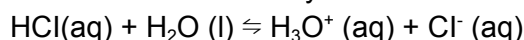


Class 11 Equilibrium Important Questions with Answers

Short Answer Type Questions

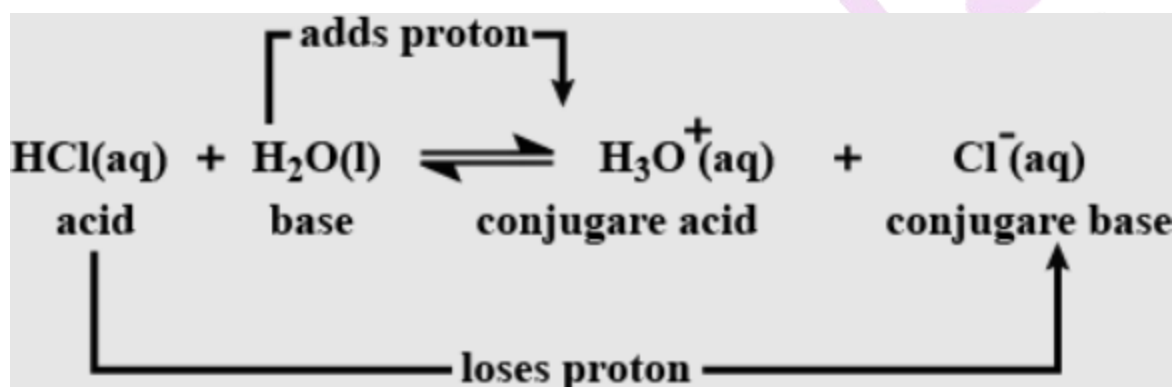
Q1. The ionisation of hydrochloric in water is given below:



Label two conjugate acid-base pairs in this ionisation.

Answer:

Conjugate acid and conjugate base are marked as follows.



Q2. The aqueous solution of sugar does not conduct electricity. However, when sodium chloride is added to water, it conducts electricity. How will you explain this statement on the basis of ionisation and how is it affected by the concentration of sodium chloride?

Answer:

Sugar is a non-electrolyte, and when it dissolves in water, there will be no ionisation. If there are no free ions, it will not conduct electricity. In contrast, When sodium chloride is added to water, it is wholly ionised to give Na^+ ions and Cl^- ions. Due to the presence of ions, there will be electrical conductance as ions are carriers of electric current.

Q3. BF_3 does not have protons but still acts as an acid and reacts with NH_3 . Why is it so? What type of bond is formed between the two?

Answer:

BF_3 is electron-deficient and hence acts as Lewis acid. NH_3 has one lone pair, which can donate to BF_3 and form a coordinate bond with NH_3 , i.e. $(\text{NH}_3: \rightarrow \text{BF}_3)$.

Q4. The ionisation constant of a weak base MOH, is given by the expression

$$K_b = \frac{[\text{M}^+][\text{OH}^-]}{[\text{MOH}]}$$

Values of the ionisation constant of some weak bases at a particular temperature are given below:

Base	Dimethylamine	Urea	Pyridine	Ammonia
K_b	5.4×10^{-4}	1.3×10^{-14}	1.77×10^{-9}	1.77×10^{-5}

Arrange the bases in decreasing order of the extent of their ionisation at equilibrium. Which of the above base is the strongest?

Answer:

Higher the value of K_b , the stronger will be the base. Among the given bases, the most substantial base is Dimethylamine.

Dimethylamine > Ammonia > Pyridine > Urea.

Q5. Conjugate acid of a weak base is always stronger. The decreasing order of basic strength of the following conjugate bases will be?

OH^- , RO^- , CH_3COO^- , Cl^-

Answer:

Conjugate acids of given bases are H_2O , ROH , CH_3COOH , HCl .

Their acidic strength is in the order.

$\text{HCl} > \text{CH}_3\text{COOH} > \text{H}_2\text{O} > \text{ROH}$.

