Equilibrium

Case Study Based Questions

Read the following passages and answer the questions that follow:

1. When we increase the pressure on a system in which NO, O_z and NO_2 are at equilibrium:

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

The formation of additional amounts of NO_2 decreases the total number of molecules in the system because each time two molecules of NO_2 form, a total of three molecules of NO and O_2 are consumed. This reduces the total pressure exerted by the system but does not completely relieve, the stress of the increased pressure. On the other hand, a decrease in the pressure on the system favours decomposition of NO_2 into NO and O_2 which tends to restore the pressure. Now consider this reaction:

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

Because there is no change in the total number of molecules in the system during the reaction, a change in pressure does not favour either formation or decomposition of gaseous nitrogen monoxide.

(A) Which type of equilibrium is established for the following: Ice Water?

(B) What is the effect of pressure on the equilibrium constant?

(C) What will happen to the concentrations of N2 O2 and NO at equilibrium if a catalyst is added?

Ans. (A) It is a heterogeneous equilibrium because all reactants and products are in the different phase. When pressure is applied to the equilibrium state, Vice> VWater Thus, an increase in pressure will show a forward reaction.

(B) Equilibrium constant is not changed if you change the pressure of the system. The only thing that changes an equilibrium constant is a change of temperature. The position of equilibrium may be changed if you change the pressure.

(C) When a catalyst is added it basically has no effect on the equilibrium of the reaction. A catalyst only increases the speed of reaction or more so the forward and reverse reaction rates are equal. When this happens, the system will reach equilibrium faster.

2. Let us consider a system at equilibrium consisting of a number of phases-vapour, solid and liquid-of various components. We shall assume that it is shut off from outside action by a rigid non-conducting envelope. This means no loss of generality, as, if the system is already in equilibrium, it will still be so if enclosed in such an envelope. The following results are obtained:

(1) Thermal Conditions: The temperature must be uniform; otherwise, heat will flow from high to low temperature.

(2) Mechanical Conditions: If only ordinary hydrostatic pressure is considered it, must be everywhere equal, but if gravity is taken into account of or membranes are present this must be modified.

(3) Chemical Conditions: The variable which determines chemical equilibrium is not so familiar as temperature, pressure or electrical potential and a few words are there for necessary in explanation of what is best called the "chemical potential" of a substance. Consider a particular substance present in two different phases and for a moment deal with these phases separately. Let a supply of the substance be available in a reference state and an infinitesimal amount of it pass from this state to one of the phases. Keep the volume and entropy of the phase constant by subtracting heat and changing the pressure.

(A) What do you mean by phase equilibrium?

(a) No reaction

(b) Study of equilibrium which exists between or within different states of matter namely solid, liquid and gas.

(c) All reactants must be in the same phase.

(d) Product concentration is more.

(B) Which property of the phases must be equal when phases of a system are in thermodynamic equilibrium with each other?

(a) Chemical potential

(b) Volume

(c) Temperature

(d) Pressure

(C) When does equilibrium exist between the phases of a substance?

(a) At 0 K

(b) At 1073 K

(c) At the triple point
(d) At 273 K

(D) Can four phases exist in equilibrium?
(a) Yes
(b) No
(c) Can't define
(d) In certain case

(E) When solid and liquid phases are in equilibrium?
(a) Melting point
(b) Freezing point
(c) Never
(d) Both (a) and (b)

Ans. (A) (b) Study of equilibrium which exists between or within different states of matter namely solid, liquid and gas.

Explanation: Phase equilibrium is the study of the equilibrium which exists between or within different states of matter namely solid, liquid and gas.

(B) (a) Chemical potential

Explanation: The chemical potential of the component of a thermodynamic system in a given phase is a thermodynamic state function.

(C) (c) At the triple point

Explanation: In thermodynamics, the triple point of a substance is the temperature and pressure at which the three phases of that substance coexist in thermodynamic equilibrium.

(D) (a) Yes

Explanation: Four phases of a substance composed of a single component can coexist at thermal equilibrium.

(E) (d) both (a)

Explanation: The melting point of a solid is the same as the freezing point of the liquid. At that temperature, the solid and liquid states of the substance are in equilibrium. For water, this equilibrium occurs at 0°C.

