Chapter 1 : Le Chatelier's principle: Worked example (video) | Khan Academy

Le Chatelier's Principle In the French chemist and engineer Henry-Louis Le Chatelier proposed one of the central concepts of chemical equilibria. Le Chatelier's principle can be stated as follows: A change in one of the variables that describe a system at equilibrium produces a shift in the position of the equilibrium that counteracts the effect of this change.

While this is true in some cases, some chemical reactions are reversible, and we can take the products of the reaction and turn them back into the reactants. Equilibrium can be a tricky concept to understand, but this graphic tries to make it a little clearer. When we have a reversible reaction taking place in a closed system â€" that is, one where no substances are being added or lost â€" at the beginning of the reaction we will have only the reactants. Once the reaction starts, the amounts of the reactants will start decreasing, and the amounts of the products will begin to increase. In a non-reversible reaction this would be about the long and short of it, but when a reaction is reversible the products can also react to produce the reactants again. This is where the forwards reaction reactants reacting to produce the products and the backwards reaction products reacting to reform the reactants are occurring at the same rate. This means that the amounts of the reactants and the products remain constant, despite the fact that both reactions are still ongoing. If both are shovelling at the same rate, then the size of the hole and the size of the pile of dirt outside the hole will remain constant, despite the fact that both of them are still shovelling dirt back and forth. Dynamic equilibrium is much the same. Though the amounts of the reactants and products may initially be the same once we reach dynamic equilibrium, we can make adjustments to the conditions of the equilibrium to change the proportions of reactants and products in the equilibrium mixture. For example, if we increase the temperature of the reaction, it will respond in a way that decreases the temperature. If we increase the concentration of a reactant, it will respond in a way that decreases the concentration of that reactant. On the face of it, this seems quite simple. However, applying it can be a little trickier than it initially seems. Concentration is simply a measure of how much of a substance we have in a particular volume. Say we increase the concentration of one of the reactants; the equilibrium can reduce its concentration by favouring the forwards reaction and producing more of the products. The net result of increasing the concentration of the reactants would be the production of more of the products at equilibrium. Temperature Changing the temperature can also affect equilibrium position. How it does this and whether it favours the reactants or the products will depend on the reaction. Chemical reactions can be either exothermic give out heat or endothermic take in heat. If the energy required to break bonds is less than the energy released when forming new bonds, the reaction will be exothermic. If the energy required to break bonds is more than the energy released when forming new bonds, the reaction is endothermic. For reversible reactions, either the forwards or backwards reaction will be exothermic, and the other will be endothermic. When we increase the temperature, the reaction will favour whichever reaction is endothermic to take in heat and reduce the temperature. On the other hand, if we decrease the temperature, the exothermic reaction will be favoured, as this will give out heat and increase the temperature. Pressure Changing the pressure of a reaction involving gases can also affect the position of equilibrium. Pressure is caused by the collisions of gas particles with the walls of the container. It follows, then, that the greater the number of gas particles, the higher the pressure will be â€" and this gives us a hint as to the effect changing pressure has on the equilibrium position. It can do this by favouring the side of the reaction with fewer gas molecules; which side this is will of course depend on the reaction in question. Conversely, if we decrease the pressure, the equilibrium will respond to increase the pressure, and will therefore favour the side of the reaction with more gas molecules. As a caveat, there are plenty of reactions where we actually have the same number of gaseous molecules shown in the balanced equation. In this case, changing the pressure will have no effect, as it will not favour either side of the reaction. Catalysts Catalysts are often used by chemists to increase the speed of chemical reactions. This is because it speeds up both the forwards and backwards reaction by an equal amount,

so overall neither side of the reaction is favoured by using one. Why does this all matter? The problem with the reaction is that it is reversible; therefore some tweaking is required to produce as much ammonia as possible. The ammonia that is produced is also removed from the reactor, so that the equilibrium responds by shifting to produce more ammonia. A iron catalyst is often used to make the reaction happen faster. In terms of temperature, the forwards reaction nitrogen and hydrogen to ammonia is exothermic. This is by no means a low temperature, so why is it used? Well, temperature also affects the rate of the reaction, as detailed in a previous graphic here â€" if the temperature is too low, the reaction will happen too slowly! Finally, the balanced equation for the reaction shows that there are more molecules of gas on the left-hand side of the equation 4 than there are on the right-hand side 2. Consequently, the reaction is favoured by a high pressure, as this shift the equilibrium towards the side with fewer molecules of gas, producing more ammonia. For this reason the Haber process is usually carried out at pressures around atmospheres. Again, there is a degree of compromise here; higher pressures would produce even more ammonia, but would require more expensive equipment and be more costly to maintain. These are better at aiding predictions, as they can avoid some of the confusion that the Principle can sometimes result in.

