NAME

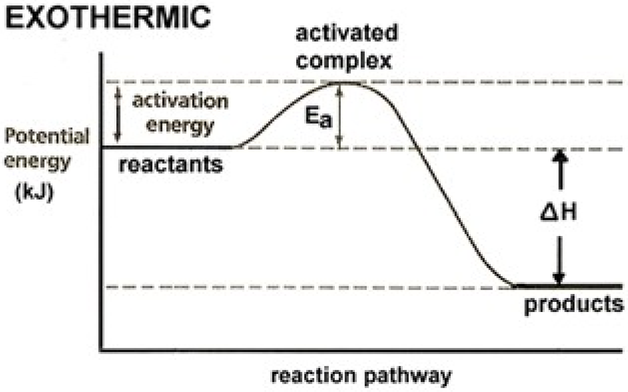
KINETICS PACK

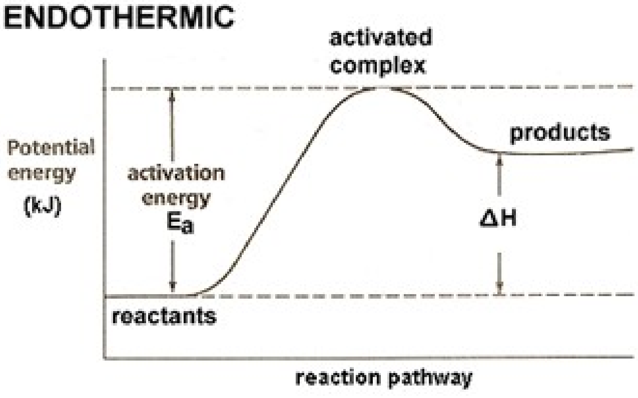
KINETICS BASICS & NOTES

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,
and 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, and also 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 as well.

Most reactions we have done and observed have been “one way” reactions. They are “IRREVERSIBLE” or the “I” in TOPIC-B. That is – the reactants become products, energy is either absorbed (endothermic) or released (exothermic) and the products that form are — more or less — permanent. The reason most reactions tend to go only one way is the energy involved is “used up and it’s no longer available.

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

Endothermic reactions absorb an exact amount of energy to go forward.
Exothermic reactions release a specific amount of energy as well.
This is all based upon the number of moles involved in the reactions. It is all quantifiable and measurable. We have Table I to tell us how much energy is released, or absorbed by many different reactions. This change in energy is called the DELTA “H”, or ΔH. The ΔH must always has a sign, positive for endothermic reactions, and negative sign for exothermic ones.

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

COLLISION THEORY of chemical reactions

It is understood that the atoms and molecules need to have enough kinetic energy to crash into each other, and be in the proper orientation, to react and combine into new substances. If the energy is insufficient, the atoms or molecules may bump into each other, but do not react. If the orientation is off, they also won’t combine. These collisions are the actual cause of chemical reactions to occur.

The more collisions, the faster the reaction. 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.

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 a reaction is directly related to the kinetic energy (temp) of the system.
Increase in Surface Area	More surface area leads to more collisions, and faster reactions. Smaller particles means “less waiting” for particles on the inside of larger particles to react after the outer ones do first.
Increase the Concentration	Stronger solutions have more ions or molecules per mL, leading to more collisions and of course, faster reactions. A lower molarity solute will lead to less collisions.
Catalysts	The use of catalysts do speed up chemical reactions, but they do not increase the collisions between particles. They do lower the energy needed to start the reaction (activation energy), or provide an alternate chemical pathway 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 will not discuss in our class 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 energy change when the solution forms. All combustion reactions are exothermic, energy is emitted as a product. All of their ΔH have negative signs. LOOK NOW at the 2nd line on the bottom of Table I. It says: "A minus sign indicates an exothermic reaction."

Energy is not negative, or positive. The negative sign is your note 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).

The six solution formations are also both positive or negative. Although these are not chemical reactions, they still have a ΔH value as they either absorb or emit energy when the solutions form.
Technically speaking, the ΔH here ARE NOT heats of reaction, these are not chemical reactions.
Here the ΔH are the HEATS OF SOLUTION.

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, this reaction emits 890.4 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 of the synthesis reactions can be reversed, that is you might combine CO and O_2 and form CO_2 which would release 566.0 kJ of energy.

Table I shows this: $2CO_{(G)} + O_{2(G)} \rightarrow 2CO_{2(G)}$ ΔH is –566.0 kJ/mole

Which means this: $2CO_{(G)} + O_{2(G)} \rightarrow 2CO_{2(G)} + 566.0\text{ kJ}$ (energy is emitted, exothermic)

But you can reverse this to: $2CO_{2(G)} + 566.0\text{ kJ} \rightarrow 2CO_{(G)} + O_{2(G)}$ (energy is absorbed, endothermic)

These reactions can all be "reversed" and the amount of energy is identical except that the sign for ΔH changes. Endothermic reactions reverse to exo. Exothermic reactions reverse to endo.