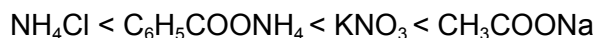
Hence, their conjugate bases will have strength in the order.

$\text{RO}^- > \text{OH}^- > \text{CH}_3\text{COO}^- > \text{Cl}^-$

Q6. Arrange the following in increasing order of pH.

$\text{KNO}_3(\text{aq})$, $\text{CH}_3\text{COONa}(\text{aq})$, $\text{NH}_4\text{Cl}(\text{aq})$, $\text{C}_6\text{H}_5\text{COONH}_4(\text{aq})$

Answer:



KNO_3 is formed from a strong acid, nitric acid and strong base, potassium hydroxide. Salts of strong acid and strong base hydrolyse to form a neutral solution. Thus its pH will be nearly 7. Sodium acetate is formed from a weak acid, acetic acid, and strong base sodium hydroxide. Thus its pH will be more than 7. $\text{C}_6\text{H}_5\text{COONH}_4$ is formed from a mild acid, $\text{C}_6\text{H}_5\text{COOH}$ and a weak base, NH_4OH . Thus its pH will be less than 7. NH_4Cl is formed from a weak base, NH_4OH , and strong acid, HCl . Thus its pH value would be less than 5.

Q7. The value of K_c for the reaction $2\text{HI}(\text{g}) \rightleftharpoons \text{H}_2(\text{g}) + \text{I}_2(\text{g})$ is 1×10^{-4} . At a given time, the composition of reaction mixture is $[\text{HI}] = 2 \times 10^{-5} \text{ mol}$, $[\text{I}_2] = 1 \times 10^{-5} \text{ mol}$. Determine the direction in which the reaction will proceed?

Answer:

At a given time, the reaction quotient Q for the reaction will be given by the expression.

$$Q = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2}$$

$$Q = \frac{1 \times 10^{-5} \times 1 \times 10^{-5}}{(2 \times 10^{-5})^2}$$

$$Q = 1/4$$

$$Q = 0.25$$

$$Q = 2.5 \times 10^{-1}$$

As the value of the reaction quotient is greater than the value of K_c , i.e. 1×10^{-4} , the reaction will proceed in the reverse direction.

Q8. On the basis of the equation $\text{pH} = -\log [\text{H}^+]$, the pH of $10^{-8} \text{ mol dm}^{-3}$ solution of HCl should be 8. However, it is observed to be less than 7.0. Explain the reason.

Answer:

Here, the water concentration cannot be neglected since the solution is very dilute.

$$[\text{H}_3\text{O}^+] = 10^{-8} + 10^{-7} \text{ M}$$

$$[\text{H}_3\text{O}^+] = 10^{-8} (1 + 10)$$

$$[\text{H}_3\text{O}^+] = 11 \times 10^{-8} \text{ M}$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log 11 \times 10^{-8} \text{ M}$$

$$\text{pH} = 8 - \log 11$$

$$\text{pH} = 8 - 1.04$$

$$\text{pH} = 6.96$$

pH will be less than 7.0.

Q9. The pH of a solution of a strong acid is 5.0. What will be the pH of the solution obtained after diluting the given solution 100 times?

Answer:

pH = 5 means $[H^+] = 10^{-5}$

On diluting 100 times ,

$[H^+] = 10^{-5} / 100 = 10^{-7}$

On calculating the pH using the equation $pH = -\log [H^+]$, the pH value comes out to be 7, which is impossible.