Page 1 of 4 Chemical Equilibrium and Le Chatelier's Principle Objectives The objective of this lab is to observe the effect of an applied stress on chemical systems at equilibrium.

Thus over time the forward reaction slows down. Eventually a point will be reached where the rate of the forward reaction will be equal to the rate of the backward reaction. When this occurs, a state of chemical equilibrium is said to exist. Chemical equilibrium is a dynamic state. A reversible reaction at equilibrium can be disturbed if a stress is applied to it. Examples of stresses include increasing or decreasing chemical concentrations, or temperature changes. If such a stress is applied, the reversible reaction will undergo a shift in order to re-establish its equilibrium. Consider a hypothetical reversible reaction already at equilibrium: Eventually the forward reaction would slow down and the forward and backward reaction rates become equal again as the system returns to a state of equilibrium. Using similar logic, the following changes in concentration are expected to cause the following shifts: In other words, if a chemical is added to a reversible reaction at equilibrium, a shift away from the added chemical occurs. When a chemical is removed from a reversible reaction at equilibrium, a shift towards the removed chemical occurs. A change in temperature will also cause a reversible reaction at equilibrium to undergo a shift. The direction of the shift largely depends on whether the reaction is exothermic or endothermic. In exothermic reactions, heat energy is released and can thus be considered a product. In endothermic reactions, heat energy is absorbed and thus can be considered a reactant. In this lab, the effect of applying stresses to a variety of chemical systems at equilibrium will be explored. The equilibrium systems to be studied are given below: Procedure Materials and Equipment Equipment: If any of these chemicals spill on you, immediately rinse the affected area under running water and notify your instructor. These spots will eventually fade after repeated rinses in water. Finally, in Part 4 you will be heating a solution in a test tube directly in a Bunsen burner flame. If the solution is overheated it will splatter out of the tube, so be careful not to point the tube towards anyone while heating. Experimental Procedure Record all observations on your report form. These should include, but not be limited to, color changes and precipitates. Note that solution volumes are approximate for all reactions below. Dispose of all chemical waste in the plastic container in the hood. Acidified Chromate Solution Place 3-mL of 0. Here the added sodium hydroxide is effectively removing acidic hydrogen ions from the equilibrium system via a neutralization reaction: Aqueous Ammonia Solution Instructor Prep: At the beginning of lab prepare a stock solution of aqueous ammonia. Label the beaker and place it on the front desk. The entire class will then use this stock solution in Part 3. Place 3-mL of the prepared stock solution into a small test tube. Label these test tubes The solution in test tube 1 remains untouched. It is a control for comparison with other tubes. Then heat this solution directly in your Bunsen burner flame moderate temperature. Then cool the solution in test tube 3 back to room temperature by holding it under running tap water, and again record your observations. At the beginning of lab prepare a stock solution of iron III thiocyanate. Add 1-mL of 0. The entire class will then use this stock solution in Part 5. Place 3-mL of the prepared stock solution into 4 small test tubes. To the solution in test tube 2, add 1-mL of 0. To the solution in test tube 3, add 1-mL of 0. To the solution in test tube 4, add 0. A light precipitate may also appear. Here the added silver nitrate is effectively removing thiocyanate ions from the equilibrium system via a precipitation reaction: What happens to the forward and reverse reaction rates when equilibrium is achieved? Consider the following exothermic reversible reaction at equilibrium: These are supplied in the Theory Section. Consider the third system you will study: Write the balanced equation for this reversible reaction. Suppose you added some excess ammonium ions to this system at equilibrium. In which direction would a shift occur? What color change might you expect to observe? List all the equipment you will use in this lab.

Chapter 3 : What Is Chemical Equilibrium? | Chemical Equilibrium | Siyavula

Le Chatelier's principle (UK: / l ɜ™ ʃ ã¦ ʌˆ t ɚ› l j eɪª /, US: / ʌˆ ʃ ɪʼ�ː• t ɜ™l j eɪª /), also called Chatelier's principle or "The Equilibrium Law", can be used to predict the effect of a change in conditions on some chemical equilibria.

