NAME

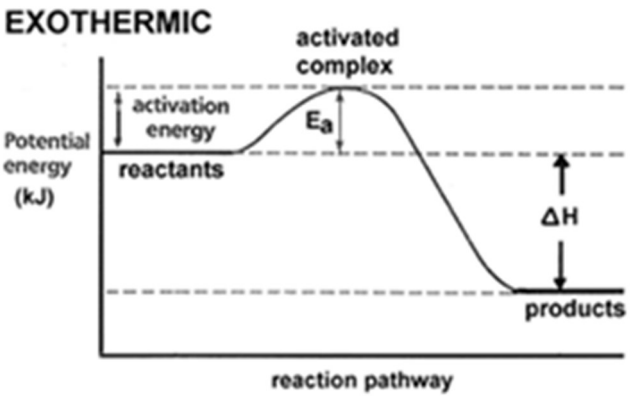
KINETICS PACK

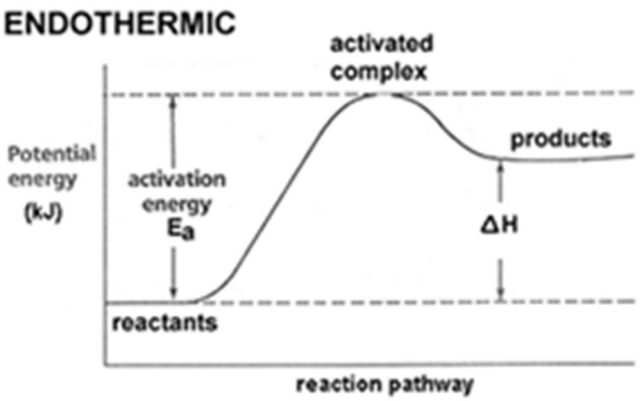
KINETICS BASICS & NOTES

In kinetics we will follow the energy.
We get to see how much energy is absorbed in an endothermic reaction,
or how much is emitted in an exothermic reaction.
We'll learn about catalysts to speed up reactions, or inhibitors to slow them down.

We'll draw graphs for both kinds of reactions (↓).

We will finally learn why a +ΔH is endothermic and a –ΔH is exothermic.





Kinetics & Equilibrium BASICS

Kinetics and Equilibrium is the part of chemistry that is measures the flow of energy through a chemical reaction, the rate of reaction, and finally the concept of dynamic equilibrium as it pertains to chemical reactions. We will examine the energy requirement to start a reaction, and the energy emitted by the reaction.

Most reactions we have done and observed have been "one way" reactions. They are "IRREVERSIBLE" - the "I" in TOPIC-B. Which means that the reactants become products, energy is either absorbed (endothermic) or released (exothermic) and the products that form are permanent. The reason most reactions tend to go only one way is the energy involved is "gone", it's absorbed into bonds, or released into the Universe. This energy is no longer available to feed the reverse reaction.

We learned matter cannot be created or destroyed in any chemical (or physical) process. Neither can energy. Energy must be accounted for. It exists, it can be measured, and it can be transferred from one kind of energy into different kinds of energy. We need to keep track of it.

Endothermic reactions absorb an exact amount of energy as the reaction occurs. Exothermic reactions release a specific amount of energy as they occur. The energy is quantifiable and measurable. We have Table I to tell us how much energy is released, or absorbed by 25 different reactions. This change in energy is called the DELTA "H", or ΔH. The ΔH must always have a sign, positive for endothermic reactions, and negative sign for exothermic ones. No sign indicates no thinking, which is frowned upon.

We will learn to draw potential energy diagrams, which show the change in energy as the REACTANTS form into the PRODUCTS.

COLLISION THEORY of chemical reactions

It is understood that the atoms and molecules need sufficient kinetic energy and proper orientation to each other to react. Too little energy will not lead to a reaction, too much energy would not let the particles to interact, they'd just bang off of each other. Molecular shape means that the particles need to "sort of" fit together to react. Not all collisions lead to a reaction. Proper collisions are the actual cause of chemical reactions to occur.

The more collisions, the faster the reaction should be. The stronger and faster the particles move, the more likely that reactions can happen. Anything that encourages more, or stronger collisions, will lead to faster chemical reactions (except too much energy).

4 factors that affect the RATE of chemical reactions	
Increase in Temperature	Hotter particles move faster, causing more collisions, causing faster reactions. The rate of reaction is directly related to the kinetic energy (temp) of the system (to a point).
Increase in Surface Area	More surface area leads to more collisions, and faster reactions. Smaller particles means "less waiting" for particle collisions. Larger chunks of matter mean the inner particles can react only after the outer ones react first. Smaller particles = less wait time to react.
Increase the Concentration	Stronger solutions have more ions or molecules per mL, leading to more collisions and of course, faster reactions. A lower solution molarity = less particles and less collisions and slower rates of reaction.
Catalysts	Catalysts do speed up chemical reactions, but they do not increase the collision rate. Catalysts are said to lower the activation energy needed to start the reaction, and/or provide an alternate chemical pathway (a chemical shortcut) which is quicker.

Table I — Heats of Reaction at 101.3 kPa and 298 Kelvin

Table I shows 6 combustion reactions, 12 synthesis reactions, and 6 ionic compounds forming a mixture (solution) in water. The last line of the table we won't discuss for a few more weeks.

All of these have what is called the DELTA "H", or ΔH. That is the change in heat of the reaction. For the ionic compounds dissolving into water, this is the ΔS or heat of solution, which is the ΔH of solution formation which is NOT a chemical reaction. All combustion reactions are exothermic, energy is a product. All have negative ΔH values. At the very bottom of the table is says "A minus sign indicates an exothermic reaction."