3. Many chemical systems, as well as physiological processes like oxygen transport by haemoglobin in blood and acid-base balance in the human body, require an understanding of equilibrium constants. The equilibrium constant of any chemical reaction is the value of its reaction quotient at chemical equilibrium, a state reached by a dynamic chemical system after a period of time has passed in which its composition shows no observable tendency to change. Equilibrium constants include stability constants, formation constants, binding constants, association constants, and dissociation constants. Allowing a single reaction to reach equilibrium and then measuring the concentrations of each chemical involved in that reaction yields the numerical value of an equilibrium constant. The ratio between the product and reactant concentrations is computed. Because concentrations are recorded at equilibrium, the equilibrium constant remains constant regardless of initial concentrations for a specific reaction.

(A) For the equilibrium, PCl_5 $PCL_3 + Cl_2$, $Kc=a^2/(1-a)V$. What effect does a change in volume have on the equilibrium constant for a particular reaction if the temperature remains constant?

(B) In the following equation, what change in temperature is seen to drive the reaction in the forward direction?

$I_{2(g)} \rightleftharpoons 2I_{(g)};$ ΔH° at (298K) = +250kJ

(C) Equilibrium concentration of H, and I, are heated to equilibrium in a 2 litre flask. At equilibrium, the forward and backward rate constants are found to be equal.

(i) What is the equilibrium concentration of H, and I?

(ii) What percentage of the initial concentration of H, has reacted at equilibrium?

Ans. (A) For the given equilibrium reaction, the K is the characteristic constant for the given reaction at a particular temperature. The equilibrium constant depends on the temperature only. It does not depend on the concentration or volume. So, if the temperature remains constant and the volume changes, then there will be no change in the value of K

(B)

$I_{2(g)} \rightleftharpoons 2I_{(g)};\ \Delta H^\circ(298\ K) = +250\ kJ$

It is an endothermic reaction because the process requires heat energy from the surrounding and the value of ΔH° is positive. For the reaction to occur, the energy consumed in the reaction acts as an inclination or activation energy. As a result of this, there is a rise in temperature and hence, the reaction shifts in the forward direction.

$H_2 + I_2 \rightleftharpoons 2HI$

(c)

(i) Initial $\quad 1 \quad 1 \quad 0$

At eq. $\quad 1-x \quad 1-x \quad 2x$

$$\frac{1-x}{2} \quad \frac{1-x}{2} \quad \frac{2x}{2}$$

$$K = \frac{K_f}{K_b} = 1 \text{ as } K_f = K_b$$

$$1 = \frac{[HI]^2}{[H_2][I_2]}$$

$$1 = \frac{\left(\frac{2x}{2}\right)^2}{\left(\frac{1-x}{2}\right)^2}$$

$$\Rightarrow 1 = \frac{2x}{1-x}$$

$$\Rightarrow x = \frac{1}{3} = 0.33$$

Hence the equilibrium concentration for both H_2 and I_2 is $1-x$.

$\Rightarrow \; 1 - 0.33 = 0.67$

(ii) If initial concentration of H_2 reacted = x.

Therefore,

$$x \times 100 = 0.33 \times 100$$
$$= 33\%$$

4. The reaction quotient is a number in chemical thermodynamics that offers a measurement of the relative amounts of products and reactants present in a reaction mixture at a given moment in time for a reaction with well- defined overall stoichiometry. The reaction quotient is constant across time and equal to the equilibrium constant when the system is at equilibrium.

(1) If the value of reaction quotient (Q_c) is greater than that of the equilibrium constant (K_c), the reaction tends to form more reactants.

(2) If (Q_c) is lower than (K_c), the reaction tends to form more products in order to keep the equilibrium.

(3) If (Q_c) and (K_c) are equal, then the reaction mixture is at equilibrium.

As the reaction progresses, the activities of the species, and therefore the reaction quotient, vary in a way that lowers the system's free energy. The Gibbs free energy of reaction governs the direction of change by the relation

: $\Delta G = \Delta G^\circ + RT\ln Q$.

(A) The equilibrium constant Kc for the

reaction $P_{4(g)} \rightleftharpoons 2P_{2(g)}$ is 1.4 at 400°C.

Suppose 3 moles of P4 and 2 moles 1 of P_2 are mixed in a 2 L container at 400°C. What is the reaction quotient Q_c ?