Half-fill two beakers with water and mark the level of the water in each case. Cover one of the beakers with a glass cover. Leave the beakers and, over the course of a day or two, observe how the water levels in the two beakers change. What do you notice? You could speed up this demonstration by placing the two beakers over a Bunsen burner, or in direct sunlight, to heat the water. Observations You should notice that in the beaker that is uncovered, the water level drops more quickly than in the covered beaker. This is because of evaporation. In the beaker that is covered, there is an initial drop in the water level, but after a while evaporation appears to stop and the water level in this beaker is higher than that in the one that is open. Discussion In the first beaker, liquid water becomes water vapour as a result of evaporation and the water level drops. A small amount of gas molecules will condense again, but because the gas molecules can escape from the system there is much less condensation than evaporation. In the second beaker, evaporation also takes place. However, in this case, the vapour comes into contact with the surface of the glass cover and it cools and condenses to form liquid water again. This water is returned to the beaker. Once condensation has begun, the rate at which the water level drops will start to decrease. At some point, the rate of evaporation will be equal to the rate of condensation, and there will be no change in the water level in the beaker. This can be represented as follows: In the forward direction there is a change in phase from liquid to gas, present here as water vapour. A reverse change can also take place, when vapour condenses to form liquid again. Evaporation is when a substance goes from the liquid phase to the gas phase it evaporates. Condensation is when a substance goes from the gas phase to the liquid phase it condenses. The bottle contains water vapour gas and water droplets are condensing on the side of the bottle. An open system is one in which matter or energy can flow into or out of the system. In the liquid-gas demonstration we used, the first beaker was an example of an open system because the beaker could be heated or cooled a change in energy , and water vapour the matter could evaporate from the beaker. Open system An open system is one whose borders allow the movement of energy and matter into and out of the system. A closed system is one in which energy can enter or leave, but matter cannot. The second beaker with the glass cover is an example of a closed system. The beaker can still be heated or cooled, but water vapour cannot leave the system because the glass cover is a barrier. Condensation changes the vapour to liquid and returns it to the beaker. In other words, there is no loss of matter from the system. Generally, some of the molecules at the surface of liquids and solids are moving into the gas phase. This means that matter is leaving the system. However, this is such a small fraction of the total volume of the liquid or solid that a reaction involving only solids or liquids can be considered a closed system. It is useful to simplify situations in science by dividing the world into the system we are studying, and the surrounding environment that might influence the reaction, but is not part of it. Closed system A closed system is one in which only energy can move into and out of the system. Matter cannot be gained by the system nor lost from the system. In a closed system it is possible for reactions to be reversible, such as in the demonstration above. In a closed system, it is also possible for a chemical reaction to reach equilibrium. In one direction the reactants combine to form the products. In the other direction, the products react to form the reactants again. In other words, the reactants and products of the reaction may reverse roles. So, in the following reversible reaction: The forward reaction is always taken from the given equation and written: The reverse reaction is always taken from the given equation and written: Chemical equilibrium Which of the following situations describes a closed system? The water vapour can leave the pot system if there is no lid. Assume that none of the solid or liquid reactants or products go into the gas phase. As the reaction gets closer to equilibrium the rate of decrease levels out until the forward reaction has a constant rate. As the reaction progresses the rate of increase levels out until the reverse reaction has a constant rate. Although it is not always possible to observe any macroscopic changes, this does not mean that

the reaction has stopped. The forward and reverse reactions continue to take place and so microscopic changes still occur in the system. This state is called dynamic equilibrium. Dynamic equilibrium There is a dynamic equilibrium in a reversible reaction when the rate of the forward reaction equals the rate of the reverse reaction. The amounts of reactants and products remain constant. In the liquid-gas phase equilibrium demonstration, dynamic equilibrium was reached when there was no observable change in the level of the water in the second beaker even though evaporation and condensation continued to take place. For more information on dynamic equilibriums watch this video:

Chapter 4 : Le Chatelier's principle (video) | Khan Academy

Examples of Chemical Equilibria mL samples of M NH 3, add a drop of phenolphthalein calendrierdelascience.com and record the color. To one sample, add 1 M NH 4Cl, a few drops at a time, with mixing.