Energy is not negative, or positive. The negative sign is the reminder that the energy is a product, it's emitted.

The synthesis reactions have either negative or positive ΔH values. Those with a + ΔH are endothermic, they absorb energy when they occur. The reactions with a –ΔH are exothermic (like the combustion reactions).

These ΔH are called the HEATS OF REACTION (that's the title of the table).

As a reminder, the first reaction, the combustion of methane, is exothermic. How exothermic?
The ΔH is –890.4 kJ/mole, which means when one mole of methane combusts, the reaction emits 890.4 kJ (kilojoules) of energy as a product.

Two moles of methane combusting releases 2 X 890.4 kJ

3.50 moles of methane combusting releases 3.50 X 890.4 kJ

The mole ratio for this reaction is 1:2:1:2, and the thermochemical mole ratio is 1:2:1:2: 890.4 kJ

All the reactions can be reversed. Find this: CO and O_2 and form CO_2 which releases 566.0 kJ of energy.

Table I shows this: $2CO_{(G)} + O_{2(G)} \rightarrow 2CO_{2(G)}$ $\Delta H = -566.0$ kJ/mole

That means: $2CO_{(G)} + O_{2(G)} \rightarrow 2CO_{2(G)} + 566.0$ kJ (energy is a product, exothermic)

You can reverse this to: $2CO_{2(G)} + 566.0$ kJ $\rightarrow 2CO_{(G)} + O_{2(G)}$ (energy is absorbed, endothermic)
These reactions can all be "reversed" and the amount of energy is identical. The sign for ΔH changes. Endothermic reactions reverse to exothermic. Exothermic reactions reverse to endothermic.

In Table I $NaCl_{(S)} \xrightarrow{water} Na^{+1}_{(AQ)} + Cl^{-1}_{(AQ)}$ $\Delta H = +3.88$ kJ/mole

which means $NaCl_{(S)} + 3.88$ kJ/mole $\xrightarrow{water} Na^{+1}_{(AQ)} + Cl^{-1}_{(AQ)}$

Reversed $Na^{+1}_{(AQ)} + Cl^{-1}_{(AQ)} \longrightarrow NaCl_{(S)} + 3.88$ kJ/mole Water evaporates

The reversed reaction has the same chemistry and the same energy, but the energy changes from endo to exothermic.

POTENTIAL ENERGY DIAGRAMS

We can show the flow of energy in a chemical reaction in a graph called a Potential Energy diagram.

They show the energy (in the bonds of the reactants and products) during the chemical reaction. The 2 kinds of PE Diagrams, which are exothermic potential energy diagrams, and endothermic potential energy diagrams. They're similar (but different). You will need to understand them to tell them apart. You must recognize these differences so you can determine what the energy flow in a reaction is by looking at the diagram.

Exothermic Potential Energy Diagrams (below)

The Y axis shows potential energy in kilojoules per mole. The X axis label is always about the start and end of the reaction. Here it's "reaction progress", but you might see "time of reaction" or "reaction pathway"

Find REACTANTS. This shows the potential energy that is in the reactants. In high school chem we do not know exactly how much potential energy these reactants have. It is a constant, and its value is available in college texts, but we will NOT have to know these exact numbers. It stays constant, or flat, forever, the potential can't change. The reactants have a certain amount of potential, forever.

When you add energy, you heat up the reactants, they gain energy and become more likely to react. For example, if your reactants are explosive when they react, hotter reactants have a greater potential to explode. They can only react if they get ENOUGH energy to react. That amount energy required to get to the "top of the hill" on the graph is called the "ACTIVATION ENERGY". That is how much energy is required to push the reaction into happening. Less than that amount of energy will encourage the reaction, but it can't actually happen with insufficient activation energy.

Once the reaction gets to "the top of the hill" it cannot be stopped and it proceeds to form products.

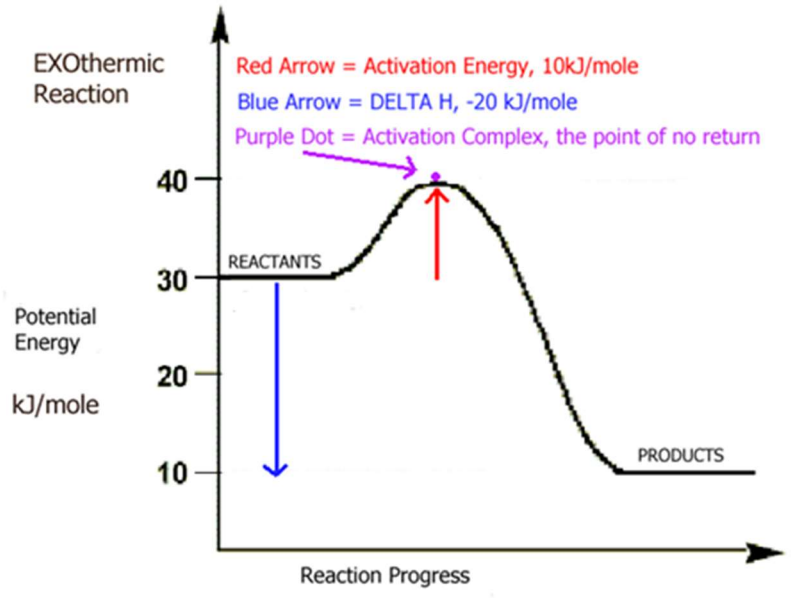
In an exothermic reaction the products have less energy than the reactants started with. Where did that energy go? It was released, exothermically. It's not lost; it is released from the reaction into the environment.

When the reactants gain the activation energy, they temporarily form into the "ACTIVATED COMPLEX".

