# Chemistry 12
# Unit 2- Equilibrium
# Notes

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

*Reactants* $\rightarrow$ *Products* *or* *Reactants* $\leftarrow$ *Products*

*Reactants form Products* *Products form Reactants*

For example, under certain conditions, one mole of the colourless gas $N_2O_4$ will *decompose* to form two moles of **brown $NO_2$ gas**:

$$N_2O_4 \rightarrow 2\,NO_2$$

colourless brown

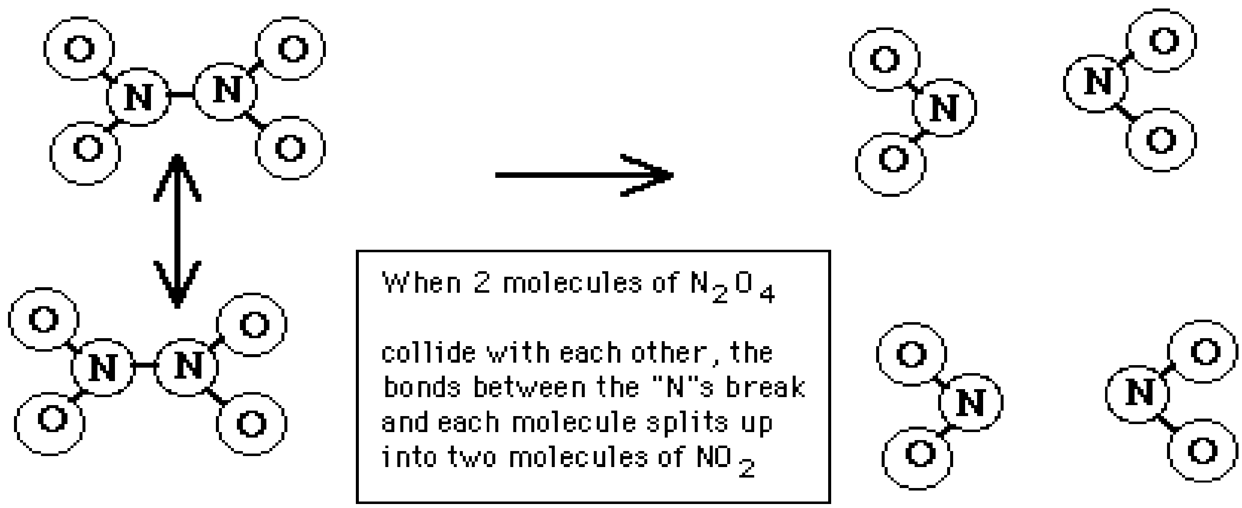
Under other conditions, you can take 2 moles of brown $NO_2$ gas and change it into one mole of $N_2O_4$ gas:

$$N_2O_4 \leftarrow 2\,NO_2$$

colourless brown

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

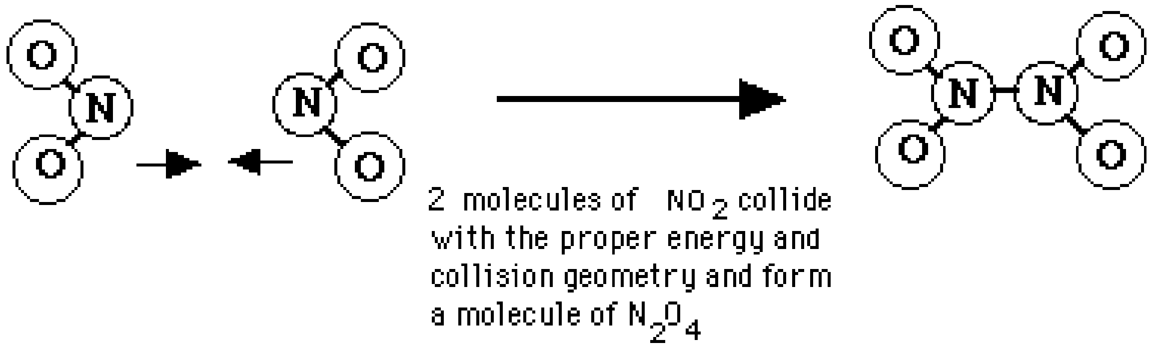
If we were to put some $N_2O_4$ in a flask, the $N_2O_4$ molecules would collide with each other and some of them would break apart to form $NO_2$ .