When a dynamic equilibrium is achieved there is a state of balance between the constant concentrations of the reactants and products because the rate at which the reactants change into products is exactly equal to the rate at which the products change back to the original reactants. However, the actual relative amounts of the original reactants left, and products formed, at equilibrium, depend on the particular reaction and reaction conditions e. A catalyst does not affect the position of the equilibrium, i. At a given constant temperature, all the final equilibrium concentrations are mathematically governed by the equilibrium expression and the equilibrium constant and these are dealt with in detail in Equilibria Part 2. Also, when considering the equilibrium rules outlined below, any change affects BOTH the rates of the forward and backward reactions. The system absorbs the heat energy from the surroundings to try to minimise the temperature increase. The system releases heat energy to the surroundings to try to minimise the temperature decrease. Increasing the pressure favours the side of the equilibrium with the least number of gaseous molecules as indicated by the balanced symbol equation. The system attempts to reduce the number of gas molecules present to reduce the pressure increase. Decreasing the pressure favours the side of the equilibrium with the most number of gaseous molecules as indicated by the balanced symbol equation. The system attempts to increase the number of gas molecules to minimise the pressure decrease. If there is NO net change in the number of gas molecules, changing total gas pressure has NO effect on the position of the equilibrium, though pressure increase effectively increases gas concentration so both the forward and backward reactions will be speeded up. Over and above rule 2, all the individual partial pressures of the gases, must comply with the mathematics of the Kp equilibrium expression described and explained in Equilibria Part 2. The rest of the rules 1, 3 and 4 apply to ANY reaction, whatever the physical states of the reactants and products. If the concentration of a reactant on the left is increased, then some of it must change to the products on the right to maintain a balanced equilibrium position. This means if you change ANY concentration, all the other concentrations must change too see Example 1. Also, any net concentration changes must comply with the Kc equilibrium expression fully explained in Equilibria Part 2. BUT it does is enable you get to the point of equilibrium faster! The importance of a catalyst lies with economics of chemical production e. So remember a catalyst cannot affect the position of the equilibrium constant or the value of the equilibrium constant K Kc or Kp. For more details of the industrial process see Equilibria Part 3. Rule 4 â€" catalyst: Rule 3 â€" concentration: If the nitrogen or hydrogen concentration was decreased, some of ammonia would change to nitrogen and hydrogen. H2 mixture is used in industry, but for academic reasoning practice, specifically More details of the Haber process are given in Equilibria Part 3. He made some of the first accurate calculations of equilibrium constants. The formation of hydrogen iodide from hydrogen and iodine: Rule 3 â€" concentration e. Please note that there would still be an overall increase in iodine at the new equilibrium point. To study the reaction, Bodenstein sealed various amounts of HI in glass containers. These were then heated to the desired temperature for various intervals of time in an accurately thermostated oven. The glass containers were then removed and cooled rapidly, broken open and the contents analyzed for all three components. From one molar measurement and the starting quantities, you can work out the quantities of the other components in the final equilibrium mixture. Not applicable, no gases involved. Rule 3 â€" concentration Sometimes it is desirable to add a large excess of the alcohol to ensure most of the acid is converted into ester. The alcohol may be cheaper than the acid and the ester and unreacted alcohol separated by distillation and the latter recycled. Rule 4 â€" catalyst The forward esterification reaction is catalysed by acids e. The reverse reaction i. Preparation of nitrogen IV oxide You can conveniently prepare a sample of nitrogen IV oxide by heating lead II nitrate - which crackles! The temperature effect can be observed by strongly heating the gases in the pyrex tube above oC. The brown

colour should fade at the lowest hottest part of the boiling tube. The equilibrium will shift from left to right, in the endothermic direction - absorbing the heat to counteract the increase in temperature. This experiment can be done after you have done the lower temperature boiling tube experiment below. Cool the tube in ice and the brown colour is considerably decreased - equilibrium shifts in the exothermic direction and more colourless N2O4 formed - heat released to counteract the decrease in temperature. Decrease in temperature favours the exothermic formation of the dimer N2O4 from NO2, so the brown colour fades on cooling the gas mixture. You can do this on an OHP, take care, its a bit awkward for all of an A level class to see it at close quarters in the confines of a fume cupboard. When you press the gas syringe plunger in, you increase the pressure of the gases. The orange-brown colour becomes more darker because ALL concentrations are increased. BUT, within a few seconds the orange-brown colour fades just a little bit as the equilibrium moves from left to right - a little more N2O4 formed - as predicted from the pressure rule. When you pull the gas syringe plunger out, you decrease the pressure of the gases. The orange-brown colour becomes lighter because ALL concentrations are decreased. BUT, within a few seconds, the colour gets a little bit darker as the equilibrium moves from right to left - a little more NO2 formed - as predicted from the pressure rule. If prepared at higher temperature, with just enough chloride added to turn the solution blue, on cooling it becomes pink. Rule 3 â€" concentration Increase in chloride concentration decreases the pink ion concentration and increases the blue ion concentration i. Diluting with water shifts the equilibrium to the left so solution is less blue and more pink. Not applicable here, despite involving transition metal complexes! If you dissolve copper II chloride in water you get a greenishâ€"blue solution as both copper II complexes are present in equilibrium. By adding water i.