This is the in-between stage when reactants come apart but are not yet formed into products. A sort of in between state that lasts for nanoseconds. The activation complex forms at the top of the curve of the potential energy diagram.

The difference in potential energy between the reactants and products is the ΔH.

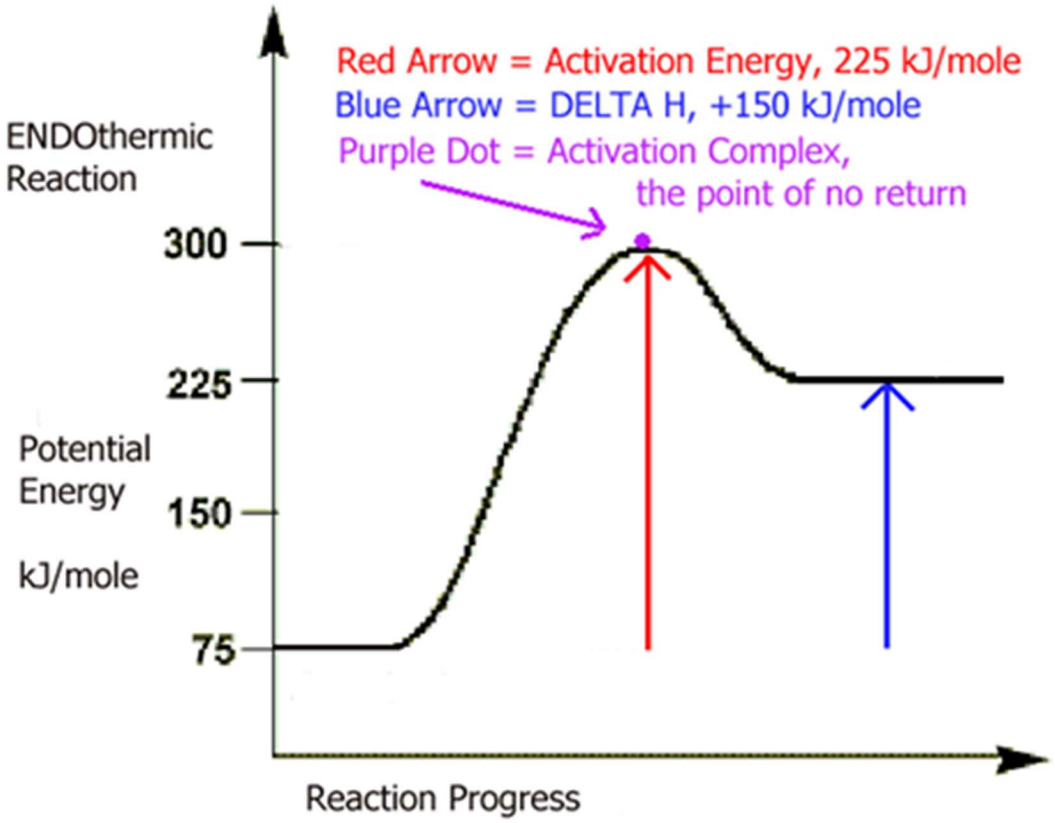
The ΔH is negative in an exothermic reaction, because the products have less potential energy than the reactants started with This "missing" energy is released into the environment.



The ENDOTHERMIC potential energy diagram

The ENDOTHERMIC potential energy diagram has all the same parts as the endothermic one, the big difference is that the products have MORE potential energy than the reactants did. For this to happen, this extra energy has to be absorbed from the immediate environment and put into the products.

That is why endothermic reactions feel "colder", they are literally absorbing energy into the bonds from the surroundings, taking this kinetic energy and converting it into potential energy stored in the products.



The reactants start with 75 kJ/mole potential energy at the "beginning of the graph".

To react they need to absorb the activation energy of 225 kJ/mole *(that's 300 — 75 = 225)*

The Activated complex has 300 kJ of energy. Once the products start to form energy is released.

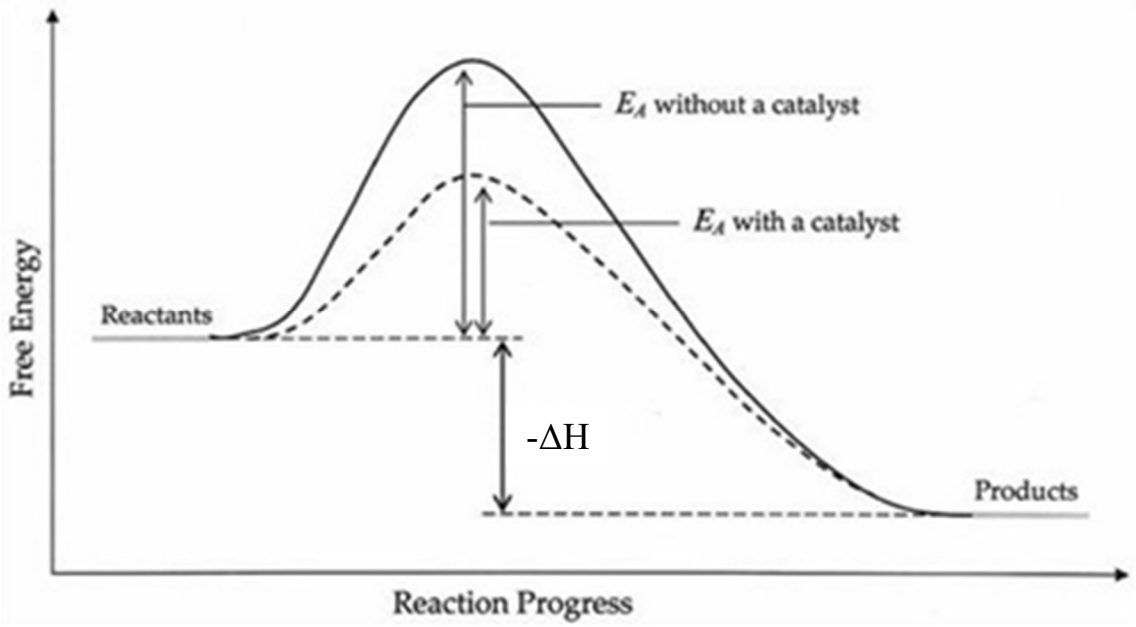
The products end up with 225 kJ/mole of potential energy.