This process is indicated by the ***forward reaction***:

$$N_2O_4 \rightarrow 2\,NO_2$$

Once this has happened for awhile, there is a build up of $NO_2$ molecules in the same flask
Once in awhile, two $NO_2$ molecules will collide with *each other* and **join** to form a molecule of $N_2O_4$ !




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

$$N_2O_4 \leftarrow 2\ NO_2$$

Two things you'll have to realize is that ***as long as there is*** $N_2O_4$ ***present, the <u>forward</u> reaction will keep on happening*** and ***as long as there is*** $NO_2$ ***present, the <u>reverse</u> 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 $N_2O_4$ might be breaking up, and ***at the same time*** two molecules of $NO_2$ might be joining to form another molecule of $N_2O_4$! 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 **<u>double arrow</u>**:

$$N_2O_4 \rightleftarrows 2\ NO_2$$

**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!

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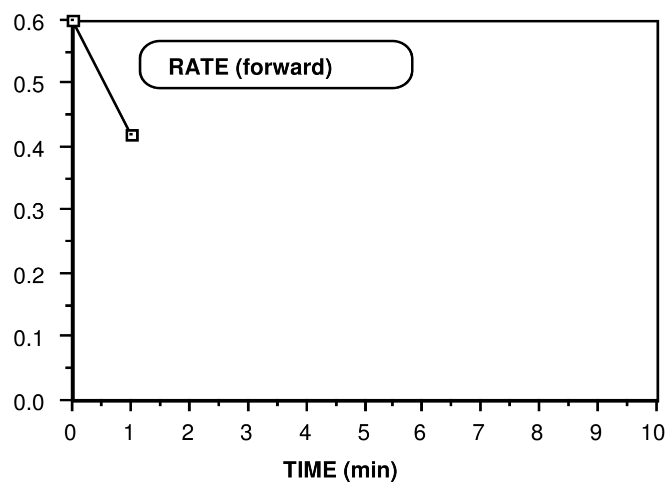
Now what we're going to do is look at how the ***rate*** of the forward reaction changes if we put some pure $N_2O_4$ in a flask:

If we put some pure $N_2O_4$ in a flask (No $NO_2$ yet!), there will be a high *concentration* of $N_2O_4$ . That is, there will be *lots* of $N_2O_4$ 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*.

$$N_2O_4 \rightarrow 2\ NO_2$$

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:



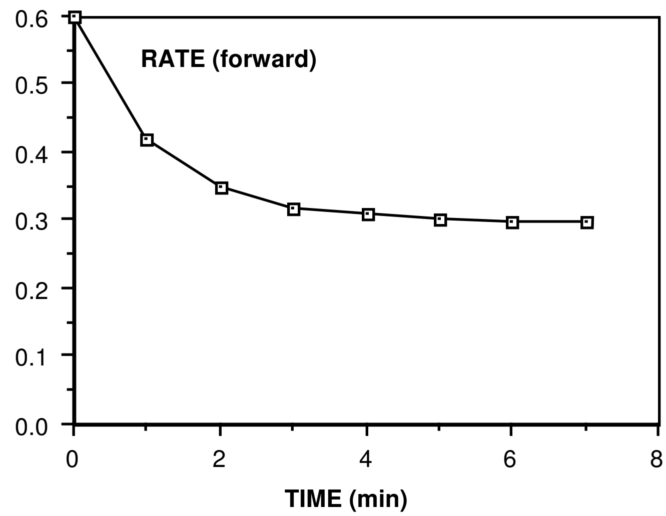

OK, now you might ask: "Why does the rate of the forward reaction go ***down*** ?

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

$$N_2O_4 \rightarrow 2\ NO_2$$

the $N_2O_4$ 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 $N_2O_4$ more slowly, so the [$N_2O_4$ ] 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:

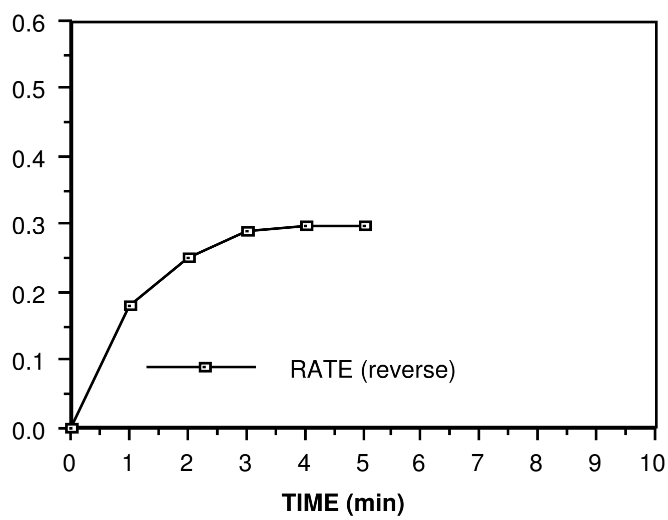



NOW, it's time we consider the rate of the ***reverse reaction***.