Chapter 5 : Le Chatelier's Principle - Chemistry LibreTexts

Using Le Chatelier's Principle. A statement of Le Chatelier's Principle. If a dynamic equilibrium is disturbed by changing the conditions, the position of equilibrium moves to counteract the change. Using Le Chatelier's Principle with a change of concentration. Suppose you have an equilibrium established between four substances A, B, C and D.

Draw a conclusion about the effect of a change in concentration of either the reactants or products on the equilibrium position. Also draw a conclusion about the effect of a change in temperature on the equilibrium position. If the pressure is increased the equilibrium will shift to favour a decrease in pressure. If the pressure is decreased the equilibrium will shift to favour an increase in pressure. When the volume of a system is decreased and the temperature is constant , the pressure will increase. There are more collisions with the walls of the container. If there are fewer gas molecules there will be fewer collisions, and therefore lower pressure. The equilibrium will shift in a direction that reduces the number of gas molecules so that the pressure is also reduced. Remember from Grade 11 that: For example, the equation for the reaction between nitrogen and hydrogen is shown below: An increase in pressure will: Favour the reaction that decreases the number of gas molecules. A decrease in pressure will: Favour the reaction that increases the number of gas molecules. Consider illustration in Figure 8. In the original system there are 12 molecules in total: If you increase the pressure shown by a decrease in volume , the equilibrium will shift to decrease the number of gas molecules. Note that the total number of nitrogen and hydrogen atoms remains the same in all three situations. Equations a and b are not balanced equations. Another example is the reaction between sulfur dioxide and oxygen: The effect of a catalyst on equilibrium ESCNS If a catalyst is added to a reaction, both the forward and reverse reaction rates will be increased. If both rates are increased then the concentrations of the reactants and products will remain the same. This means that a catalyst has no effect on the equilibrium position. However, a catalyst will affect how quickly equilibrium is reached. This is very important in industry where the longer a process takes, the more money it costs. So if a catalyst reduces the amount of time it takes to form specific products, it also reduces the cost of production. Changing the concentration of a reactant or product results in one of the reactions forward or reverse being favoured. This change in reaction rate minimises the effect of the change and restores the concentration ratio between reactants and products. There will just be more reactants and products. Changing the pressure of the system will change the ratio between the reactant and product concentrations. The equilibrium then shifts to minimise the effect of the change and restores the ratio between reactant and product concentrations. Both the forward and reverse reactions rates are increased. Therefore the ratio between reactant and product concentrations will remain the same. Changing the temperature will favour either the endothermic or exothermic reaction. The ratio between the concentration of the reactants and products will change. Identify the disturbance or stress on the system. Look at the given equation and decide whether the rate of the forward reaction or the rate of the reverse reaction is increased. State the shift in equilibrium. Where appropriate, link equilibrium shift to any observed change in the system. What would be observed? The reverse reaction is favoured and the equilibrium will shift to the left. What would the colour change be due to this equilibrium shift? The following are some points to keep in mind when presented with a graph. Identify the type of graph by looking at the label on the y-axis. You will find either: For concentration-time graphs or mole-time graphs equilibrium occurs where the concentration or number of moles of the reactants and products are constant. These values need not be equal to one another. Rate-time graphs A change in concentration of a substance would favour the reaction that decreases the amount of that substance. This will appear as a sharp increase in the rate of either the forward or reverse reaction and a sharp decrease in the rate of the other reaction. The increased rate will then gradually decrease and the decreased rate will gradually increase until they are equal again. A change in pressure of the reaction would cause a sharp increase or decrease in all the reactants and products. This will have the same effect as a change in concentration, although the increase or decrease would be more gradual. A change in temperature of a