Find this one now… $NaCl_{(S)} \xrightarrow{\text{water}} Na^{+1}_{(AQ)} + Cl^{-1}_{(AQ)}$ ΔH is +3.88 kJ/mole

When table salt, sodium chloride dissolves and ionizes in water to form salt water, it will absorb a bit of energy and get a touch cooler (only 3.88 kJ per mole of NaCl). When salty water evaporates away, this is completely reversed.

$Na^{+1}_{(AQ)} + Cl^{-1}_{(AQ)} \rightarrow NaCl_{(S)} + H_2O_{(G)}$ *which evaporates away* ΔH is –3.88 kJ/mole

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 at all times during the chemical reaction. The 2 kinds are exothermic potential energy diagrams, and endothermic potential energy diagrams.

They are very similar (and different). You will need to understand them to tell them apart. It is important to see the difference, and recognize these differences so you can determine what the energy flow in a reaction is, just by looking over the diagram/graph.

Exothermic Potential Energy Diagrams

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 will 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 amount, and it is available in many in college texts, but we will NOT have to know the 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 instance, 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 go forward with insufficient activation energy.

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

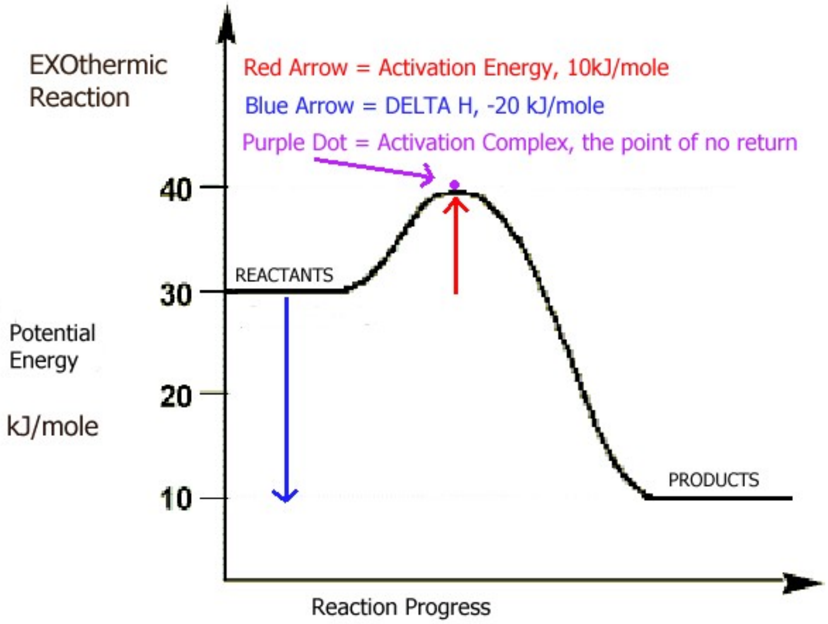
Here the products have less energy than the reactants started with. Where did that energy go? It was released, exothermically. It's not lost, it is released from the reaction into the environment.

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

This is the in-between stage when reactants are not really reactants anymore, but they are not quite yet the products. This activation complex forms exactly at the top of the curve of the potential energy diagram.
It lasts a nanosecond.

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

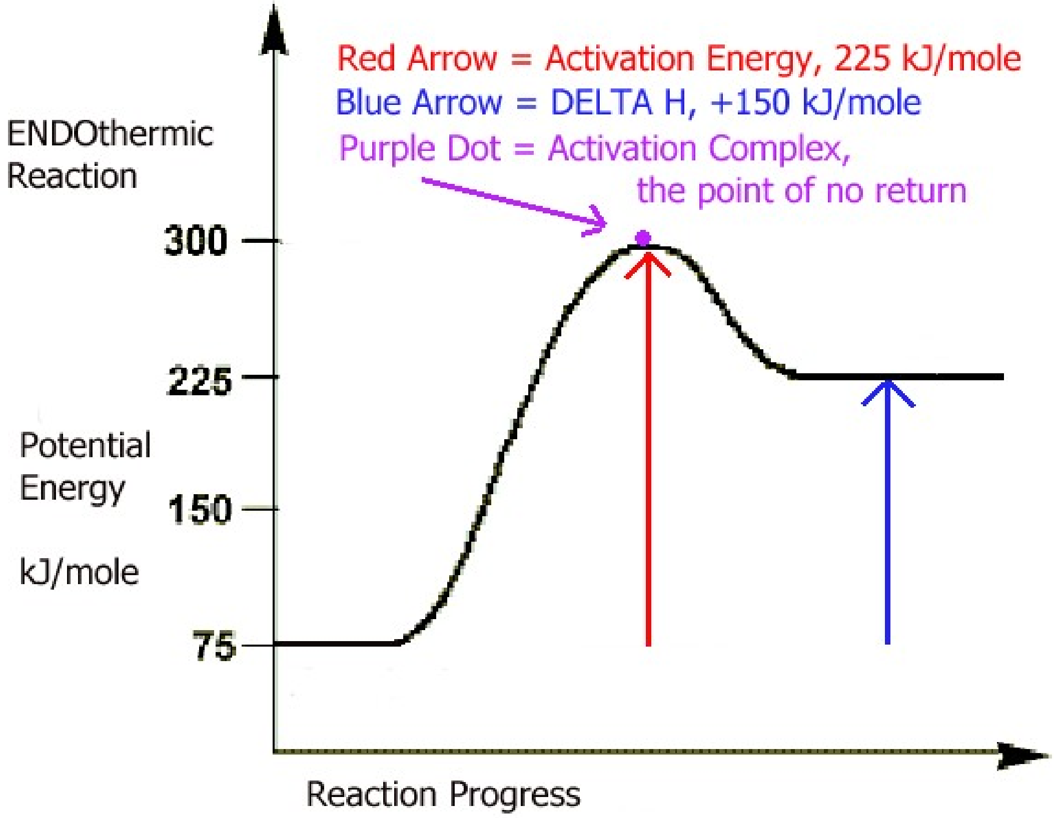
The ΔH is always negative in an exothermic reaction, because the products always have less potential energy than the reactants started with This extra energy is released into the environment.



