which controlled reactions was **enthalpy**. (chemical potential energy).

Also remember that:

**Chemical systems will tend toward a state of minimum enthalpy if sufficient activation energy is available and no other factors are considered.**

or

**A chemical reaction will favour the side (reactants or products) with minimum enthalpy if no other factors are considered.**

You figure out which has the most **enthalpy** (reactants or products) by looking at the ΔH or the heat term.

Also, remember that you can figure out which has the more **entropy** (reactants or products) by looking at the subscripts which represent the phases.

Also, we can combine the rules about "natural tendencies" to come up with this:

**In nature, there is a tendency toward minimum enthalpy and maximum entropy.**

Now, let's consider this process again:

\[ \text{H}_2\text{O}_\text{s}, + \text{heat} \Leftrightarrow \text{H}_2\text{O}_\text{l} \]

The two tendencies are said to "**oppose each other**" in this case:

The tendency toward **minimum enthalpy** would favour the reactant ! (since you have to add heat energy to \( \text{H}_2\text{O}_\text{s} \) to get \( \text{H}_2\text{O}_\text{l} \). \( \text{H}_2\text{O}_\text{s} \) has **minimum enthalpy**)

In this case the tendency toward **maximum entropy** would tend to favour the product. (A liquid has more **entropy** (disorder) than a **solid**)

We say that:

When the two tendencies **oppose each other** (one favours reactants, the other favours products), the reaction will **reach a state of equilibrium**.

That is, there will be some reactants and some products present. The relative amounts of each depends on conditions like temperature, pressure, concentration etc.
Since this is the case with \( H_2O(s) + \text{heat} \Leftrightarrow H_2O(l) \), there is some solid water and some liquid water in the universe. (In other words, there is a state of equilibrium) Which one is present in the greater amount is determined largely by the temperature.

Now, let's consider another simple process: A glass bottle is knocked down from a high shelf onto a concrete floor and the glass shatters:

**Bottle on a high shelf → Thousands of pieces of glass on a concrete floor**

The bottle falls down and not up! This happens because there is a natural tendency toward minimum gravitational potential energy (like minimum enthalpy in chemistry).

In other words the tendency toward minimum gravitational potential energy favours the products (the low bottle rather than the high)

(The person who knocked the bottle off of the shelf was simply supplying the "activation energy")

Remember that the bottle broke into thousands of pieces when it hit the concrete. The broken pieces of glass have more disorder (entropy) than the bottle, so in this process, the tendency toward maximum entropy also favours the products!

There is no "equilibrium" here when the process is finished. That bottle has completely fallen down and it is all broken. (This bottle is no longer on the shelf and it is no longer an "unbroken bottle")

We can summarize what happened here:

**Processes in which both the tendency toward minimum enthalpy and toward maximum entropy favour the products, will go to completion.**

(i.e. All reactants will be converted into products. There will be no reactants left once the process is finished!)
Here's an example of a chemical reaction in which this happens:

\[ 2K(s) + 2H_2O(l) \rightleftharpoons 2KOH(aq) + H_2(g) + \text{heat} \]

This process is *exothermic* (the heat term is on the right) so the *products have lower enthalpy*.

The tendency toward *minimum enthalpy* favours the *products*.

There is a mole of gas on the right (\(H_2(g)\)) and no gases in the reactants. Therefore, the *products have greater entropy*.

The tendency toward *maximum entropy* favours the *products*.

Since both tendencies favour the products, this reaction **will go to completion**.

That is, all of the reactants (assuming you have the correct mole ratios eg. 2 moles of K to 2 moles of \(H_2O\)) will be converted to products.

If one reactant is in excess, the *limiting reactant* will be completely consumed.

So, if you put a little bit of potassium in a beaker of water, the reaction will keep going until all of the potassium is used up. There will be *no* potassium left once the reaction is complete.

In other words, the reverse reaction does not occur!