Hence, Total H^+ ion concentration = H^+ ions from acid + H^+ ion from water

$[H^+] = 10^{-7} + 10^{-7} M$

$[H^+] = 2 \times 10^{-7}$

$pH = 7 - 0.3010$

$pH = 6.699$

Q10. A sparingly soluble salt gets precipitated only when the product of the concentration of its ions in the solution (Q_{sp}) becomes greater than its solubility product. If the solubility of $BaSO_4$ in water is $8 \times 10^{-4} \text{ mol dm}^{-3}$. Calculate its solubility in 0.01 mol dm^{-3} of H_2SO_4

Answer:

	$BaSO_4(s)$	\rightleftharpoons	$Ba^{2+}(aq)$	+	$SO_4^{2-}(aq)$
At $t = 0$	1		0		0
At equilibrium in water	$1 - S$		S		S
At equilibrium in the presence of sulphuric acid	$1 - S$		S		$S + 0.01$

K_{sp} for $BaSO_4$ in water = $[Ba^{2+}][SO_4^{2-}] = (S)(S) = S^2$

But $S = 8 \times 10^{-4} \text{ mol dm}^{-3}$

$\therefore K_{sp} = (8 \times 10^{-4})^2$

$= 64 \times 10^{-8} \dots (1)$

The expression for K_{sp} in the presence of sulphuric acid will be as follows.

$K_{sp} = (S)(S + 0.01) \dots (2)$

Since the value of K_{sp} will not change in the presence of sulphuric acid, therefore from (1) and (2)

$$(S)(S + 0.01) = 64 \times 10^{-8}$$

$$S^2 + 0.01 S = 64 \times 10^{-8}$$

$$S^2 + 0.01 S - 64 \times 10^{-8} = 0$$

$$\begin{aligned} \Rightarrow S &= \frac{-0.01 \pm \sqrt{(0.01)^2 + (4 \times 64 \times 10^{-8})}}{2} \\ &= \frac{-0.01 \pm \sqrt{10^{-4} + (256 \times 10^{-8})}}{2} \\ &= \frac{-0.01 \pm \sqrt{10^{-4} (1 + 256 \times 10^{-2})}}{2} \\ &= \frac{-0.01 \pm 10^{-2} \sqrt{1 + 0.256}}{2} \\ &= \frac{-0.01 \pm 10^{-2} \sqrt{1.256}}{2} \\ &= \frac{-10^{-2} + (1.12 \times 10^{-2})}{2} \\ &= \frac{(-1 + 1.12) \times 10^{-2}}{2} = \frac{0.12}{2} \times 10^{-2} \\ &= 6 \times 10^{-4} \text{ mol dm}^{-3} \end{aligned}$$

Q11. The pH of 0.08 mol dm^{-3} HOCl solution is 2.85. Calculate its ionisation constant.

Answer:

$$\text{pH of HOCl} = 2.85$$

$$\text{But, } -\text{pH} = \log [H^+]$$

$$\therefore -2.85 = \log [H^+]$$

-

$$3.15 = \log [H^+]$$

$$[H^+] = 1.413 \times 10^{-3}$$

For the weak monobasic acid $[H^+] = (K_a \times C)^{1/2}$

$$K_a = [H^+]^2 / C$$

$$K_a = (1.413 \times 10^{-3})^2 / 0.08$$

$$K_a = 24.957 \times 10^{-6}$$

$$K_a = 2.4957 \times 10^{-5}$$

Q12. Calculate the pH of a solution formed by mixing equal volumes of two solutions A and B of a strong acid having pH = 6 and pH = 4 respectively.

Answer:

The pH of solution A = 6,

Therefore, the concentration of $[H^+]$ ion in solution A = 10^{-6} mol / L

The pH of solution B = 4

Therefore, the concentration of $[H^+]$ ion in solution B = 10^{-4} mol / L.

On mixing one litre of each solution, total volume = 1L + 1L = 2L.

Amount of H^+ ions in 1L of solution A = concentration \times Volume (V) =

$$\text{Amount of } H^+ \text{ ions in 1L of solution A} = 10^{-6} \times 1 = 10^{-6}$$

$$\text{Amount of } H^+ \text{ ions in 1L of solution B} = 10^{-4} \times 1 = 10^{-4}$$

\therefore The total amount of H^+ ions in the solution formed by mixing solutions A and B is (10^{-5} mol + 10^{-4} mol)

