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Homogeneous chemical equilibrium lab report

The equilibrium constants for homogenous and heterogeneous solutions should be calculated differently. Describe the difference between homogenous and heterogeneous solution equilibrium. Key takeaways Key Points The Equilibrium Chances (K) are measured by the ratio of products to reactors at equilibrium. When calculating equilibrium const constants, pure solids and liquids are not taken into account. This is because their concentration does not change during the response. In homogenous equilibrium, all substances are in the same phase. In heterogeneous equilibrium, substances are in different phases. Key terms equilibrium: The state of a response in which the rates of the forward and reverse responses are the same. Heterogeneous solution: A solution consists of different states of matter. Homogeneous solution: A solution consists of matter that all exist in the same state. The equilibrium constants for reactions containing substances that are all in the same phase, and reactions containing substances in different phases, should be calculated differently. The former is called homogenous reactions, and the later is called heterogeneous reactions. Equilibrium constants The equilibrium constant K for a given response is defined as the ratio of the products of a response to the reactors, measured at equilibrium. In a general response $aA + bB \rightleftharpoons cC + dD$ The response quote measured at equilibrium is the equilibrium constant K. $K = \frac{[C]^c [D]^d}{[A]^a [B]^b}$ The chemical species in question can be molecules, ions or a mixture of both. For the homogeneous response $C_2H_2(aq) + 2Br_2(aq) \rightleftharpoons C_2H_2Br_4(aq)$ The equilibrium constant $K = \frac{[C_2H_2Br_4]}{[C_2H_2][Br_2]^2}$ Heterogeneous Equilibrium A heterogeneous equilibrium is a system in which reactors and products are found in two or more phases. The phases can be any combination of solid, liquid or gas phases and solutions. When dealing with this equilibrium, remember that solids and pure liquids do not appear in equilibrium constant expressions. For example, for the response $Br_2(l) \rightleftharpoons Br_2(g)$ The equilibrium constant K is simple $[Br_2]$, with the concentration of the pure liquid Br_2 excluded. The common ion effect describes the changes that occur with the introduction of ions to a solution that contains the same ion. Explain the general ion effect. Key takeaways key points The role that the general ion effect plays in solutions is mostly visible in the of solubility of solids. By adding general ions reduces the solubility of a compound generally due to a shift in equilibrium. The general ion effect also plays a role in regulating buffers. Buffering solutions contain either an acid or base, accompanied by its conjugate counterpart. Adding more like conjugate ions will eventually shift the pH of the solution. The general ion effect should be taken into account when determining solution equilibrium when adding by ions that are already present in the solution. Key terms Le Chatelier's principle: The principle used to predict the effect of a change in conditions on a chemical equilibrium. buffer: A solution used to stabilize the pH (acidity) of a liquid. assessed basis: The species created after the donation of a proton. assess acid: The species created when a base accepts a proton. Solubility refers to the amount of material that can be dissolved in a particular solvent. For example, table salt (NaCl) placed in water eventually disses. However, if more table salt is added continuously, the solution will reach a point at which to no longer be resolved; In other words, the solution is saturated, and the table salt has effectively reached its solubility limit. Chemical equilibrium is the chemical condition where there are no net physical or chemical changes between the reactor and the products of a reaction. This is because the rates of the forward (reactor for product) and reverse (product to reactor) reactions are equal. Solubility equilibrium refers to the state of chemical equilibrium between a chemical compound in the solid state and a solution consisting of that decomposed compound. This equilibrium is fixed when the rates of migration between the solid and aqueous phases of the molecules (or ion) are equal. General Ion Effect When a solution of an ionic substance comes into contact with another ionic connection to a common ion, the solubility of the ionic substance decreases significantly. For example, it