(a) $\frac{2}{3}$ (b) $\frac{1}{2}$

(c) 1 (d) 2

(B) The ratio of equilibrium constant for the reaction and the reaction quotient is

(given as $\frac{K}{Q}$) 1.33. This means that:

(a) The reaction mixture will equilibrate to form more reactant species.
(b) The reaction mixture will equilibrate to form more product species.
(c) The equilibrium ratio of reactant and product concentration will be 1.33.
(d) The equilibrium ratio of reactant and product concentration will be 0.33.

(C) Which of the following statement is incorrect?

(a) The reaction quotient is useful in predicting the volume of reaction by comparing the values of K_c, and Q_c.
(b) The value of the equilibrium constant is independent of the initial and final concentration of reactants and products.
(c) Equilibrium constant is temperature dependent.
(d) If the value of $K_c > Q_c$, the reaction goes in the forward direction i.e., in the direction of products.

(D) Large value of Kc and Kp, favours the product..................

(a) rarely
(b) strongly
(c) comparable
(d) does not favour the products

Ans. (A) (a) 2/3

Explanation: $Q_c = \frac{[P_2]^2}{[P_4]} = \frac{\left(\frac{2}{2}\right)^2}{\left(\frac{3}{2}\right)} = \frac{2}{3}$

(B) (a) The reaction mixture will equilibrate to form more reactant species.

Explanation:

$$\because \quad \frac{K}{Q} = 1.33$$
$$\Rightarrow \quad K < Q$$

This means the reaction will move in the backward direction to achieve equilibrium. Hence, more reactants will form.

(C) (a) The reaction quotient is useful in predicting the volume of reaction by comparing the values of K_c and Q¿$_c$

Explanation: The reaction quotient is useful in predicting the direction of reaction, not the volume by comparing the values of K_c and Q_c

(D) (b) strongly

Explanation: Large value of K_c, and K_p favours the product strongly. The values are larger than about 10^3.

5. Le Chatelier's principle is also known as the equilibrium law, used to predict the effect of change on a system at chemical equilibrium. This principle states that equilibrium adjusts the forward and backward reactions in such a way as to accept the change affecting the equilibrium condition. When factors like concentration, pressure, temperature and inert gas that affect equilibrium are changed, the equilibrium will shift in that direction where the effects that caused by these changes are nullified. This principle is also used to manipulate reversible reactions in order to obtain suitable outcomes. Reactions proceed in both directions (reactants go to products and products go to reactants). We can tell a reaction is at equilibrium if the reaction quotient (Q) is equal to the equilibrium constant (K). We next address what happens when a system at equilibrium is disturbed so that Q is no longer equal to K. If a system at equilibrium is subjected to a perturbance or stress (such as a change in concentration) the position of equilibrium changes. Since this stress affects the concentrations of the reactants and the products, the value of Q will no longer equal the value of K. To re-establish equilibrium,

the system will either shift toward the products (if
Q<K) or the reactants (if Q>K) until Q returns to the same value as K. This process is described by Le Chatelier's principle: When a chemical system at equilibrium is disturbed, it returns to equilibrium by counteracting the disturbance. As described in the previous paragraph, the disturbance causes a change in Q; the reaction will shift to re-establish Q = K.

(A) In a vessel N_2, H_2 and NH, are at equilibrium. Some helium gas is introduced into the vessel so that the total pressure increases while temperature and volume remain constant. According to Le Chatelier's principle, the dissociation of NH,:

(a) Increases

(b) Decreases

(c) Remains unchanged

(d) Equilibrium is disturbed

(B) The following reaction is at equilibrium,

$$\underset{\text{Yellow}}{Fe^{3+}_{(aq)}} + \underset{\text{Colourless}}{SCN^{-}_{(aq)}} \rightleftharpoons \underset{\text{Deep red}}{\left[Fe(SCN)\right]^{2+}_{(aq)}}$$

In the above reaction, colour intensity of red colour can be increased by:

(a) Addition of KSCN

(b) Addition of oxalic acid which reacts with Fe^2+ ions.

(c) Addition of Hg^2+ ions which react with SCN-ions.

(d) Red colour intensity cannot be changed.

(C) The reaction

$$2SO_{2(g)} + O_{2(g)} \rightleftharpoons 2SO_{3(g)}; K_c = 1.7 \times 10^{26},$$

will be favoured by:

(a) high temperature and low pressure

(b) low temperature and high pressure

(c) high temperature and high pressure

(d) low temperature and low pressure

(D) Which statement is incorrect about catalyst?