This amount is present in 2L solution,

$$\therefore \text{Total } [H^+] = 10^{-4} (1 + 1.01) / 2 \text{ mol / L}$$

$$\text{Total } [H^+] = 1.01 \times 10^{-4} / 2 \text{ mol / L}$$

$$\text{Total } [H^+] = 0.5 \times 10^{-4} \text{ mol / L}$$

$$\text{Total } [H^+] = 5 \times 10^{-5} \text{ mol / L}$$

$$\text{pH} = -\log [H^+]$$

$$\text{pH} = -\log [5 \times 10^{-5}]$$

$$\text{pH} = -[\log 5 - 5 \log 10]$$

$$\text{pH} = -\log 5 + 5$$

$$\text{pH} = 5 - \log 5$$

$$\text{pH} = 5 - 0.6990$$

$$\text{pH} = 4.3010$$

$$\text{pH} = 4.3$$

Thus, the pH will be 4.3.

Q13. The solubility product of $\text{Al}(\text{OH})_3$ is 2.7×10^{-11} . Calculate its solubility in g / L and also find out the pH of this solution. (Atomic mass of Al = 27 u).

Answer:

Let S be the solubility of $\text{Al}(\text{OH})_3$

$$K_{sp} = [\text{Al}^{3+}] [\text{OH}^-]^3$$

$$K_{sp} = (S) (3S)^3$$

$$K_{sp} = 27S^4$$

$$S^4 = K_{sp} / 27$$

$$S^4 = 27 \times 10^{-11} / 27 \times 10$$

$$S^4 = 1 \times 10^{-12} \text{ mol / L.}$$

$$S = 1 \times 10^{-3} \text{ mol / L.}$$

(i) Solubility of $\text{Al}(\text{OH})_3$: Molar mass of $\text{Al}(\text{OH})_3$ is 78g.

Therefore, the solubility of $\text{Al}(\text{OH})_3$ in g / L = $1 \times 10^{-3} \times 78 \text{ g / L}$

Solubility of $\text{Al}(\text{OH})_3$ in g / L = $78 \times 10^{-3} \text{ g / L}$

Solubility of $\text{Al}(\text{OH})_3$ in g / L = $7.8 \times 10^{-2} \text{ g / L}$

(ii) pH of the solution: $S = 1 \times 10^{-3} \text{ mol / L}$

$$[\text{OH}] = 3S = 3 \times 1 \times 10^{-3} = 3 \times 10^{-3}$$

$$\text{pOH} = 3 - \log 3$$

$$\text{pH} = 14 - \text{pOH} = 11 + \log 3 = 11.4771.$$

Q14. Calculate the volume of water required to dissolve 0.1 g lead (II) chloride to get a saturated solution. (K_{sp} of $\text{PbCl}_2 = 3.2 \times 10^{-8}$, atomic mass of Pb= 207 u).

Answer:

$$K_{sp} \text{ of } \text{PbCl}_2 = 3.2 \times 10^{-8}$$

Let S be the solubility of PbCl_2 .

	$\text{PbCl}_2 (\text{s})$	\rightleftharpoons	$\text{Pb}^{2+} (\text{aq})$	+	$2 \text{Cl}^- (\text{aq})$
Concentration of specie at t = 0	1		0		0
Concentration of various specie at equilibrium	$1 - S$		S		2S

$$K_{sp} = [\text{Pb}^{2+}][\text{Cl}^-]^2$$

$$K_{sp} = (S)(2S)^2$$

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

$$S^3 = K_{sp} / 4$$

$$S^3 = 3.2 \times 10^{-8} / 4 \text{ mol / L}$$

$$S^3 = 8 \times 10^{-9} \text{ mol / L}$$

$$S = 2 \times 10^{-3} \text{ mol / L}$$

Molar Mass of $\text{PbCl}_2 = 278$

Hence, solubility of PbCl_2 in g / L = $2 \times 10^{-3} \times 278 \text{ g / L}$

Solubility of PbCl_2 in g / L = $556 \times 10^{-3} \text{ g / L}$

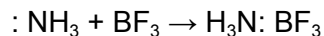
Solubility of PbCl_2 in g / L = 0.556 g / L

To get saturated solution, 0.556 g of PbCl_2 is dissolved in 1 L water.