That's a difference of (net gain) of 150 kJ/mole. The ΔH is +150 kJ/mole *(that's 225 — 75 = 150)*

Make sure you touch the diagram and do some simple math to make sure you "get" the numbers and grasp the "difference between the reactants and the products" is the +ΔH.

Catalysts in Potential Energy Diagrams

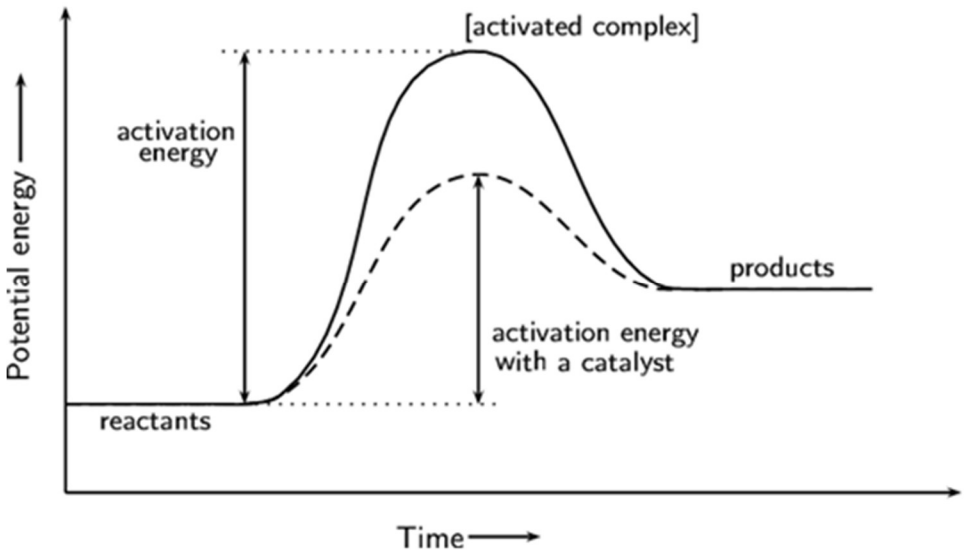
A catalyst lowers the activation energy of the reaction. It takes less energy to get to the "top" of the hill, so the reaction happens faster. Or the catalyst provides an "alternate pathway" for the reaction to occur, which is very hard to describe at our level of chemistry, but trust me, it means faster. You will show the effect of a catalyst with a dotted line. The catalyst must lower the AE, but the reaction "pathway" needs still to go both "up and down" inside the middle of the graph.



Here E_A means the energy of activation which is activation energy.

Normally this reaction follows the solid line to the "top" of the hill. With a catalyst the reaction can follow an alternate path (a shortcut so to speak) which is the dotted line.

The ΔH does not change, the catalyst does not change the start or end points. ΔH is this difference.
Since this graph goes "DOWN", the ΔH is negative.



Here the catalyst affect (dotted line) still goes "up and down" inside the dome of the graph, which lowers the AE but not changing the ΔH. Since the graph goes "UP", the ΔH is POSITIVE.
ΔH is not indicated on this graph.

An inhibitor will of course inhibit or slow or even stop a chemical reaction.
Often, they block the catalyst from working; they can also just stop the reaction.

Dynamic Equilibrium

Most chemical reactions you have seen so far are "irreversible" meaning once they occur, they stay done. The products form, the products do not spontaneously revert back into the reactants. They can't, as the energy required is insurmountable without help from a scientist.

Some reactions are more easily reversed than others, the energy required to go forward, or even in reverse is low enough, and similar enough, that air temperature is sufficient to push reactions in either direction. Some reactions can alternate back and forth over and over again, in a sort of balanced way. When the rate of the forward reaction is equal to the rate of the reverse reaction, this chemical system is said to be in dynamic equilibrium. That means, always changing and always equal at the same time.

When a chemical system is in dynamic equilibrium (sometimes just called equilibrium), it can appear that NOTHING is happening. This is of course totally untrue. In equilibrium the forward reaction keeps happening at same rate as the reverse reaction. Constantly changing and constantly staying the same.

The symbol is this: ⇋
In our class, ANY reaction that has this double arrow will be considered to be in equilibrium, even if you might not think that it's possible. In high school we sometimes use reactions that are not readily reversible, but we use this double arrow anyway, to help you think.

The forward reaction here is the synthesis of AB. The forward reaction is endothermic; energy is a reactant.
The reverse reaction is the decomposition of AB and it's exothermic, energy is a product in the reverse reaction.
The double arrows mean that the synthesis happens at a certain rate, decomposition occurs at the same rate.

$$A + B + \text{energy} \leftrightharpoons AB$$

The HEAT OF REACTION for the forward reaction is unknown, but it certainly is a $+\Delta H$
The reverse reaction must be a $-\Delta H$, of unknown value, but equal to its opposite.