substance would affect both rates in the same direction either both increase or both decrease. However, the effect will be unequal, with the endothermic reaction favoured by an increase in temperature, and the exothermic reaction favoured by a decrease in temperature. The addition of a catalyst would favour both the forward and reverse reactions by the same amount. What stress has occurred in this system? Label the graph with what is happening at each stage. Check the axes so that you know what the variables are on this graph The axes are labelled rate and time. Therefore this is a rate-time graph. Are the rates both affected in the same way or is one rate increased and the other decreased when the stress is applied? Both rates are affected in the same way increased therefore the stress must be a catalyst or a change in temperature. A change in pressure or concentration would favour one reaction direction only Are both rates affected equally? No, the forward rate is increased more than the reverse rate. Therefore the stress must be a change in temperature. A catalyst would increase both rates equally Was the temperature increased or decreased? The forward reaction was favoured more than the reverse reaction. An increase in temperature will favour the reaction that cools the reaction vessel the endothermic reaction. Therefore the stress must have been an increase in temperature. Label what is occuring at each stage on the graph Worked example 8: Yes, both rates are increased by the same amount. The addition of a catalyst a change in temperature would affect both rates, but unequally. Concentration-time and mole-time graphs A change in concentration of a substance would appear as a sharp increase or decrease in the concentration or number of moles of that substance and a gradual change in the other substances. A change in temperature would affect both the forward and reverse reactions. However, one reaction would be affected more than the other. So the reactants and products would be affected gradually, in the opposite direction one increased, the other decreased. For an increase in pressure, if the forward reaction is then favoured the reactant concentrations will decrease, and if the reverse reaction is then favoured the product concentrations will decrease. The addition of a catalyst would increase both the forward and reverse reaction rates, meaning the equilibrium is reached faster. So if the reaction is already at equilibrium there will be no effect on a concentration-time or moles-time graph. Concentration-time graphs Consider the following chemical equilibrium and graph and answer the questions that follow. How do the rates of the forward and reverse reactions compare at the following times: Check the axes so that you know what the variables are on this graph The axes are labelled concentration and time. Therefore this is a concentration-time graph. How much time was necessary for the system to reach equilibrium? This means that the reaction has reached equilibrium. How can you determine which rate is faster from the concentration graphs? Therefore the rate of the forward reaction is faster than the rate of the reverse reaction. Therefore the reaction is in equilibrium and the rate of the forward reaction equals the rate of the reverse reaction. Reading off the graph you can see that:

Chapter 6 : LeChatelier's principle in equilibria

1 Chemical Equilibria, LeChatelier's Principle v You are encouraged to carefully read the following sections in Tro (2nd ed.) to prepare for this experiment: Sec (LeChatelier's principle), pp - 3 (indicators).

The duration of adjustment depends on the strength of the negative feedback to the initial shock. Where a shock initially induces positive feedback such as thermal runaway , the new equilibrium can be far from the old one, and can take a long time to reach. In some dynamic systems , the end-state cannot be determined from the shock. The principle is typically used to describe closed negative-feedback systems, but applies, in general, to thermodynamically closed and isolated systems in nature, since the second law of thermodynamics ensures that the disequilibrium caused by an instantaneous shock must have a finite half-life. The principle while well rooted in chemical equilibrium and extended into economic theory, can also be used in describing mechanical systems in that the system put under stress will respond in a way such as to reduce or minimize that stress. Moreover, the response will generally be via the mechanism that most easily relieves that stress. Chemistry[edit] Effect of change in concentration[edit] Changing the concentration of a chemical will shift the equilibrium to the side that would reduce that change in concentration. The chemical system will attempt to partly oppose the change affected to the original state of equilibrium. In turn, the rate of reaction, extent, and yield of products will be altered corresponding to the impact on the system. This can be illustrated by the equilibrium of carbon monoxide and hydrogen gas, reacting to form methanol. If we are to add a species to the overall reaction, the reaction will favor the side opposing the addition of the species. Likewise, the subtraction of a species would cause the reaction to "fill the gap" and favor the side where the species was reduced. This observation is supported by the collision theory. As the concentration of CO is increased, the frequency of successful collisions of that reactant would increase also, allowing for an increase in forward reaction, and generation of the product. Even if the desired product is not thermodynamically favored, the end-product can be obtained if it is continuously removed from the solution. The effect of a change in concentration is often exploited synthetically for condensation reactions i. This can be achieved by physically sequestering water, by adding desiccants like anhydrous magnesium sulfate or molecular sieves, or by continuous removal of water by distillation, often facilitated by a Dean-Stark apparatus. Effect of change in temperature[edit] The effect of changing the temperature in the equilibrium can be made clear by 1 incorporating heat as either a reactant or a product, and 2 assuming that an increase in temperature increases the heat content of a system. Hence, whether increasing or decreasing the temperature would favor the forward or the reverse reaction can be determined by applying the same principle as with concentration changes. Take, for example, the reversible reaction of nitrogen gas with hydrogen gas to form ammonia: More ammonia would be produced if the reaction was run at a lower temperature, but a lower temperature also lowers the rate of the process, so, in practice the Haber process the temperature is set at a compromise value that allows ammonia to be made at a reasonable rate with an equilibrium concentration that is not too unfavorable. In exothermic reactions , increase in temperature decreases the equilibrium constant , K, whereas, in endothermic reactions , increase in temperature increases the K value. The effect of temperature on equilibria, however, involves a change in the equilibrium constant. The value of K changes with temperature. When heat is added and the temperature increases, the reaction shifts to the right and the flask turns reddish brown due to an increase in NO2. When heat is removed and the temperature decreases, the reaction shifts to the left and flask turns colorless due to an increase in N2O4. Effect of change in pressure[edit] The equilibrium concentrations of the products and reactants do not directly depend on the total pressure of the system. They may depend on the partial pressures of the products and reactants, but if the number of moles of gaseous reactants is equal to the number of moles of gaseous products, pressure has no effect on equilibrium. Effect of change in volume[edit] Changing the volume of the system changes the partial pressures of the products and reactants and can affect the equilibrium concentrations. With a pressure increase due to a decrease in volume, the side of the equilibrium with fewer