(a) It alters the equilibrium position in a reversible reaction.

(b) It accelerates the reaction.

(c) Its small quantity is enough to bring the considerate amount of reaction.

(d) It is unchanged at the end of reaction.

(E) The equilibrium constant for the reaction

$2NO_{(g)} + O_{2(g)} \rightleftharpoons 2NO_{2(g)}$ is 4×10^{-4} at

2000 K. In presence of a catalyst, the equilibrium is attained three times faster. The equilibrium constant in the presence of the catalyst at 2000 K is:

(a) 40×10^{1}

(b) 4×10^{-4}

(c) 4×10^{-3}

(d) Difficult to compute without more data.

Ans. (A) (c) Remains unchanged

Explanation: Backward reaction is not favoured at high pressure.

(B) (a) Addition of KSCN

Explanation: Addition of KSCN increases the colour intensity of the solution as it shifts the equilibrium to right. Addition of reagents like oxalic acid or Hg2+ ions which remove Fe3+ or SCN ions shift the equilibrium to the left and colour intensity decreases.

(C) (b) low temperature and high pressure

Explanation: Since the reaction is exothermic it is favoured at low temperature. There is a decrease in the number of moles hence reaction is favoured at high pressure.

(D) (a) It alters the equilibrium position in a reversible reaction.

Explanation: A catalyst is a substance which speeds up the process of reaction but remains unchanged chemically at the end of reaction. Catalyst does not alter the equilibrium position for reversible reaction.

(E) (b) 4 × 10-4

Explanation: A catalyst increases the rate of forward and backward reactions by the same factor hence, does not change the equilibrium constant.

6. An acid-base reaction is a chemical reaction that occurs between an acid and a base. Several concepts exist that provide alternative definitions for the reaction mechanisms involved and their application in solving related problems. Despite several differences in definitions, their importance as different methods of analysis becomes apparent when they are applied to acid-base reactions for gaseous or liquid species, or when acid or base character may be somewhat less apparent. The Arrhenius definition of acid-base reactions, which was devised by Svante Arrhenius, is a development of the hydrogen theory of acids. It was used to provide a modern definition of acids and bases, and followed from Arrhenius's work with Friedrich Wilhelm Ostwald in establishing the presence of ions in an aqueous solution in 1884. This led to Arrhenius receiving the Nobel Prize in Chemistry in 1903. As defined by Arrhenius:

- An Arrhenius acid is a substance that dissociates in water to form hydrogen ions(H). In other words, an acid increases the concentration of H+ ions in an aqueous solution. This protonation of water yields the hydronium ion (H_3O^+); in modern times, H is used as a shorthand for (H_3O^+) because it is now known that a bare proton (H) does not exist as a free species in aqueous solution.
- An Arrhenius base is a substance that dissociates in water to form hydroxide(OH^-) ions. In other words, a base increases the concentration of (OH^-) ions in an aqueous solution.

(A) Comment on the statement: All Arrhenius acids are Bronsted acids but all Arrhenius bases may not be Bronsted bases.

(B) Why heat of neutralisation for a strong acid and strong base is always constant?

(C) Give two limitations of Arrhenius's theory?

Ans. (A) The statement is true. In the Arrhenius concept as well as in the Bronsted concept, an acid is regarded as a substance which can give an H+ ion or proton. Therefore, all Arrhenius acids are Bronsted acids. On the other hand, the definitions of base aredifferent in the two concepts. In Arrhenius's concept, a base is regarded as a substance which gives OH- ions in solution whereas a Bronsted base is a proton acceptor. Hence, Arrhenius bases may not be Bronsted bases. For example, NaOH can give OH- ions but cannot accept a proton. Hence, NaOH is an Arrhenius base but not a Bronsted base.

(B) Heat of neutralisation is always constant for a strong acid and strong base because they are completely ionised in dilute solution.

(C) Limitations of Arrhenius's theory are:

(1) It only applies to substances that produce hydrogen ions (H^+) or hydroxide ions (OH^-).

(2) Arrhenius did not explain in his theory the behaviour of acids and bases in a non-aqueous solution.