When a reaction is at dynamic equilibrium, this balance can be disrupted by three ways:
- changing the temperature (adding or removing energy)
- changing the pressure (this only affects the gases in the reaction)
- adding or removing stuff (reactants or products)

Adding a catalyst will not shift an equilibrium, catalysts just bring about equilibrium faster.

Equilibrium is a state that can exist depending upon the current conditions. When these conditions are altered (change temp, pressure, or adding/removing stuff), this equilibrium accommodates the stress, and creates a new dynamic equilibrium at the new conditions. Every time you adjust the conditions, the dynamic equilibrium will adjust o relieve the stress you have applied, and another new equilibrium forms.

This seems basic when you think about it, but it was proposed first by the French chemist named Henry Louis LeChatelier in the early 1880's. Memorize "LeChatelier's Principle".

"A dynamic equilibrium will stay in equilibrium. If a stress is added to a chemical equilibrium, the system will adjust to relieve the stress, and a new dynamic equilibrium will form."

When you have a chemical reaction in dynamic equilibrium, you can adjust temperature, pressure, or amounts of reactants, and you can predict which way the equilibrium will shift: FORWARD or REVERSE when it accommodates that stress. A new equilibrium will form.

$N_{2(G)} + 3H_{2(G)} \leftrightharpoons 2NH_{3(G)} + \text{energy}$

Add N_2 →	Forward shift, adding N_2 causes more collisions to speed up synthesis, until a new dynamic equilibrium is reached.
Add H_2 →	Forward shift, adding H_2 causes more collisions to speed up synthesis, until a new dynamic equilibrium is reached.
← Add NH_3	Reverse shift, adding NH_3 (ammonia), the only thing NH_3 does here is decompose. More NH_3 means faster decomposition, until a new dynamic equilibrium is reached.
← Add heat	Reverse shift, more heat will cause faster decomp, until a new dynamic equilibrium is reached.
Increase Pressure →	This is special. Higher pressure favors less moles of gas. There are less moles as NH_3 than as N_2 and H_2. A forward shift occurs until a new dynamic equilibrium is reached.
← Decrease pressure	This is special. Lower pressure favors more moles of gas. There are more moles of gas as N_2 and H_2, so it shifts to more to that side, reverse, until a new dynamic equilibrium is reached.
X Add a Catalyst X	Catalysts do not favor forward or reverse, They bring about an equilibrium faster. Indicate this with a "double X" , not arrows.
← Remove H_2	Removing H_2 makes the forward reaction slow down. With less H_2 we get less collisions. Reverse shift, until a new dynamic equilibrium is reached.
Remove NH_3 →	Less decomp occurs so we get a forward shift, until a new dynamic equilibrium is reached.
Remove heat →	Less heat slows the reverse reaction while the forward reaction keeps going. Here we get a forward shift, until a new dynamic equilibrium is reached.

This reaction is technically NOT a real dynamic equilibrium. The forward reaction is very exothermic ($\Delta H = -3351$ kJ), but if the double arrow are present, we will use this as a thinking exercise.

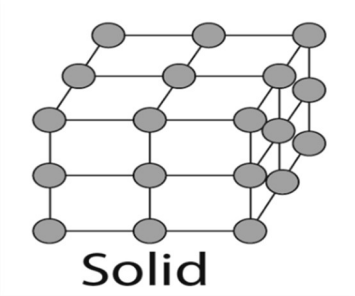
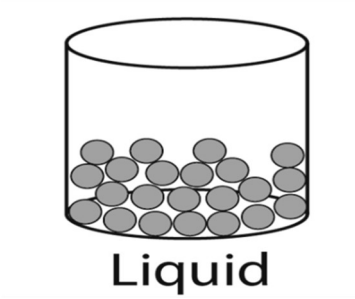
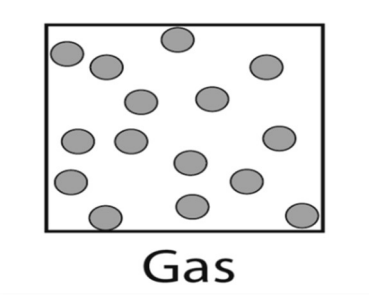
$$4Al_{(S)} + 3O_{2(G)} \leftrightharpoons 2Al_2O_{3(S)} + \text{energy}$$

← Add Al_2O_3	Reverse shift, more Al_2O_3causes more decomposition, until a new dynamic equilibrium is reached.
Add O_2 →	Forward shift, causes more collisions to speed up synthesis, until a new dynamic equilibrium is reached.
Add Al →	Forward shift, causes more collisions to speed up synthesis, until a new dynamic equilibrium is reached.
← Decrease pressure	Lower pressure favors more moles of gas. There are 3 moles of gas on the left, and NO moles of gas on the right. Reverse shift until a new dynamic equilibrium is reached.
← Add heat	Reverse shift, more heat will cause faster decomp, until a new dynamic equilibrium is reached.
Increase Pressure →	Higher pressure favors less moles of gas. There are NO moles of gas on the right, so this will favor a forward shift until a new dynamic equilibrium is reached.
X Add a Catalyst X	Catalysts do not favor forward or reverse, but they do get to equilibrium faster. We indicate this will a "double X" instead of two arrows.
← Remove Al	Removing Al makes the forward reaction slow down. With less Al we get less collisions. Reverse shift, until a new dynamic equilibrium is reached.
Remove Al_2O_3 →	By removing Al_2O_3 there will be forward reaction but much less of the reverse, so we will get a forward shift, until a new dynamic equilibrium is reached.
Remove heat →	Less heat slows the reverse reaction while the forward reaction keeps going. Here we get a forward shift, until a new dynamic equilibrium is reached.

ENTROPY

Entropy is the measure of disorder, or chaos, in a chemical system. Technically it is quantifiable, with complicated math, but that is outside the scope of our class. We will just compare entropy.