The ENDOTHERMIC potential energy diagram

The ENDOTHERMIC potential energy diagram has all of the same parts, the big difference is that the products have MORE potential energy than the reactants did. In order 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.



Here 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 AC has 300 kJ of energy, but once the products start to form energy is released.

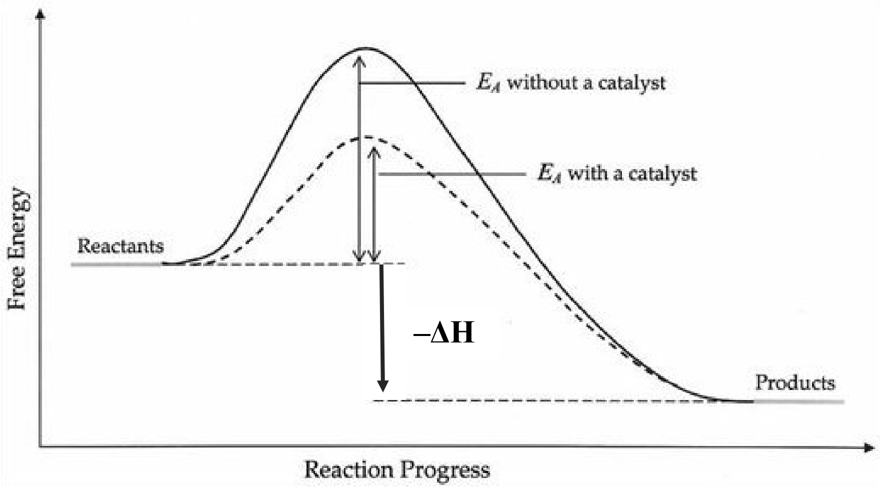
The products finish up with 225 kJ/mole of potential energy.

That's a difference of (net gain) of 150 kJ/mole.

The ΔH is +150 kJ/mole (that's 225 — 75 = 150)

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. We 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.

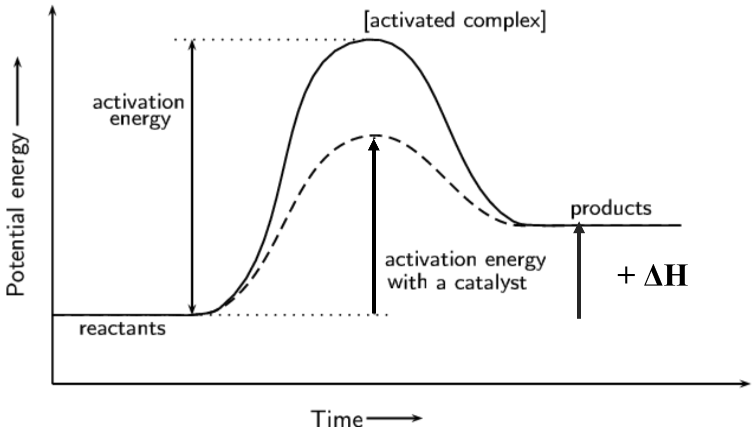


Here E_A means energy of activation.

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.

An inhibitor will of course inhibit or slow or even stop a chemical reaction.
Often they block the catalyst from working, but can also work by stopping 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: $\leftrightharpoons$

In our class ANY reaction that is provided with that 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. It will help you think.
The double arrow MEANS THAT THE REACTION IS REVERSIBLE in this moment.

The forward reaction below is the synthesis of AB. The forward reaction is endothermic as energy is a reactant. The reverse reaction is the decomposition of AB and it's exothermic as the reverse reaction releases energy as a product. The double arrows mean that the synthesis happens at a certain rate, which is matched by the decomposition.

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

The ΔH for the forward reaction is unknown exactly, but it has a $+\Delta H$
The reverse reaction has a $-\Delta H$, also unknown 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)

The addition of a catalyst will not change an equilibrium, catalysts just bring it to 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 aaccommodates 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. The famous "LeChatelier's Principle is this:

> "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 at the new conditions."