moles is more favorable [4] and with a pressure decrease due to an increase in volume, the side with more moles is more favorable. There is no effect on a reaction where the number of moles of gas is the same on each side of the chemical equation. Considering the reaction of nitrogen gas with hydrogen gas to form ammonia: When the volume of the system is changed, the partial pressures of the gases change. If we were to decrease pressure by increasing volume, the equilibrium of the above reaction will shift to the left, because the reactant side has a greater number of moles than does the product side. The system tries to counteract the decrease in partial pressure of gas molecules by shifting to the side that exerts greater pressure. Similarly, if we were to increase pressure by decreasing volume, the equilibrium shifts to the right, counteracting the pressure increase by shifting to the side with fewer moles of gas that exert less pressure. If the volume is increased because there are more moles of gas on the reactant side, this change is more significant in the denominator of the equilibrium constant expression, causing a shift in equilibrium. Effect of adding an inert gas[edit] See also: Adding an inert gas into a gas-phase equilibrium at constant volume does not result in a shift.

Chapter 7 : Reversible Reactions, Equilibrium, and Le Châtelier's Principle | Compound Interest

Le Chatlier's principle (also known as "Chatelier's principle" or "The Equilibrium Law") states that when a system experiences a disturbance (such as concentration, temperature, or pressure changes), it will respond to restore a new equilibrium state.