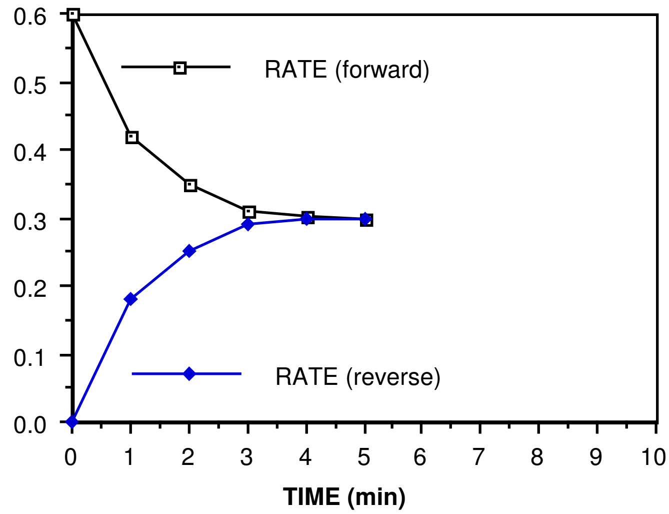
As you might recall, when we have a container full of pure $N_2O_4$ , initially there is ***no*** $NO_2$ in the container. Since there is ***no*** $NO_2$ , there are no $NO_2$ molecules to collide with each other, and the ***rate of the reverse reaction is zero.***
But of course, as time goes on, $NO_2$ is formed from the ***forward reaction*** (***$N_2O_4 \rightarrow 2\ NO_2$***) so in a short time, some $NO_2$ molecules can start colliding and the ***reverse reaction*** will begin:

$$( N_2O_4 \leftarrow 2\ NO_2 )$$

As MORE $NO_2$ is formed by the forward reaction, ***the rate of the reverse reaction gradually <u>increases</u>.*** Now, for you "graph buffs", the graph of the *Rate of the reverse reaction* vs. *Time* might look like this:



OK. Now, lets look at the graph for the forward rate *and* the reverse rate together:




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

***the rate of the forward reaction = the rate of the reverse reaction***

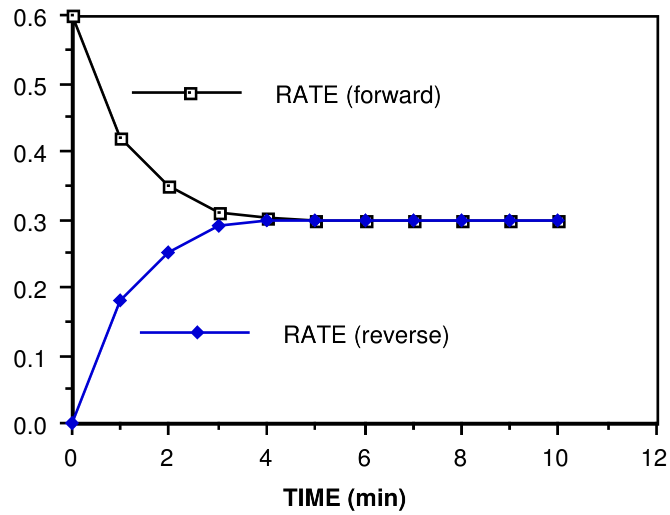
At this point, $NO_2$ is being used up ***at the same rate*** that it is being formed:

$$N_2O_4 \rightleftarrows 2\ NO_2$$

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, $N_2O_4$ is being formed ***at the same rate*** it is being used up. So, also ***the [$N_2O_4$] 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 ***[$N_2O_4$] 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 $N_2O_4$ and the $NO_2$ ) 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:

$$N_2O_4 \rightleftarrows 2\ NO_2$$

for each $N_2O_4$ molecule that breaks up to form two $NO_2$ molecules, two *other* $NO_2$ ***microscopic level***, 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: $N_2O_4 \rightleftarrows 2\ NO_2$
   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:

## Characteristics of a System at Dynamic Equilibrium

1. The ***rate*** of the ***forward*** reaction = The ***rate*** of the ***reverse*** reaction

2. ***Microscopic*** processes (the forward and reverse reaction) continue in a *balance* which yields ***no macroscopic changes***. (so nothing *appears* to be happening.)