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

For example, this reaction will be examined in detail now.

$N_{2(G)} + 3H_{2(G)} \leftrightharpoons 2NH_{3(G)}$ + energy	The stresses added in the left column will cause this to happen because…
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, but they do get to equilibrium faster. We indicate this will a "double X" instead of two 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 can happen 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.

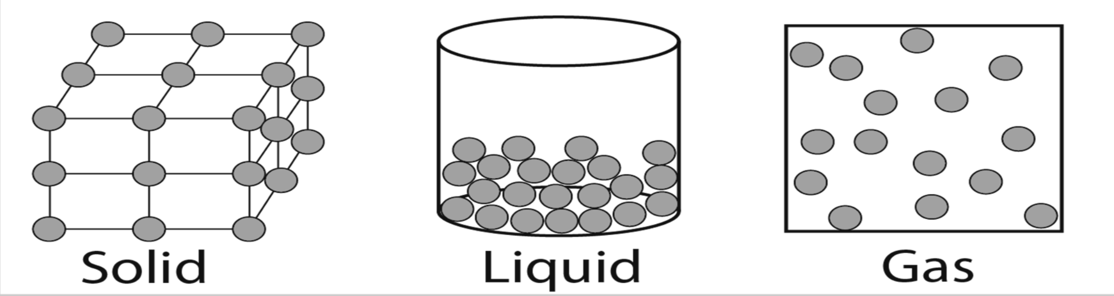
This reaction is technically NOT a real dynamic equilibrium, the forward reaction is very exothermic ($\Delta H = -3351$ kJ), but if the double arrow are present, we will use this more as a thinking exercise.
This is more of an exercise on LeChatelier's Principle than actual chemistry.
It's practice, and the shifting rules will be followed because of the double arrows.

$4Al_{(S)} + 3O_{2(G)} \leftrightharpoons 2Al_2O_{3(S)}$ + energy	The stresses added in the left column will cause this to happen because…
← 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 a complex math, but that is outside the scope of our class. Entropy for us is just going to be comparable.

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



Lowest Entropy	Medium Entropy	Highest Entropy
Ice has the lowest entropy because the moleculess are all stuck in place. They can "hum" but they don't really move around. They are well organized, and not chaotic.	Water has medium entropy. Liquids stay within the beaker, but the molecules are in constant motion.	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 two phases.

Another way to compare different substances and entropy is by particle size. Larger particles have a lower entropy than smaller particles when they are at the same conditions of pressure and temperature.

These three gases are all at the same Temperature and Pressure

octane C_8H_{18}	propane C_3H_8	methane CH_4
26 atoms bonded together LOWEST ENTROPY	11 atoms bonded together MEDIUM ENTROPY	5 atoms bonded together HIGHEST ENTROPY

A final 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.

Things are breaking down into smaller particles, the Universe is falling apart. Energy dissipates.

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 $\xrightarrow{\qquad\qquad\qquad}$ water + oxygen gas + Energy!
4	The CATALYST is not a
5	Kinetics is the part of chemistry that studies the
6	We will again examine reactions that absorb energy when they occur, called:
7	And their opposites, reactions that emit energy as a product, 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 It's ________________________________, because heat is a ________________________________ The ____ the _________ table I = ____________________
9	The forward reaction is __________THERMIC because energy is a ____________________________.
10	At the bottom of Table I it says that

11	$CH_{4(G)} + 2O_{2(G)} \rightarrow 2H_2O_{(G)} + CO_{2(G)} + \text{energy}$ 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? $2H_2O_{(L)} + \text{energy} \rightarrow O_{2(G)} + 2H_{2(G)}$ Here, energy is a __________________________, this reaction is _______________________________
15	Table I shows this… $2H_{2(G)} + O_{2(G)} \rightarrow 2H_2O_{(L)} + 571.6\text{ kJ}$ This is EXOTHERMIC! $2H_2O_{(L)} + 571.6\text{ kJ} \rightarrow O_{2(G)} + 2H_{2(G)}$ 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
$C_3H_{8(G)} + 5O_{2(G)} \rightarrow 3CO_{2(G)} + 4H_2O_{(G)}$		
$CO_{2(G)} \rightarrow C_{(S)} + O_{2(G)}$		
$NaOH_{(S)} \xrightarrow{\text{water}} Na^{+1}_{(AQ)} + OH^{-1}_{(AQ)}$		
$2NH_{3(G)} \rightarrow N_{2(G)} + 3H_{2(G)}$		