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Let's consider one more process:

\[ 2KOH(aq) + H_2(g) + \text{heat} \rightleftharpoons 2K(s) + 2H_2O(l) \]

In this case, the tendency toward *minimum enthalpy* favours the *reactants*, and the tendency toward *maximum entropy* also favours the *reactants*.

Processes in which **both** the tendency toward *minimum enthalpy* and toward *maximum entropy* favour the *reactants*, will **not occur at all!**

(i.e. None of the reactants will be converted into products. There will be no products formed!)

**NOTE:** This would be like thousands of pieces of glass spontaneously sticking together, forming a bottle and jumping up onto a high shelf! This does not occur at all. (At least I've never seen it happen!)
To summarize:

When the two tendencies oppose each other (one favours reactants, the other favours products), the reaction will reach a state of equilibrium.

Processes in which both the tendency toward minimum enthalpy and toward maximum entropy favour the products, will go to completion.

Processes in which both the tendency toward minimum enthalpy and toward maximum entropy favour the reactants, will not occur at all!

10. For each of the following reactions decide which has minimum enthalpy (reactants or products), which has maximum entropy (reactants or products), and if the reactants are mixed, what will happen? (go to completion/ reach a state of equilibrium/not occur at all).

   a) $\text{PCl}_3(g) + \text{Cl}_2(g) \rightleftharpoons \text{PCl}_5(g) ; \Delta H = -92.5 \text{ kJ}$

   The _____________________________ has/have minimum enthalpy.

   The _____________________________ has/have maximum entropy.

   If PCl$_3$ and Cl$_2$ are put together, what should happen? (go to completion/ reach a state of equilibrium/not occur at all)

   ________________________________________________________________

   b) $2\text{NO}_2(g) \rightleftharpoons \text{N}_2\text{O}_4(g) + \text{energy}$

   The _____________________________ has/have minimum enthalpy.

   The _____________________________ has/have maximum entropy.

   If NO$_2$ was put in a flask, what should happen? (go to completion/ reach a state of equilibrium/not occur at all)

   ________________________________________________________________

   c) $\text{P}_4(s) + 6\text{H}_2(g) + 37 \text{ kJ} \rightleftharpoons 4\text{PH}_3(g)$

   The _____________________________ has/have minimum enthalpy.
The _________________________________ has/have maximum entropy.

If \( \text{P}_4(\text{s}) \) and \( 6\text{H}_2(\text{g}) \) was put in a flask, what should happen? (go to completion/ reach a state of equilibrium/not occur at all)

___________________________________________________________________

\( \text{d) } 2\text{PbO}_\text{(s)} + 4\text{NO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{Pb(NO}_3)_2(\text{s}) \); \( \Delta \text{H} = -597 \text{ kJ} \)

The _________________________________ has/have minimum enthalpy.

The _________________________________ has/have maximum entropy.

If \( \text{PbO}_\text{(s)} \) and \( \text{NO}_2(\text{g}) \) were put in a flask, what should happen? (go to completion/ reach a state of equilibrium/not occur at all)

___________________________________________________________________

**More Questions:**

1. What is meant by **enthalpy**? _____________________________________________

_______________________________________________________________________

2. What is meant by **entropy**? _____________________________________________

3. In an **endothermic reaction**, the ________________________________ have minimum enthalpy.

4. In an **exothermic reaction**, the ________________________________ have minimum enthalpy.

5. Arrange the following in order from **least entropy** to **greatest entropy**:
   a) liquids b) gases c) aqueous solutions d) solids
   
   ___________ < ___________ < ___________ < ___________

6. There is a natural tendency toward ____________________________ enthalpy and ____________________________ entropy.

7. A process in which **entropy increases** and **enthalpy decreases** will
   (go to completion/ reach a state of equilibrium/not occur at all) _______________________________