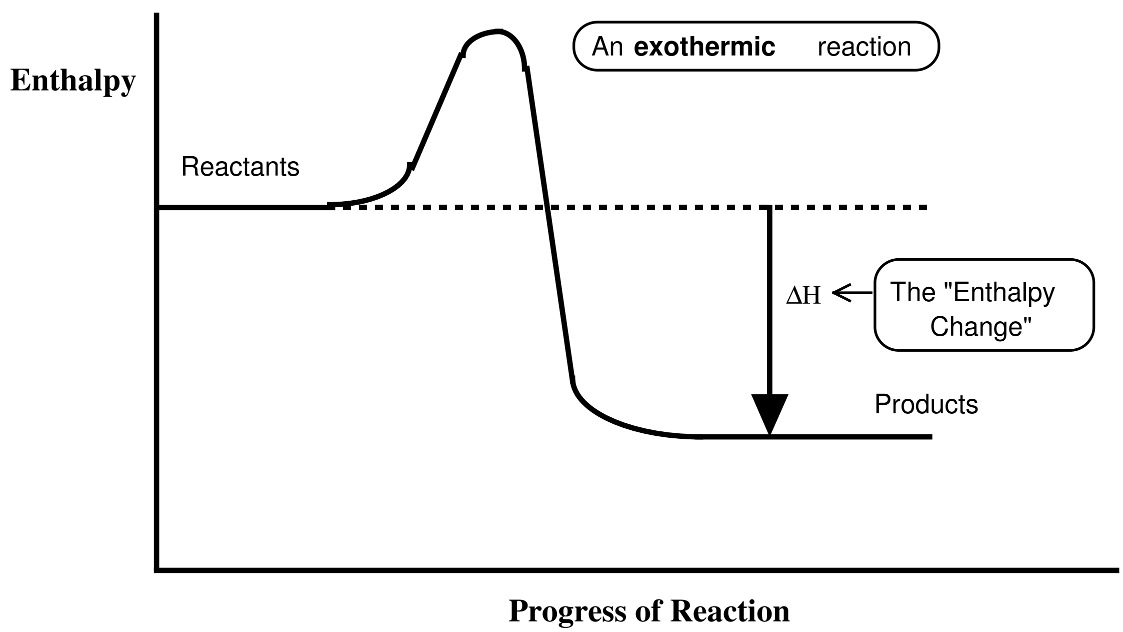
3. The system is ***closed*** and the ***temperature*** is ***constant*** and ***uniform*** throughout.

4. The equilibrium can be approached from the ***left*** (starting with *reactants*) or from the ***right*** (starting with *products*).

# 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 "**ΔH**". In Chemistry 12, a Potential Energy Diagram is the same thing as an "Enthalpy Diagram




The reaction shown on this graph is Exothermic. This means that heat is released or given off. What was given off as heat energy was lost as Enthalpy from the reactants. The net energy change between the products and the reactants is called the Enthalpy Change ( ΔH). As you can see, the enthalpy change ( ΔH) in this reaction is **negative** . (This is always the case when the Products are lower on the Potential Energy (Enthalpy) Graph.)

*In an **Exothermic** Reaction (ΔH is **negative**), the **Enthalpy** is* **decreasing**.

*In an **Endothermic** Reaction (ΔH 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. $A + B \rightleftarrows C + D \quad \Delta H = -24\ kJ$ is *exothermic* so enthalpy is ***decreasing***.
2. $X + Y \rightleftarrows Z \quad \Delta H = 87\ kJ$ is *endothermic* so enthalpy is ***increasing***.
3. $E + D \rightleftarrows F + 45\ kJ$ is *exothermic* so enthalpy is ***decreasing***.
4. $G + J + 36\ kJ \rightleftarrows L + M$ is *endothermic* so enthalpy is ***increasing***.