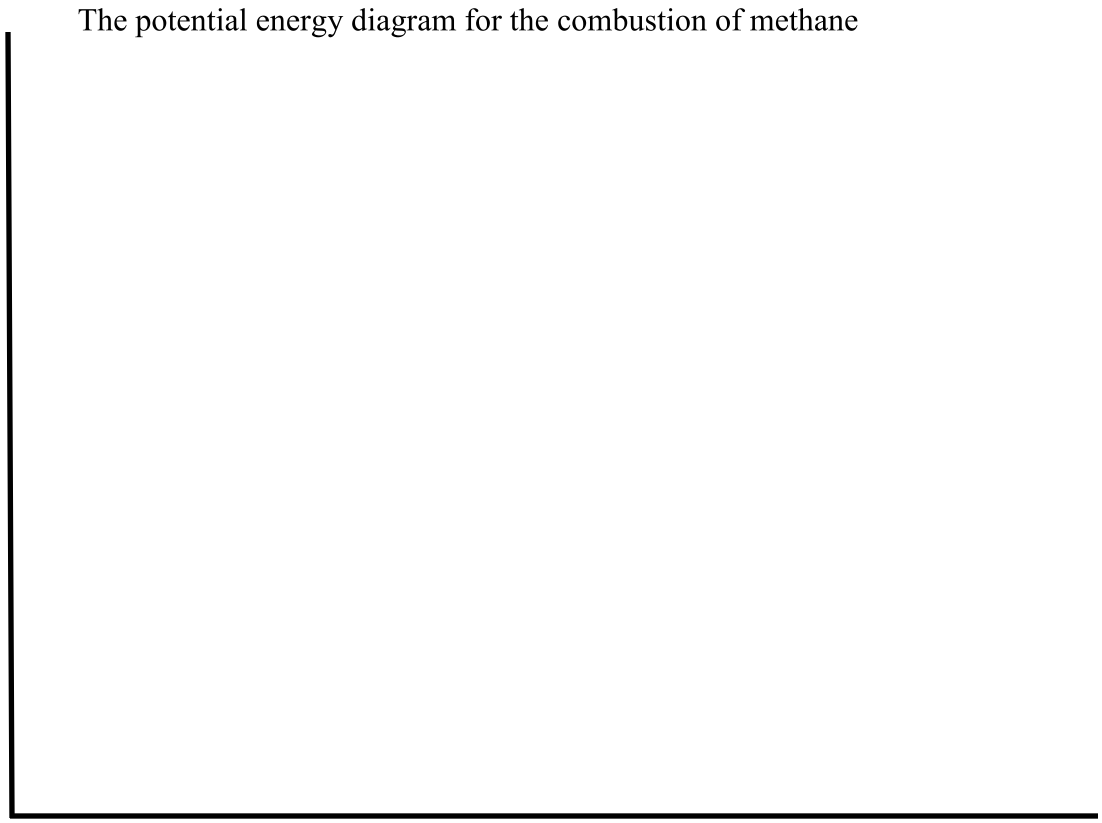
To drive from Vestal to Johnson City High School it’s about 8 miles. That should take about 20 minutes.	$\frac{8 \text{ miles}}{0.33 \text{ hours}}$ = ~ 24 mph
It you drive fast, getting there in just 10 minutes…	$\frac{8 \text{ miles}}{0.17 \text{ hours}}$ = ~ 47 mph
It you drive fast, getting there in just 5 minutes…	$\frac{8 \text{ miles}}{.083 \text{ hours}}$ = ~ 96 mph!

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	When you add a catalyst, it has no affect 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

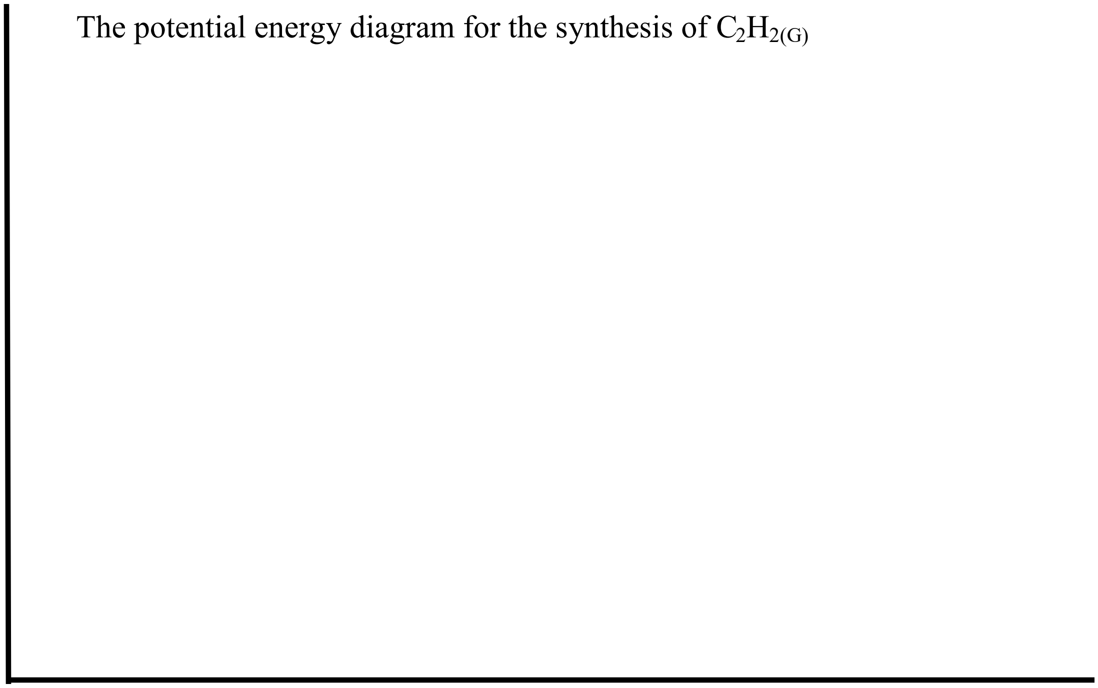
42. Draw the potential energy diagram showing the combustion of CH_4