would be like trying to resolve solid table salt (NaCl) in a solution where the chloride ion (Cl) is already present. The amount of NaCl that can decompose to reach the saturation point will be lowered. This phenomenon is the common ion effect and plays important roles in pharmaceutical and environmental areas. The general ion effect can be explained by Le Chatelier's principle of chemical equilibrium: $AB \rightleftharpoons A^+ + B^-$ For a simple decomposition process, adding more of one of the ion (A^+) from another connection will shift the composition to the left, reducing the concentration of the other ion (B^-), which effectively reduces the solubility of the solid (AB). For example, when dissolving in calcium and fluoride ions is the solubility product expression: $CaF_2 \rightleftharpoons Ca^{2+} + 2F^-$ This expression should always last even if some ionic specs come from other sources. Therefore, if more Ca^{2+} is placed in solution, the equilibrium will shift to the left, favor the solid shape and reduce the solubility of the solid. Lithium hydroxide with carbon growth: Lithium hydroxide forms less soluble lithium carbonate, which precipitates due to the general ion effect. Bumper Solutions A bumper solution consists of a weak acidity and its conjugate base, or a weak base and its conjugate acidity. Adding excess ions will change the pH of the buffer solution. Therefore, the common ion effect takes a role in pH regulation. According to Le Chatelier's principle, adding more ions changes the equilibrium and shifts the response to benefit the solid or deionized form. In the case of an acid buffer, the hydrogen ion concentration decreases, and the resulting solution is less acidic than a solution that contains the purely weak acid. For example in the response: $HCN \rightleftharpoons H^+ + CN^-$ The addition of cyanide ions (CN^-) will suppress the ionization of hydrogen cyanide (HCN) and move its equilibrium to the left. The percent dissociation of the hydrogen cyanide will decrease, thus reducing the H^+ ions and increasing the pH of the solution. Academia.edu use cookies to personalize content, customize ads, and improve the user experience. By using our website, you agree to our collection of information by using cookies. To learn more, see our Privacy Policy. Experiment 6: Chemical Equilibrium—The Hydrolysis of Ethyl Acetate Goals: \dot{u} Determines the Value of the Equilibrium Constant for a Response \dot{u} Use acid-based titrations and solution stoichiometry in determining the equilibrium constant definitions. \dot{u} Equilibrium – a balance condition in a chemical reaction in which the forward and deterioration is equal. Background material: For the response below, when A and B were mixed, the response continues in the forward direction to produce C and D. However, as time progresses, the concentration of C and D increases an increase in the rate of the reverse response. Simultaneously with this increased rate of reverse response is a reduction of the forward rate due to the decrease in the concentration of A and B. At some point, the rate of the forward and reverse reactions will become the same and we will achieve a state of dynamic equilibrium: $aA + bB \rightleftharpoons cC + dD$ Equilibrium does not mean that the forward and reverse reactions are not. Molecules of A and B still respond to form C and D and molecules of C and D respond to form A and B. As the rate of the forward and reverse reactions is the same, however, nothing will appear to be happening. As such, all physical and chemical properties such as pH, color and will remain constant. For a general equilibrium equation, we can specify an equilibrium constant, Kequil, associated with the concentrations of all product and reactive species, where [A], [B], [C], and [D] are the molar concentration of all species present in equilibrium. The eponents, a, b, c, and d represent the stoichiometry coefficients of the balance chemical reaction. Kequil is really the proportion of the rate of the reverse reaction divided by the rate of the forward reaction and so is a dimensional constant at a given temperature. We will have the acid catalysed (HCl) hydrolysis of an ester (ethyl acetate, EtAc), to an alcohol (ethanol, EtOH) and an acid (acetic acid, HAc) to form: $CH_3CH_2CO_2CH_3 + H_2O \rightleftharpoons CH_3CH_2OH + CH_3CO_2H$ EtAc + H₂O \rightleftharpoons EtOH + HAc This hydrolysis response will occur spontaneously at room temperature and adding acid only affects the amount of time it takes the response to achieve in practice it is