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

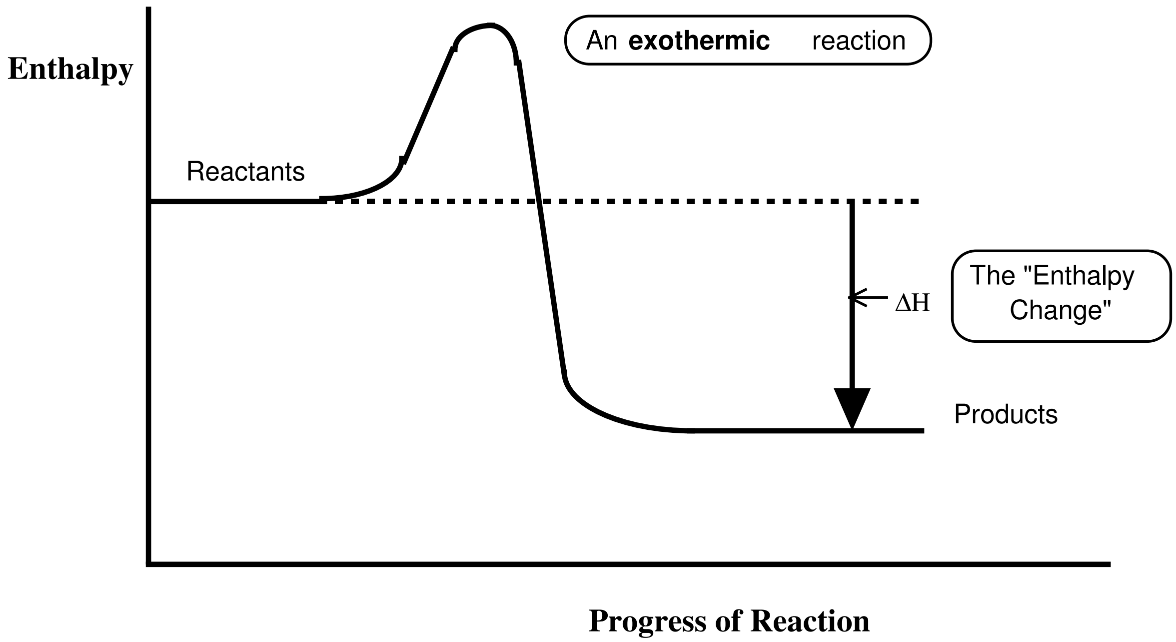
In Chemistry, we ***are*** interested in is *chemical potential energy*, otherwise known as ***enthalpy***!

***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:




Here, the Products have <u>lower enthalpy</u> 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.

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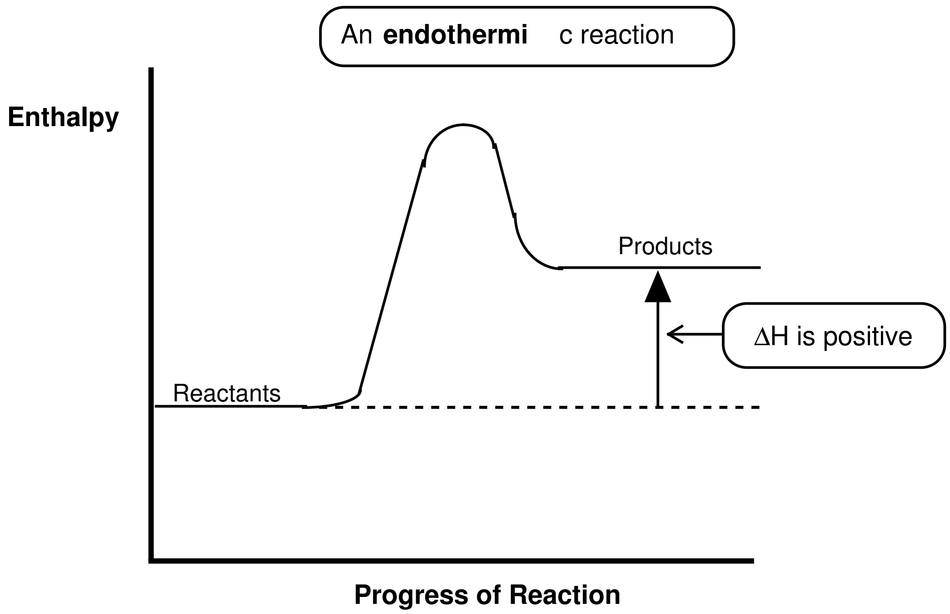
In an ***endothermic reaction***, the ____________________________________ have minimum enthalpy, so the _____________________________________ will be favoured. In other words, if the reactants are mixed they will (tend to remain as reactants / spontaneously form products) ______________

______________________________________________________________________________________

Let's look at a diagram for an *endothermic reaction:*




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.