43. Potential Energy Diagram vocabulary

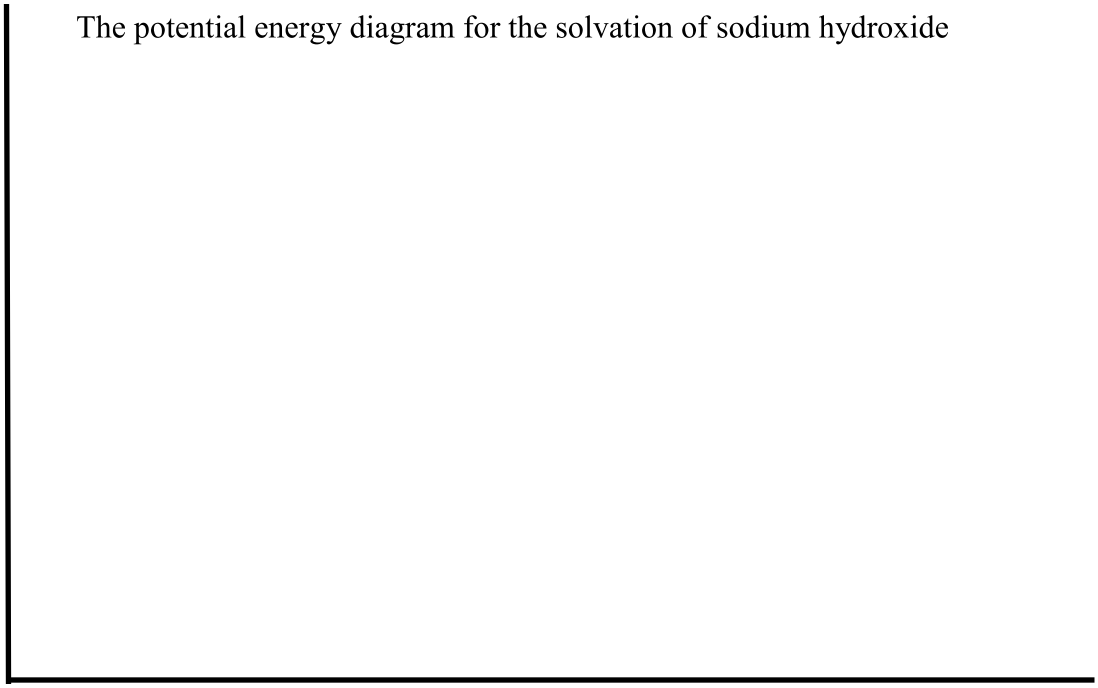
	the kJ/mole of energy that the reactants start with
	the kJ/mole of energy that the products end up with
	the difference in PE between reactants and products this PE diagram has a $-\Delta H$
	the energy required start the reaction (to get to the top of the hill)

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



The potential energy diagram for the synthesis of $C_2H_{2(G)}$

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



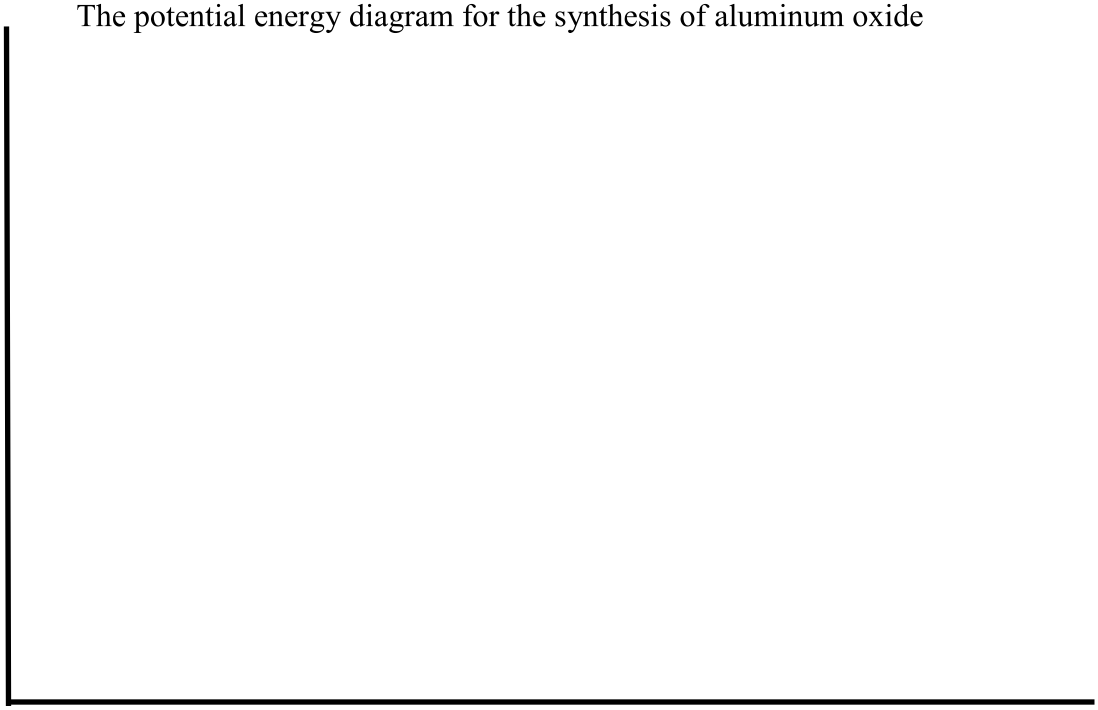
The potential energy diagram for the solvation of sodium hydroxide