often difficult to determine the equilibrium concentration of all these species However, we can calculate the equilibrium constant by determining the concentration of a single species in equilibrium, if we know the initial concentration of all species. We don't have to worry about the concentration of the hydrochloric acid, as it only serves as a catalyst. For our specific response, the equilibrium constant will be equal: Since we will know the initial concentration of each component, we only need to determine the equilibrium concentration of one component to calculate the equilibrium constant. Although we have several chemical species to choose from, we will monitor the acetic acid concentration as it can be determined very accurately via a simple titration with standardized base. In the above discussion, we decided to monitor the concentration of the HAc to determine the equilibrium constant. However, the presence of the strong acid HCl complicates our task. A simple titration with sodium hydroxide will neutralize all the acids present and we won't be able to discern the amount of base needed to neutralize the HCl versus the amount needed to neutralize the HAc. This is a common problem in chemistry and is solved by an empty, which contains only the HCl (Solution 1). Since it contains a known amount of HCl and water, a simple titration will determine the amount of base needed to neutralize the HCl contained in the other solutions. This amount of HCl will be constant in the other solutions because the HCl is a catalyst not consumed by the response. Now that we know the amount of base needed to account for the HCl, any additional base used in the titration of Solution #3/4, Solution #5/6, and Solution #7/8, should represent the amount of HAc present. Sample calculation: The volume of 1.090 M NaOH used for titrations was: Scales 1 & 2 Vials & 4 Scales 5 & 6 Scales 7 & 8 1 34.74 ml 20.25 ml 29.60 ml Scales 1 and 2: Remember that mobility x volume = moles so: $M \times L = .0147 \text{ moles HCl}$ Getting grams HCl present: $.0147 \text{ mole HCl} \times 36.46 \text{ g/mole} = 0.536 \text{ g HCl}$ in 5 ml What is the total mass of the 5 ml HCl? $1.0436 \text{ g/mL} \times 5 \text{ ml} = 5.218 \text{ g HCl}$ What is the mass of water in the 5 ml HCl? $5.218 \text{ g} - 0.536 \text{ g} = 4.682 \text{ g H}_2\text{O}$ How much mole water present in the 5 ml HCl? $4.682 \text{ g H}_2\text{O} \div 18 \text{ g/mole} = 0.260 \text{ moles of water in 5 ml}$ Scales 3 and 4: This can help you set up a small table that is after the initially and final condition watch the response EtAc H₂O EtOH HAc Initial: 3 mL 2 mL + 0.260 mole 0 ml 0 ml Change: - x - x + x Calculate x and you are set. You get x from the titration information: Mole base used in titration: $1.090 \text{ M NaOH} \times .03474 \text{ L} = 0.03787 \text{ mole base}$ Since the stoichiometry of the response is 1 to 1, mole base = mole acid: $0.03787 \text{ mole base} = 0.03787 \text{ mole acid}$ This is NOT x because the titration only gives you moles of total acidity, do not forget about HCl: $0.03787 \text{ mole acid} - 0.0147 \text{ mole HCl} = 0.02317 \text{ mole acetic acid} = x$ Final mole ethanol is x to so mole ethanol = x How much moles of ethyl acetate present initially: $3 \text{ m} \times 0.894 \text{ g/ml} \div 88 \text{ g/mole} = 0.03056 \text{ mole ethielate}$ Initially How much mole water initially: $2 \text{ mL} \times 1 \text{ g/mL} \div 18 \text{ g/mole} = 0.1111 \text{ moles of water initially}$ Let's revisit the table : EtAc(aq) H₂O(soln) EtOH(aq) HAc (aq) Initial: .03056 mole 0.1111+ .026 mole mole 0 moles 0 moles Change: -.02317 moles -.02317 moles +.02317 moles +.02317 moles+.02317 mole Equilibrium: .00739 .34793 mole .02317 mole .02317 mole Now we can calculate the equilibrium constant for this response as: Kequil = Kequil = 0.209 Scales 5 & 6: Repeat the above. Be careful with x!!! Remember that the titration only tells you the moles of total acidity present. Scales 5 and 6 already initially have acid present. Watch out!!! Scales 7 & 8: Repeat the above. Items for lab report: \dot{u} Create a table like the one on page 59. \dot{u} Make sure in your table you include the volume of NaOH used in the titrations (you will have duplicates for each set of mixtures, average the volume of NaOH used and do 1 calculation for each set of mixtures). \dot{u} Show all calculations for sets 1/2, 3/4, 5/6, and 7/8 \dot{u} Be very clear in your presentation—I will not look for your answers!!! Answers!!!!

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