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

a. $PCl_{5(g)} \rightleftarrows Cl_{2(g)} + PCl_{3(g)}$ $\Delta H = 92.5$ kJ

__________thermic and the ___________________ have *minimum enthalpy.*

b $2NH_{3(g)} + 92.4 \text{ kJ} \rightleftarrows N_{2(g)} + 3H_{2(g)}$

__________thermic and the ___________________ have *minimum enthalpy.*

c $CO_{(g)} + 3H_{2(g)} \rightleftarrows CH_{4(g)} + H_2O_{(g)} + 49.3 \text{ kJ}$

__________thermic and the ___________________ have *minimum enthalpy.*

d. $Cl_{2(g)} \rightleftarrows Cl_{2(aq)}$ $\Delta H = -25$ 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 ____________________________________________ enthalpy.

3 Given the equation: $2NH_{3(g)} + 92.4\ kJ \rightleftarrows N_{2(g)} + 3H_{2(g)}$

If only the *enthalpy* is considered, the (reactant / products) ____________________________ will be favoured at equilibrium.

4. Given the equation: $Cl_{2(g)} \rightleftarrows Cl_{2(aq)}$ $\Delta H = -25\ kJ$

If only the *enthalpy* is considered, the (reactant / products) ____________________________ will be favoured at equilibrium.

5 If the reaction : $CO_{(g)} + 3H_{2(g)} \rightleftarrows CH_{4(g)} + H_2O_{(g)} + 49.3\ kJ$

was proceeding to the *right*, the enthalpy would be _____________________ing. Is this a *favourable* change? _____________.

6. If the reaction: $PCl_{5(g)} \rightleftarrows Cl_{2(g)} + PCl_{3(g)}$ $\Delta H = 92.5\ kJ$

was proceeding to the *right*, the enthalpy would be ____________________ing. Is this a *favourable* change? _____________.

7. If the reaction: $Cl_{2(g)} \rightleftarrows Cl_{2(aq)}$ $\Delta H = -25\ kJ$

was proceeding to the *right*, the enthalpy would be ____________________ing. Is this a *favourable* change? _____________.

8 If the reaction: $2NH_{3(g)} + 92.4\ kJ \rightleftarrows N_{2(g)} + 3H_{2(g)}$

was proceeding to the *right*, the enthalpy would be ____________________ing. Is this a *favourable* change? _____________.

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:

$H_2O_{(s)}$. + heat $\rightleftarrows$ $H_2O_{(l)}$ (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 ( $H_2O_{(s)}$) 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

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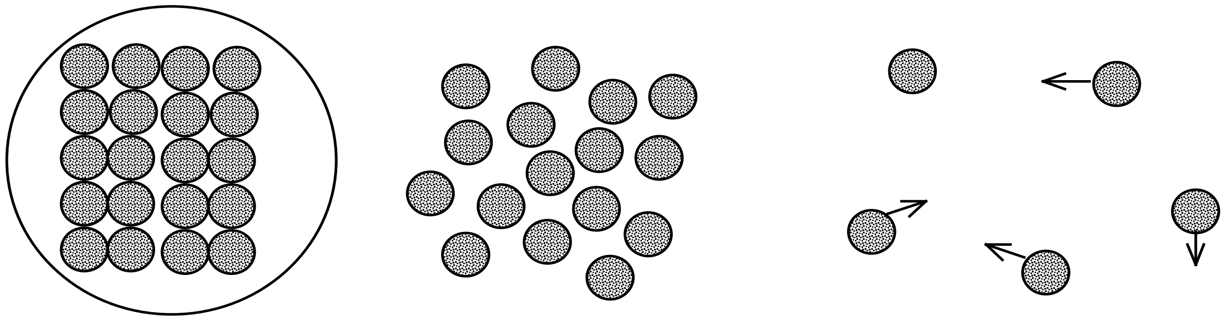
| ***Entropy simply means disorder, or lack of order.*** |

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:

***Entropy of a Solid < Entropy of a Liquid < 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 ***<u>entropy is increasing</u>*** (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:

$$CaCO_{3(s)} + 2\,HCl_{(aq)} \rightleftarrows CaCl_{2(aq)} + \mathbf{CO_{2(g)}} + H_2O_{(l)}$$