46. This energy is the

47. More Kinetics Vocabulary	
	The energy stored in the bonds of the reactants or the products
	A transitional state; the reactants 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. 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.

The potential energy diagram for the synthesis of aluminum oxide



49. Draw the PE diagram for the solvation of sodium chloride. $NaCl_{(S)} + \text{energy} \xrightarrow{\text{water}} Na^{+1}_{(AQ)} + Cl^{-1}_{(AQ)}$

The potential energy diagram for the solvation of sodium chloride

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

The potential energy diagram for the combustion of propane

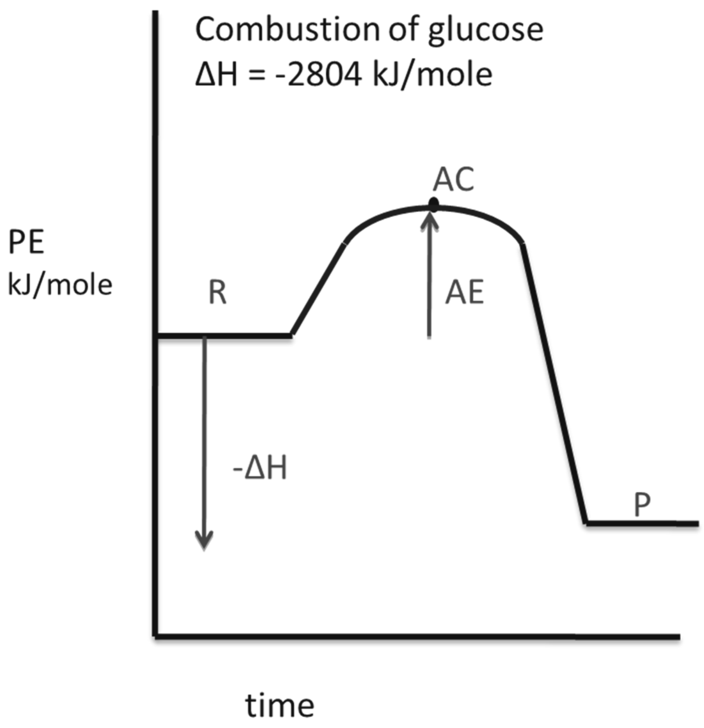
The affect of catalysts on chemical reactions

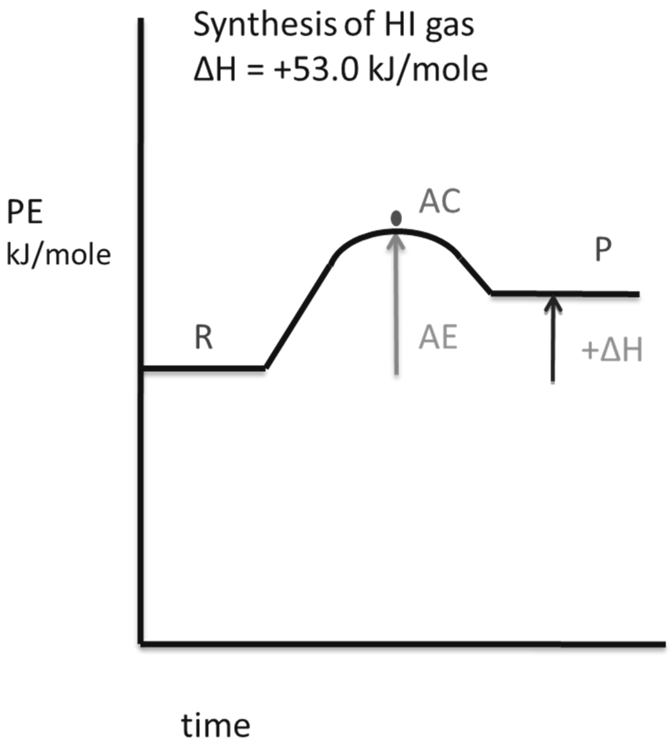
51. Draw the PE diagram for the combustion of glucose and for the synthesis of HI gas. Make 2 graphs, label PE of reactants and products, AE, AC, and ΔH (+ or -), and a title, with labels for the X and Y axis.