We can compare entropy by PHASE water in its three phases.

		
Lowest Entropy Ice has the lowest entropy because the molecules are stuck in place. They can "hum" but they don't really move around. They are well organized, and not chaotic.	**Medium Entropy** Water has medium entropy. Liquids stay within the beaker, but the molecules are in constant motion.	**Highest Entropy** Steam has the highest entropy. Gas molecules move the fastest and are not likely to stay contained. They have high chaos compared to the other phases.

Another way to compare entropy is by temperature.

1.5 M $NaCl_{(AQ)}$ at 10°C	1.5 M $NaCl_{(AQ)}$ at 40°C	1.5 M $NaCl_{(AQ)}$ at 85°C
COLDEST has the LOWEST ENTROPY	Medium temperature has MEDIUM ENTROPY	HOTTEST has the HIGHEST ENTROPY

The Universe is tending towards higher entropy and lower energy. Matter is breaking down into smaller particles; the Universe is falling apart. Energy is dissipating.

The end. ☺

Kinetics Class #1
Objective: Introduction to kinetics and equilibrium chemistry. Four factors that affect the rate of chemical reactions.

1	hydrogen gas + oxygen gas →
2	hydrogen peroxide →
3	That second one was SO SLOW that it took a <u>catalyst</u> so we could watch it! The catalyst was potassium iodide. Write it in the equation where it belongs… hydrogen peroxide ⟶ water + oxygen gas + Energy!
4	The CATALYST is not a
5	Kinetics is the part of chemistry that studies the
6	Reactions that absorb energy are called:
7	Their opposites, reactions that emit energy as a product are called
8	One of the simplest reactions we know is the combustion of methane. $CH_{4(G)} + 2O_{2(G)} \rightarrow 2H_2O_{(G)} + CO_{2(G)}$ + energy This reaction is _____, because heat is a _____ The _____ the _____ table I = _____
9	The forward reaction is _____<u>THERMIC</u> because energy is a
10	At the bottom of Table I it says

11	$CH_{4(G)} + 2O_{2(G)} \rightarrow 2H_2O_{(G)} + CO_{2(G)}$ + energy The mole ratio of this equation would be: _____ The thermochemical mole ratio would be: _____
12	Why is this chemical reaction irreversible?
14	What is different here compared to the combustion of methane? $2H_2O_{(L)}$ + energy → $O_{2(G)}$ + $2H_{2(G)}$ Here, energy is a _____, this reaction is _____
15	Table I shows this… $2H_{2(G)} + O_{2(G)} \rightarrow 2H_2O_{(L)}$ + 571.6 kJ EXOTHERMIC! $2H_2O_{(L)}$ + 571.6 kJ → $O_{2(G)}$ + $2H_{2(G)}$ This is _____
16	This decomposition reaction is _____, of table I where water is the product. You are allowed to do this reverse a reaction, but if you do, you must reverse the sign for ΔH

17. Reaction	ΔH	Exo or endo
$2C_8H_{18(L)} + 25O_{2(G)} \rightarrow 16CO_{2(G)} + 18H_2O_{(G)}$	-10943 kJ/mole	exo
$N_{2(G)} + O_{2(G)} \rightarrow 2NO_{(G)}$		
$C_2H_{2(G)} \rightarrow 2C_{(S)} + H_{2(G)}$		
$4Al_{(S)} + 3O_{2(G)} \rightarrow 2Al_2O_{3(S)}$		

18. Reaction	ΔH	Exo or endo
$C_3H_{8(G)} + 5O_{2(G)} \rightarrow 3CO_{2(G)} + 4H_2O_{(G)}$		
$CO_{2(G)} \rightarrow C_{(S)} + O_{2(G)}$		
$NaOH_{(S)} \xrightarrow{\text{water}} Na^{+1}_{(AQ)} + OH^{-1}_{(AQ)}$		
$2NH_{3(G)} \rightarrow N_{2(G)} + 3H_{2(G)}$		

To drive from Vestal to Johnson City High School it's about 8 miles. That should take about 20 minutes.	$\frac{8 \text{ miles}}{0.33 \text{ hours}} = \sim 24 \text{ mph}$
It you drive fast, getting there in just 10 minutes…	
It you drive fast, getting there in just 5 minutes…	

What is rate of reaction?	
28	TIME is _____ IT TAKES to complete a chemical reaction, usually measured in seconds.
29	Rate is different. RATE =
30	Rate ( ) Time *Not in driving, not in chem.*
31	Rate has a weird unit:

These factors affect the **rate** of a reaction	
32	Increase in
33	Increase in
34	Increase in
35	Adding
36	Increase in Temperature causes more
37	Increase in reactant surface area causes more
38	Increase in concentration of reactants causes more
39	Adding a catalyst has no effect on the number of (but they still speed up chemical reactions. More to follow)

Objective: Potential Energy Diagrams are the graphs that show the flow of energy through chemical reactions. They come in 2 flavors, exothermic and endothermic.	
40	The Law of Conservation of Energy
41	Potential energy diagrams

42. Draw the potential energy diagram for the combustion of CH_4

The potential energy diagram for the combustion of methane

43. Potential Energy Diagram vocabulary

	the kJ/mole of energy that the reactants <u>start</u> with
	the kJ/mole of energy that the products <u>end up</u> with
	the difference in PE between reactants and products this PE diagram has a $-\Delta H$
	the energy required start the reaction (to get to the top of the hill)

44. Draw the potential energy diagram for the synthesis of $C_2H_{2(G)}$ (look for that on table I now)

45. Draw the potential energy diagram for the solvation of sodium hydroxide

46. This energy is the

47. More Kinetics Vocabulary	
	The energy stored in the bonds of the reactants or products
	A transitional state; the reactants have come apart – but have not yet turned into the products
	The minimum amount of energy it takes to start a reaction.
	The difference between the potential energy of the reactants and products. This can be the Heat of Reaction or Heat of Solution
	A graph showing the flow of energy in a chemical reaction. These can be exothermic or endothermic.