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.):

$$4NH_{3(g)} + 5O_{2(g)} \rightleftarrows 4NO_{(g)} + 6H_2O_{(g)}$$

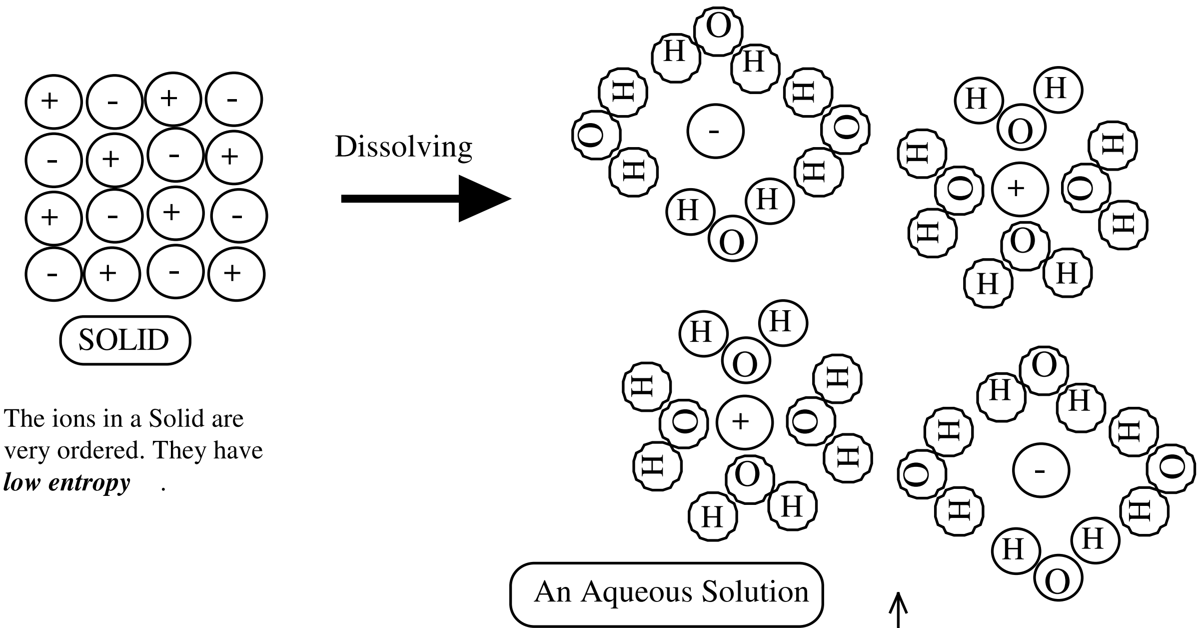
There are (4 + 5) = <u>9</u> moles of gas on the left

There are (4 + 6) = <u>10</u> 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 <u>higher</u> than that of a solid*** .

So in order of lowest to highest **entropy**:

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

<u>Here are few exercises for you:</u>

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

a) $I_{2(s)} \rightleftarrows I_{2(aq)}$ The ___________________________ have greater entropy.

b) $2NH_{3(g)} \rightleftarrows N_{2(g)} + 3H_{2(g)}$

The ________________________________________________have greater entropy.

c) $NH_{3(g)} \rightleftarrows NH_{3(aq)}$

The ________________________________________________have greater entropy.

d) $CO_{(g)} + Cl_{2(g)} \rightleftarrows COCl_{2(g)}$

The ________________________________________________have greater entropy.

e) $MgCO_{3(s)} + 2HCl_{(aq)} \rightleftarrows MgCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

The ________________________________________________have greater entropy.

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

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Remember: $H_2O_{(s)}$ . + ***heat*** $\rightleftarrows$ $H_2O_{(l)}$

We decided that all the $H_2O$ in the universe should remain as a *solid* because $H_2O_{(s)}$ has ***lower enthalpy*** than $H_2O_{(l)}$ and nature ***favours a state of <u>minimum enthalpy</u>***.

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

$H_2O_{(l)}$ has ***<u>higher entropy</u>*** than $H_2O_{(s)}$

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

| |
|---|
| ***Chemical systems will tend toward a state of <u>maximum entropy</u> if no other factors are considered.*** |

Another was of stating this might be:

| |
|---|
| ***A chemical reaction will favour the side (reactants or products) with <u>maximum entropy</u> 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:

$$H_2O_{(s)} + heat \rightleftarrows H_2O_{(l)}$$

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 $H_2O_{(s)}$ to get $H_2O_{(l)}$ , $H_2O_{(s)}$ 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 : $H_2O_{(s)}$ **+ *heat*** $\rightleftarrows$ $H_2O_{(l)}$ , 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***.

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Now, lets 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*** $\rightarrow$ ***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***!