52	What is a Catalyst?
53	Catalysts…

54 **Indicate WHAT HAPPENS IF YOU ADD A CATALYST TO THESE REACTIONS?**

Exothermic at left   Endothermic at right





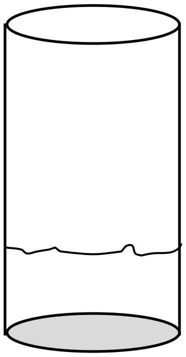
Do you know what R, P, ΔH, AC and AE mean?

55 Catalysts <u>DO NOT</u> change the number of collisions.
Catalysts are said to work

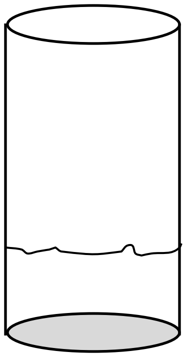
56 Catalysts ____of a reaction.
This lets the reaction start in a less energetic way, so it can happen quicker than normal. Once it starts, it moves right along.

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.

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
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	$N_2 + 3H_2 \leftrightarrow 2NH_3 + \text{energy}$ 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. 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 is 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}

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

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” because of the double arrows.

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

Add methane

Add water

Add heat

Remove carbon dioxide

Remove heat

Remove methane

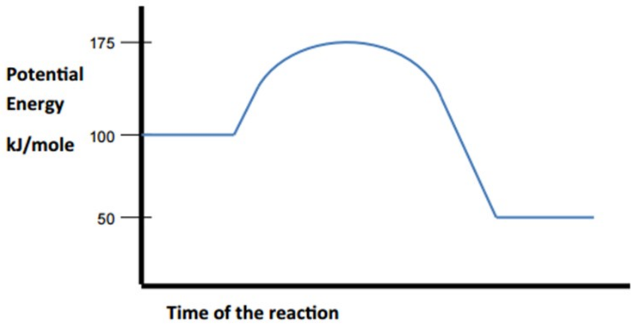
Add carbon dioxide

INCREASE PRESSUE

DECREASE PRESSURE

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

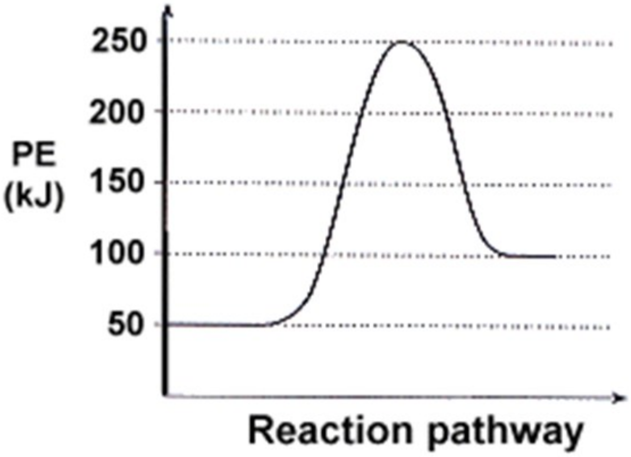
Use units where appropriate.



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	Of the 4 factors that would that speed up 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	

101. Three solid compounds are at the same temperature and pressure, which has the most entropy, which has the least entropy, and why?	
C_4H_{10} butane	
MgO magnesium oxide	
$C_{12}H_{22}O_{11}$ sucrose (sugar)	

102

103

104 $2H_{2(G)} + C_{(S)} \leftrightarrow CH_{4(G)}$ What's equal in a dynamic equilibrium?

105 Below, draw and label an exothermic and then endothermic potential energy diagram.

KINETICS AND EQUILIBRIUM

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.

The Law of Conservation of Energy... Know it?

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 a 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 is ΔH for solution formation.

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

Collision Theory

Anything that increases the rate and strength of collisions increases the rate of reaction
INCREASE
Temperature, Surface Area,
or the
Concentration of the Solutions,
makes the reaction rate ↑

Learn to draw exothermic + endothermic potential energy diagrams (draw on back). Always label the reactants R, products P, activation energy AE, activated complex AC, the change of energy ΔH, and the affect of a catalyst (shown with a dotted line <u>inside</u> the bell curve).

Most reactions only go "forward" →
← Some reactions can go in "reverse"
Some reactions are dynamic equilibriums,
they go both ways ⇌

Catalysts increase the rate of reactions, or lower the time it takes to start reactions, by being able to lower the activation energy, or they provide an alternate pathway forward for the reaction. They do not affect the collisions of particles.

Exothermic reactions have energy written with the products.

Endothermic reactions have energy written with the reactants.

ENTROPY is the measure of chaos in a chemical system. For us, it's easy.
Low entropy to high is SOLID → LIQUID → GAS.
Gases move fast + randomly, solids are stuck in place,
in between the liquids have medium entropy.
When comparing different compounds at the same temp + pressure,
bigger molecules with more bonds have less entropy,
smaller molecules have more entropy.

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