48. Draw the potential energy diagram for the synthesis of aluminum oxide.
Label the reactants, products, AE, AC, ΔH. Graph needs axis labels, and a title.

49. Draw the PE diagram for the solvation of sodium chloride.

$$NaCl_{(S)} + \text{energy} \xrightarrow{\text{water}} Na^{+1}_{(AQ)} + Cl^{-1}_{(AQ)}$$

50. Draw the PE diagram for the combustion of propane.
Indicate the PE of reactants, products, AE, AC, ΔH, and a title.

51. The effect of catalysts on chemical reactions	
Draw the PE diagram for the combustion of glucose	Draw the PE diagram for the synthesis of HI gas.

52	What is a Catalyst?
53	Catalysts…

55	Catalysts <u>DO NOT</u> change the number of collisions. Catalysts are said to work
56	Catalysts _____ _____ _____. *This lets the reaction start in a less energetic way, so it can happen quicker than normal. Once it starts, it moves right along.*
57	Catalysts provide "_____" for the reaction to proceed. *This lets this reaction happen quicker. (Like a shortcut).*

58. DEMO "Catalysts speed up the rate of a chemical reaction."

_____

write chemical equation on this line

Objective: Describing chemical reactions that are in dynamic equilibrium; learning how to "push" them forward, or in reverse, using LeChatelier's Principle.	
On August 7, 1974 Phillipe Petit decides it's his day.	
59	Most chemical reactions are "_____", meaning that once they happen, they are done. Spontaneous reversals in chemistry are not common, because most reactions are without enough energy to go the other way.
60	Some reactions are _____ because their energy requirements, their ΔH is very low in _____ directions, forward and reverse.

61	One of the most important reversible chemical reactions is…
62	is when the rate of the forward reaction is equal to the rate of the reverse reaction.
63	In this reaction… $N_{2(G)} + 3H_{2(G)} \leftrightarrow 2NH_{3(G)} +$ energy The rate of the forward reaction (synthesis) is the same as the _____ reaction (decomposition)
64	A dynamic equilibrium is
65	
66	
67	

68	LeChatelier's Principle (memorize)
69	You can disrupt a chemical equilibrium by
70	The chemical system must accommodate this stress, and
71	What happens if we pump in extra $NH_{3(G)}$ to this closed system? nitrogen + hydrogen ↔ ammonia + energy add Ammonia
72	This closed system is in dynamic equilibrium. Let's apply different stresses and see which way the system will "push" to create a new dynamic equilibrium. $N_{2(G)} + 3H_{2(G)} \leftrightarrow 2NH_{3(G)} + \text{energy}$ Add nitrogen Add hydrogen Add ammonia Add energy (heat) Add pressure
73	Remove nitrogen Remove hydrogen Remove ammonia Remove energy (cool system) Lower pressure

Objective
Students will see a dynamic equilibrium reaction undergo a forward shift and a reverse shift. Then we'll practice with dynamic equilibrium equations to drill the LeChatelier's Principle.

This compound is called Bromthymol Blue. It's big, it's ionic, and in water it will ionize into an acid H^{+1} cation, and a big $C_{27}H_{29}O_5S^{-1}$ anion.

Copy the dynamic equilibrium and write blue and yellow as well

74

75 This will always change but always be the same.

76

$$HC_{27}H_{29}O_5S_{(S)} \leftrightarrow H^{+1}_{(AQ)} + C_{27}H_{29}O_5S^{-1}_{(AQ)}$$

molecule YELLOW — Anion BLUE

Add H^{+1}

Add OH^{-1}

$$4Al_{(S)} + 3O_{2(G)} \leftrightarrow 2Al_2O_{3(S)} + Energy$$

This is NOT really a dynamic equilibrium. We will believe in ghosts, or dinosaurs now. If the reaction has the double arrows, we will use the concepts of dynamic equilibrium and LeChatelier's Principle to "play" with the theory. This will NOT reverse with just heat in real life, but on paper, it can.

	$4Al_{(S)} + 3O_{2(G)} \leftrightarrow 2Al_2O_{3(S)} + \text{Energy}$
77	The forward reaction is _____; the reverse reaction is _____. The forward reaction is _____; the reverse is _____.
78	$4Al_{(S)} + 3O_{2(G)} \leftrightarrow 2Al_2O_{3(S)} + \text{Energy}$ Add aluminum oxide Remove oxygen Remove heat (cool system) Add aluminum Add Heat Increase pressure
79	Pressure only effects gases. Pressure does not effect...
80	If a stress "stops" or slows down a forward reaction,
81	If a stress "stops" or slows down a reverse reaction,
82	A new

83. This is NOT a real dynamic equilibrium, but we will "make believe" it is, and accept this so we can think about double arrows and dynamic equilibrium.

$$CH_{4(G)} + 2O_{2(G)} \leftrightarrow CO_{2(G)} + 2H_2O_{(G)} + \text{energy}$$