Bottle on a high shelf $\rightarrow$ Thousands of pieces of glass on a concrete floor

Products have minimum "enthalpy", so the tendency toward minimum enthalpy favours the ***products***.

Products have maximum entropy, so the tendency toward maximum entropy 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***.

(ie. 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:

$$2K_{(s)} + 2H_2O_{(l)} \rightleftarrows 2KOH_{(aq)} + H_{2(g)} + heat$$

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

The tendency toward ***minimum enthalpy*** favours the <u>***products***</u>.

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 <u>***products***</u>.

Since <u>***both tendencies favour the products***</u>, this reaction <u>***will go to completion***</u>.

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!

*************************************************

Let's consider one more process:

$$2KOH_{(aq)} + H_{2(g)} + heat \rightleftarrows 2K_{(s)} + 2H_2O_{(l)}$$

In this case, *the tendency toward* <u>***minimum enthalpy***</u> *favours the* <u>***reactants***</u>, and the

*tendency toward* <u>***maximum entropy***</u> *also favours the* <u>***reactants***</u>.

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

(ie. 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) $PCl_{3(g)} + Cl_{2(g)} \rightleftarrows PCl_{5(g)}$ ; $\Delta H = -92.5$ 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) $2NO_{2(g)} \rightleftarrows N_2O_{4(g)}$ + 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) $P_{4(s)} + 6H_{2(g)} + 37$ kJ $\rightleftarrows 4PH_{3(g)}$

The ________________________________ has/have minimum enthalpy.

The ____________________________________ has/have maximum entropy.

If $P_{4(s)}$ and $6H_{2(g)}$ was put in a flask, what should happen?(go to completion/ reach a state of equilibrium/not occur at all)

_____________________________________________________________________

d) $2PbO_{(s)} + 4NO_{2(g)} + O_{2(g)} \rightleftarrows 2Pb(NO_3)_{2(s)}$ ; $\Delta H = - 597$ kJ

The ____________________________________ has/have minimum enthalpy.

The _____________________________________ has/have maximum entropy.

If $PbO_{(s)}$ and $NO_{2(g)}$ 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) $C_{(s)}. + O_{2(g)} \rightleftarrows CO_{2(g)}$ ____________________

    b) $2Al(s) + 6HCl_{(aq)} \rightleftarrows 3H_{2(g)} + 2AlCl_{3(aq)}$ ____________________

    c) $2SO_{3(g)} \rightleftarrows 2SO_{2(g)} . + O_{2(g)}$ ____________________

    d) $HCl_{(g)} \rightleftarrows H^{+}{}_{(aq)} + Cl^{-}{}_{(aq)}$ ____________________

    e) $KOH_{(s)} \rightleftarrows K^{+}{}_{(aq)} + OH^{-}{}_{(aq)}$ ____________________

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) $PCl_{5(g)} \rightleftarrows PCl_{3(g)} + Cl_{2(g)}$ ; $\Delta H = +92.5$ kJ

    The ____________________ has/have minimum enthalpy.

    The ____________________ has/havemaximum entropy.

    If $PCl_5$ is put in a flask what should happen?(go to completion/ reach a state of equilibrium/not occur at all)

    ____________________

b) $2NO_{(g)} + O_{2(g)} \rightleftharpoons 2NO_{2(g)}$ + energy

The ________________________________ has/have minimum enthalpy.

The ________________________________ has/have maximum entropy.

If NO and $O_2$ were put in a flask, what should happen?(go to completion/ reach a state of equilibrium/not occur at all)

______________________________________________________________

c) $Na_2CO_{3(s)} + 2HCl_{(aq)} \rightleftharpoons 2NaCl_{(aq)} + CO_{2(g)} + H_2O_{(l)}$ + 27.7 kJ

The ________________________________ has/have minimum enthalpy.

The ________________________________ has/have maximum entropy.

If $Na_2CO_{3(s)}$ . + $2HCl_{(aq)}$ were put in a flask, what should happen?(go to completion/ reach a state of equilibrium/not occur at all)

______________________________________________________________

d) $2Pb(NO_3)_{2(s)}$ + 597 kJ $\rightleftharpoons 2PbO_{(s)} + 4NO_{2(g)} + O_{2(g)}$ ;

The ________________________________ has/have minimum enthalpy.

The ________________________________ has/have maximum entropy.

If $Pb(NO_3)_2$ was put in a flask, what should happen?(go to completion/ reach a state of equilibrium/not occur at all)

______________________________________________________________

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**