7. A very strong base would give a very weak conjugate acid. A weak acid HA is only partially dissociated in an aqueous medium and thus, the solution mainly contains undissociated HA molecules. Example: HCL is a strong acid. If HCL is a strong acid, it must be a good proton donor. HCL can only be a good proton donor, however, if the Ction is a poor proton acceptor. Thus, the Cl-ion must be a weak base. Strong acid-like perchloric acid, hydrochloric acid will release conjugate base ions CLO_4 Cl- these are weaker bases than water. Similarly, a very strong base would give a very weak conjugated acid. On the other hand, a weak acid say HA is only partially dissociated in an aqueous medium and the solution mainly contains undissociated HA molecules. The value of K_a for an acid can be used to decide whether it is a strong acid or a weak acid, in an absolute sense. It can also be used to compare the relative strengths of a pair of acids.

(A) Common ion effect can be used in which of the following cases:

(a) Cloth making

(b) Alcohol purification

(c) Quantitative analysis

(d) Qualitative analysis

(B) The K is equilibrium dependent upon:

(a) temperature

(b) pressure

(c) volume

(d) concentration of acid or base

(C) Choose the correct option for a weak acid.

(a) HF

(b) CH_3COONa

(c) Na_2CO_3

(d) H_2S

(D) Which of the following is the strongest conjugate base?

(a) F-

(b) Cl-

(c) I-

(d) Br

(E) Among following the one pair that does not act as conjugate acid-base pair is:

(a) H_2O, H_3O+

(b) HCO_3-H_2CO_3

(c) HSO_4-,H_2,SO_4

(d) NH_4+NH_2-,

Ans. (A) (d) Qualitative analysis

Explanation: Common ion effect is used in the purification of common salt, qualitative analysis. In the qualitative analysis, Ammonium hydroxide is added in the presence of ammonium chloride to avoid the precipitation of V group radicals.

(B) (a) temperature

Explanation: The value K_w is an equilibrium constant, therefore it is temperature-dependent.

(C) (a) HF

Explanation: In HF, hydroxonium ion is attached to the fluoride ion, it is not free to function as an acid, thus limiting the strength of HF in water.

(D) (a) P

Explanation: The stronger the acid, the weaker will be its conjugate base.

Order of acidic nature: HI > HBr >HCL > HF Conjugate base strength order:

F-> CL-> Br>I

(E) (d) NH_4^{+}”, NH_2^{-}

Explanation: NH4+ and NH3- are conjugate acid-base pair, not NH_4^{+}, NH_2^{-}

8. The common ion effect helps to finish the particular ion precipitated as sparingly soluble water, with a very low value of solubility product for gravimetric estimation. Thus, we can precipitate silver ion as silver chloride, ferric ion as its hydroxide and barium ion as its sulphate for quantitative estimation. The solubility of a solute in a particular solvent is the maximum concentration that may be achieved under given conditions when the dissolution process is at equilibrium. When a solute's concentration is equal to its solubility, the solution is said to be saturated with that solute. If the solute's concentration is less than its solubility, the solution is said to be unsaturated. A solution that contains a relatively low concentration of solute is called dilute, and one with a relatively high concentration is called concentrated.

(A) Differentiate the solutions depending upon the solubility of the salt.

(B) Explain the effect of common ions on the solubility of ions salts.

(C) What is buffer capacity?

(D) The molar solubility of $PbBr_2$, is 2.17×10^{-3} M at a certain temperature.

Calculate K_{sp} for $PbBr_2$

Ans. (A) As a general rule, for a salt to be able to dissolve in a particular solvent its solvation enthalpy must be greater than its lattice enthalpy so that the latter may be overcome by the former. Each salt has its characteristics of solubility which depends upon the temperature.

Depending upon the solubility of the salt differentiate into three:

Category I	Soluble	Solubility > 0.1 M
Category II	Slightly Soluble	0.1M < Solubility < 0.01 M
Category III	Sparingly Soluble	Solubility < 0.01 M

(B) (1) One of the ionic concentrations increase then it will combine with the opposite charge ions and certain salt will undergo precipitation till it reaches

a point where, K_{sp}, Q_{sp}

(2) If one of the ionic concentrations is decreasing, then higher will be the salt dissolve which increases both ions concentration till it reaches a point where $K_{sp}=Q_{sp}$ this is used for even the salt like sodium chloride.