8. A process in which **entropy increases** and **enthalpy increases** will
   (go to completion/ reach a state of equilibrium/not occur at all) _______________________________
9. A process in which *entropy decreases* and *enthalpy decreases* will
   (go to completion/ reach a state of equilibrium/not occur at all) _______________________________

10. A process in which *entropy decreases* and *enthalpy increases* will
   (go to completion/ reach a state of equilibrium/not occur at all) _______________________________

11. A process in which *both the enthalpy and entropy trends favour reactants* will
   (go to completion/ reach a state of equilibrium/not occur at all) _______________________________

12. A process in which *both the enthalpy and entropy trends favour products* will
   (go to completion/ reach a state of equilibrium/not occur at all) _______________________________

13. A process in which *the enthalpy and entropy trends oppose each other* will
   (go to completion/ reach a state of equilibrium/not occur at all) _______________________________

14. In each of the following, state which has the *maximum entropy*, (reactants or products)
   a) \( \text{C(s)} + \text{O}_2(\text{g}) \rightleftharpoons \text{CO}_2(\text{g}) \) __________________________________________
   b) \( 2\text{Al(s)} + 6\text{HCl(aq)} \rightleftharpoons 3\text{H}_2(\text{g}) + 2\text{AlCl}_3(\text{aq}) \) _____________________________
   c) \( 2\text{SO}_3(\text{g}) \rightleftharpoons 2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \) ______________________________________
   d) \( \text{HCl}(\text{g}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \) ______________________________________
   e) \( \text{KOH(s)} \rightleftharpoons \text{K}^+(\text{aq}) + \text{OH}^-(\text{aq}) \) ______________________________________

15. For each of the following reactions decide which has *minimum enthalpy* (reactants or products),
   which has *maximum entropy* (reactants or products), and if the reactants are mixed, what will
   happen? (go to completion/ reach a state of equilibrium/not occur at all). Assume there is sufficient
   activation energy to initiate any spontaneous reaction.
   a) \( \text{PCl}_5(\text{g}) \rightleftharpoons \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g}) \); \( \Delta H = +92.5 \text{ kJ} \)

   The __________________________________________ has/have minimum enthalpy.
   The __________________________________________ has/have maximum entropy.

   If \( \text{PCl}_5 \) is put in a flask what should happen? (go to completion/ reach a state of
   equilibrium/not occur at all) __________________________________________________________
b) \[ 2\text{NO}(g) + \text{O}_2(g) \rightleftharpoons 2\text{NO}_2(g) + \text{energy} \]

The ___________________________ has/have minimum enthalpy.

The ___________________________ has/have maximum entropy.

If NO and \( \text{O}_2 \) were put in a flask, what should happen? (go to completion/reach a state of equilibrium/not occur at all)

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c) \[ \text{Na}_2\text{CO}_3(s) + 2\text{HCl(aq)} \rightleftharpoons 2\text{NaCl(aq)} + \text{CO}_2(g) + \text{H}_2\text{O(l)} + 27.7 \text{kJ} \]

The ___________________________ has/have minimum enthalpy.

The ___________________________ has/have maximum entropy.

If \( \text{Na}_2\text{CO}_3(s) \), + 2\text{HCl(aq)} were put in a flask, what should happen? (go to completion/reach a state of equilibrium/not occur at all)

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d) \[ 2\text{Pb(NO}_3)_2(s) + 597 \text{kJ} \rightleftharpoons 2\text{PbO(s)} + 4\text{NO}_2(g) + \text{O}_2(g) \]

The ___________________________ has/have minimum enthalpy.

The ___________________________ has/have maximum entropy.

If \( \text{Pb(NO}_3)_2 \) was put in a flask, what should happen? (go to completion/reach a state of equilibrium/not occur at all)

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16. Reactions which result in a/an ___________________________ in enthalpy and a/an ___________________________ in entropy will **always** be spontaneous.

17. Reactions which result in a/an ___________________________ in enthalpy and a/an ___________________________ in entropy will **always** be non-spontaneous.

**Do Worksheet 2-1**