Add methane

Add water

Add heat

Remove carbon dioxide

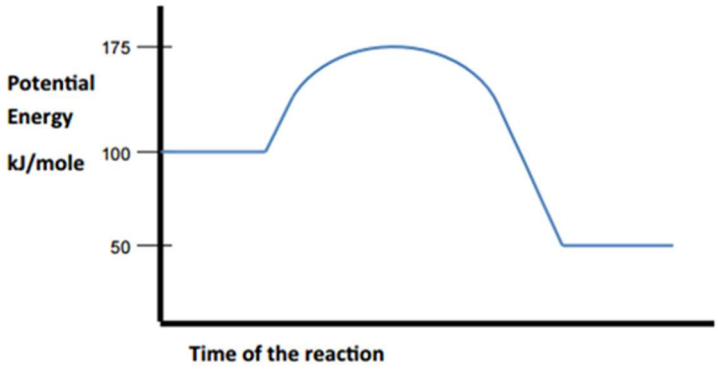
Remove heat

Remove methane

Add carbon dioxide

INCREASE PRESSUE

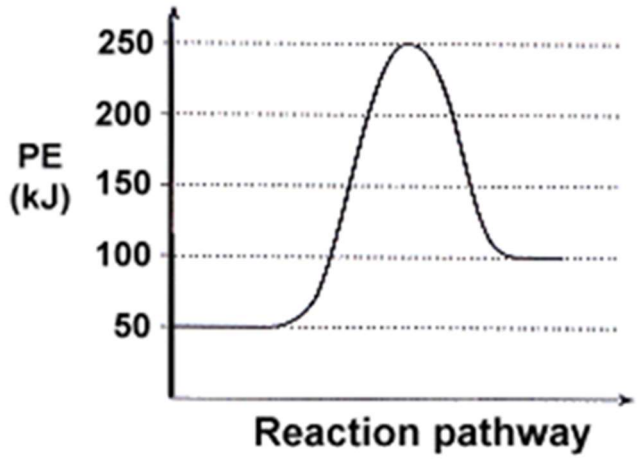
DECREASE PRESSURE



Using this diagram,
answer questions #84 to #88.

Use units where appropriate.

84	Is this an exothermic or endothermic reaction?	
85	What is the potential energy of the activated complex?	
86	What is the PE of the products?	
87	What is the ΔH?	
88	What would be a possible activation energy with a catalyst?	



Using this diagram,
answer questions #89 to #94.

Use units where appropriate.

89	What is the PE of the reactants?	
90	What is the activation energy for this reaction?	
91	Is this reaction exothermic or endothermic?	
92	What are possible AE values for this reaction with a catalyst?	
93	What is the ΔH for this reaction?	
94	Would the ΔH for this reaction change with a catalyst?	

95. In this dynamic equilibrium…
Which way does each stress push the reaction with all of these stresses?

$$3A_{(G)} + B_{(S)} \leftrightarrow 3C_{(G)} + D_{(G)} + \text{energy}$$

Add heat

Add B

Inc. pressure

Remove D

Add C

96	Write LeChatlier's Principle
97	State 4 ways to stress a dynamic equilibrium
98	How do the 4 factors that would that <u>speed up</u> a chemical reaction, three of them work one way, but the 4th factor works a different way.
99	Entropy is the

100. Compare the phases of matter for entropy	
Solid	
Liquid	
Gas	
102	
103	

There are more slides for you to review. These notes are done, you are not done.

Energy of Reactions
ΔH+ is endothermic
ΔH- is exothermic
AE is activation energy
PE is energy stored in bonds
When bonds form, energy is released.

KINETICS
and
EQUILIBRIUM

Heats of Reaction are in Table I. If you reverse an arrow, then reverse the sign for ΔH too.

Collision Theory
Anything that increases the rate and strength of collisions increases the rate of reaction
INCREASE in Temperature, Surface Area, or the Concentration of the Solutions, increases reaction rate

LeChatleier's Principle
A chemical system at equilibrium will stay at equilibrium. If a stress is applied to a dynamic equilibrium, the system will shift (forward or reverse) to relieve that stress, and create a new dynamic equilibrium.
Chemical stresses include increasing or decreasing heat, pressure, the concentrations of the reactants or products.

Pressure ONLY affects gases,
Catalysts move systems to new equilibrium quicker, but don't favor forward or reverse.
EXAMPLES

$N_{2(G)} + 3H_{2(G)} \leftrightarrow 2NH_{3(G)} + Energy$
← Add Heat
Remove Heat →
Add Pressure →
← Remove Pressure
Add N_2 →
← Remove N_2
← Add NH_3
Remove NH_3 →
Add H_2 →
← Remove H_2
X Add a catalyst X

these "X's" mean NO EFFECT, catalysts do not favor forward or reverse

Most reactions only go "forward" Some reactions go in reverse"

Some reactions are dynamic equilibriums, and go both ways ⇌

ENTROPY is the measure of chaos in a chemical system.

Low entropy to high is SOLID → LIQUID → GAS

Potential Energy Diagrams
always label the reactants R, the products P, the activation energy AE, the activated complex AC, the change of energy ΔH, and the effect of a catalyst (with a dotted line inside the bell curve).

Catalysts increase the rate of reactions (or lower the time it takes to start reactions) by lowering the activation energy, and/or they provide an alternate pathway forward for the reaction.

Catalysts do not affect the collisions of particles.

Exothermic reactions have energy written with the products.

Endothermic reactions have energy written with the reactants.

Potential Energy Diagrams
show the flow of energy in a reaction. Reactants have the POTENTIAL to react, some more than others (helium balloons are fun, methane balloons are scary). Reactions require activation energy to react. All reactions pay some energy back. If they pay back only a little, the reaction is endothermic. If they pay back a lot, the products have less energy than the reactants did, the reaction is exothermic.

The difference between the potential of the reactants and the products is called the HEAT OF REACTION, or ΔH. Heat of Solution ΔS is the ΔH for solution formation