(C) The number of millions of acids or bases to be added to a litre of buffer solution to change the pH by one unit is the buffer capacity of the buffer. Buffer capacity no. of moles of acid or base per litre/Change in pH

(D) For PbBr2, the expression of solubility product is

$K_{sp} = [Pb^{2+}] [Br^{-}]^{2} = (S) (2S)^2 = 4S^3$

Substitute S = 2.17 x 10-3 in the above

expression

Then, $K_{sp} = 4S^3$

$= 4 (2.17 \times 10^{-3})^3 = 4.1 \times 10^{-8}$

9. The idea of pH concerns the 'logarithm of the equilibrium' process of water autoionisation. Therefore, the notion is practically limited to aqueous solutions.

The pH of an aqueous solution' is determined using the pH scale, which in water varies typically from 0 to 14. A pH of 7 is neutral. A pH of less than 7 is acidic, while that of 7 or above is considered basic.

A salt is an ionic compound that is formed when an acid and a base neutralize each other. While it may seem that salt solutions are always neutral, they can frequently be either acidic or basic.

Salt hydrolysis is a reaction in which one of the ions from a salt reacts with water,

forming either an acidic or basic solution.
Salt formed from: Salt Solution
Strong acid + Strong base Neutral
Strong acid + Weak base Acidic
Weak acid + Strong base Basic

(A) The PK of weak acid HA is 4.80 and pK of a weak base BOH is 4.78. The pH of an aqueous solutions of corresponding salt BA will be:

(a) 7.01
(b) 4.79
(c) 9.22
(d) 10.0

(B) What will be the value of pH of 0.01 mol dm-^{3} CH,COOH (K = 1.74 x 10-5)?

(a) 3.4
(b) 3.6
(c) 3.9
(d) 3.0

(C) A solution of an acid has a pH = 5.20. Find out the concentration of OH- ions. (pK$_w$= 14):

(a) 6.58 x 10^{-8}
(b) 1.58 x 10^{-9}
(c) 8.8 x 10^{-9}
(d) 5.5 x 10^{-8}

(D) Which of the following salt gives the highest pH in water?

(a) KCL
(b) NaCl
(c) Na$_2$,CO$_3$,
(d) CuSO$_4$

(E) Lemon juice (10 ml) is diluted with an equal amount of water (10 ml). What impact on the pH of the solution is most likely to be seen?

(a) pH of diluted lemon juice would be more than pure lemon juice.
(b) pH of diluted lemon juice would be less than pure lemon juice.
(c) pH of diluted lemon juice would be equal to than pure lemon juice.
(d) No effect on pH.

Ans. (A) (a) 7.01

Explanation: For a salt of a weak acid and weak base:

$$pH = \frac{1}{2}\,[pK_w + pK_a - pK_b]$$

$$= \frac{1}{2}(14 + 4.80 - 4.78) = 7.01$$

(B)

(a) 3.4

Explanation: Given: $C = 0.01$ mol/dm³ and $K_a = 1.74 \times 10^{-5}$

So, we have $pH = -\log[H^+]$ and

$$[H^+] = \sqrt{K_a} \times C = \sqrt{1.74} \times 10^{-5} \times 0.01$$

$$= 4.2 \times 10^{-4}$$

$$pH = -\log(4.2\times10^{-4}) = -[\log 4.2 + \log 10^{-4}]$$

$$= 3.4$$

(C) (b) 1.58×10^{-9}

Explanation: $pH + pOH = pK_w = 14$ (given)

$pOH = pK_w - pH = 14 - 5.2 = 8.8$

$[OH^-] = 10^{-pOH} = 10^{-8.8} = 1.58 \times 10^{-9}$

(D) (c) Na,CO,

Explanation: Salt of weak acid + strong base: pH >7 e.g. Na_2CO_3,

Salt of strong acid + weak base: pH < 7 e.g.$CuSO_4$

Salt of strong acid + strong base: pH = 7, e.g. KCL and NaCl

Hence out of the given options, Na,CO, will have the highest pH.

(E) (a) pH of diluted lemon juice would be more than pure lemon juice.

Explanation: Therefore, the pH of the solution (diluted lemon juice) would be more than that of pure lemon juice. Lemon juice helps lower the pH levels by raising the number of hydronium ions and decreasing hydroxide ions. As the hydronium ions in the water are what determine the pH level, it is probable that the higher the hydronium content, the higher the pH level.