Firstly, every reversible reaction reaches its own specific equilibrium under a given set of conditions. This equilibrium state is dependent on Temperature of the reacting system; Pressure of the reacting system for gases ; Concentration of the reacting system. A change in any one of these factors will upset the balance of the system and result in a shift in the position of the equilibrium. Conversely, if the temperature is raised, then the equilibrium position will shift so as to annul the heating effect by absorbing more heat. In the following reversible reaction, the forward reaction is endothermic. Hence, the value of the equilibrium constant K also increases. So, for a given reversible reaction, a higher value of K would mean a greater yield of product s. Lowering the temperature of the system will shift the equilibrium position to the left, favouring the backward reaction, i. This results in a lowering of K So, for a given reversible reaction a lower value of K would mean a greater yield of reactant s. Conversely, a decrease in temperature will cause the equilibrium position to shift to the right, i. Dinitrogen IV oxide molecules are pale yellow, whereas nitrogen IV oxide molecules are reddish-brown in colour. At low temperatures, dinitrogen IV oxide predominates while at higher temperatures, it dissociates to give the monomer. If a sealed glass vessel containing an equilibrium mixture of the two is warmed or cooled, the colour changes give a visual demonstration of the effect of temperature on the equilibrium position of the reaction. Effect Of A Change In Pressure For a change in pressure to affect a chemical system in equilibrium, One of the reactants or products in the reversible reaction must be gaseous; The total number of moles of gaseous molecules on the left side of the equation must be different from the total number of moles of gaseous molecules on the right side. An example of such a reaction is the formation of ammonia from nitrogen and hydrogen. Here, 4 moles or volumes of reactants give only 2 moles or volumes of products. The forward reaction results in a decrease in the pressure of the system. The backward reaction results in an increase in the number of moles and hence an increase in the pressure of the system. Conversely, if low pressure is imposed on an equilibrium system, then the reaction which results in an increase in pressure will be favoured. In the formation of ammonia, a high pressure will favour the forward reaction where a reduction in volume occurs, causing the equilibrium position to shift to the right. This is because the forward reaction will counteract the high pressure imposed on the system by lowering the pressure, thus keeping the equilibrium constant, K, unchanged. In the process, a high percentage yield of the product is obtained On the other hand, a low pressure will cause the equilibrium position of this system to shift to the left, favouring the backward reaction. This will cause an increase in the pressure of the system and, thus, keep the equilibrium constant unchanged. It will also result in high yields of the reactants. In the following endothermic reaction, the forward reaction results in an increase in volume. How to Easily Memorize Elements of the Electrochemical Series In this case, a low pressure will favor the forward reaction and product formation, while a high pressure will favor the backward reaction. In reversible reactions involving the same number of moles of the equation as shown in the reaction below, a change in pressure has no effect on the position of equilibrium. However, the rate at which equilibrium is attained is faster at higher pressures. Effect Of A Change In Concentration In an equilibrium mixture, there is a balance between the concentrations of the reactants and the products, i. If more reactants are introduced into the equilibrium system, the balance will be upset. In order to relieve this constraint i. This results in a proportional increase in the concentration of the products, and so the equilibrium constant remains unchanged. Another way of increasing the amount of products formed is by continually removing the products from the system. Again, in order to relieve the constraint, the equilibrium will shift to the right, favouring the forward reaction. In the reversible reaction of iron with steam, The hydrogen is constantly being removed by a current of steam that passes over the heated

iron, until all the iron is converted to iron II diiron III oxide. The reverse process occurs if water vapor is removed. The preparation of many important chemicals from reversible systems depends upon the principle of removing one of the products from the reaction vessel. In the manufacture of calcium oxide by the heating of limestone in a limekiln, the carbon IV oxide produced is removed by an air current. We can illustrate the effect of concentration on the position of equilibrium by using the following chemical systems. Iron III tricyanate system: When we add some iron III trioxonitrate V solution to an equal volume of potassium thiocyanate solution, a blood-red colouration of a certain intensity results when equilibrium is reached. The reaction is represented as follows: The intensity of the red colour in the reaction mixture remains the same as long as the equilibrium is not disturbed. If the concentration of a reactant or product is changed, the equilibrium position shifts to annul the change. This is seen by a change in the colour intensity of the reaction mixture. Potassium heptaoxochromate VI system: Potassium heptaoxochromate VI is an orange colored solid which when dissolved in water, dissociates into hydrogen ions and yellow tetraoxochromate VI ions as shown. The yellow color of the tetraoxochromate VI ions masks the orange color of the heptaoxochromate VI ions. Thus, a neutral solution of potassium heptaoxochromate VI looks yellow. When the hydrogen ion concentration is increased by adding an acid , the equilibrium position shifts to the left. This is reflected by the colour of the solution changing from yellow to orange. When the hydrogen ion concentration is decreased by adding an alkali , the equilibrium position shifts to the right as indicated by an increase in the intensity of the yellow colour. Effect Of Catalyst We have learnt that a catalyst does not change the position of equilibrium. Thus, it cannot be used to increase product formation. However, catalysts are extremely important in chemical industry because they affect the time taken for a reaction to reach equilibrium. A positive catalyst increases the rates of both the forward and backward reactions to reach equilibrium. A positive catalyst increases the rates of both the forward and backward reactions to the same extent by lowering the activation energy of the reaction. As a result, equilibrium is reached in a shorter time than in an uncatalyzed reaction. A negative catalyst slows down the rate of reaction so that a longer time is needed to reach equilibrium. More from my site.

Chapter 8 : Equilibrium and Le Chatelier's Principle (Experiment) - Chemistry LibreTexts

Equilibrium: Applying Le Chatelier's Principle Purpose: This experiment illustrated Le Chatelier's principle, which states that a system at equilibrium that is disturbed will shift to counteract the disturbance and reestablished equilibrium.

Chapter 9 : Le Chatelier's Principle

Some Examples of Chemical Equilibria, Le Chatelier's Principle Introduction: The purpose of this experiment is to observe and understand chemical reactions that are examples of chemical system at equilibrium.