0.1 g PbCl_2 is dissolved in $0.1 / 0.556 \text{ L} = 0.1798 \text{ L}$ water.

To make a saturated solution, dissolution of 0.1 g PbCl_2 in $0.1798 \text{ L} \approx 0.2 \text{ L}$ of water will be required.

Q15. A reaction between ammonia and boron trifluoride is given below:



Identify the acid and base in this reaction. Which theory explains it? What is the hybridisation of B and N in the reactants?

Answer:

NH_3 is Lewis base, while BF_3 is Lewis acid. Lewis's electronic theory of acids and bases explains it. The hybridisation state of nitrogen in NH_3 is sp^3 hybridised, and boron in BF_3 is sp^2 hybridised.

Q16. Following data is given for the reaction: $\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$

$$\Delta_f H^\circ [\text{CaO}(\text{s})] = -635.1 \text{ kJ / mol}$$

$$\Delta_f H^\circ [\text{CO}_2(\text{g})] = -393.5 \text{ kJ / mol}$$

$$\Delta_f H^\circ [\text{CaCO}_3(\text{s})] = -1206.9 \text{ kJ / mol}$$

Predict the effect of temperature on the equilibrium constant of the above reaction.

Answer:

$$\Delta_r H^\circ = \Delta_f H^\circ [\text{CaO}(\text{s})] + \Delta_f H^\circ [\text{CO}_2(\text{g})] - \Delta_f H^\circ [\text{CaCO}_3(\text{s})]$$

$$\therefore \Delta_r H^\circ = 178.3 \text{ kJ / mol.}$$

The reaction is endothermic.

Hence, according to Le Chatelier's principle, the reaction will proceed in the forward direction on increasing temperature.

Long Answer Type Questions

Q1. How can you predict the following stages of a reaction by comparing the value of K_c and Q_c ?

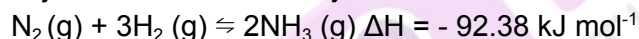
- (i) Net reaction proceeds in the forward direction.
- (ii) Net reaction proceeds in the backward direction.
- (iii) No net reaction occurs.

Answer:

The values of K_c and Q_c are self-explanatory and less than or greater than one another decides the direction in which reaction will proceed as follows-

- (i) As $Q_c < K_c$, the reaction will proceed in the direction of products (forward direction).
- (ii) If $Q_c > K_c$, the reaction will proceed in the direction of reactants (reverse direction).
- (iii) If $Q_c = K_c$, no net reaction will occur.

Q2. On the basis of the Le Chatelier principle explain how temperature and pressure can be adjusted to increase the yield of ammonia in the following reaction.



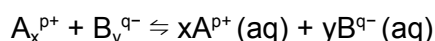
What will be the effect of addition of argon to the above reaction mixture at constant volume?

Answer:

According to Le Chatelier's principle, raising the temperature shifts the equilibrium to the left and decreases the equilibrium concentration of ammonia since it is an exothermic reaction. In other words, low temperature and high pressure are favourable for a high yield of ammonia. There will be no change in equilibria on the addition of argon (Ar).

Q3. A sparingly soluble salt having general formula $A_x^{p+} B_y^{q-}$ and molar solubility S is in equilibrium with its saturated solution. Derive a relationship between the solubility and solubility product for such salt.

Answer:



S moles of A B dissolve to give x S moles of A^{p+} and y S moles of Bq^{-} .

$$K_{sp} = [A^{p+}]^x [B^{q-}]^y$$

$$K_{sp} = [x5]^x [y5]^y$$

$$x^x y^y 5^{x+y}$$

Q4. Write a relation between ΔG and Q and define the meaning of each term and answer the following:

(a) Why a reaction proceeds forward when $Q < K$ and no net reaction occurs when $Q = K$.

(b) Explain the effect of an increase in pressure in terms of reaction quotient O. for the reaction: $CO(g) + 3H_2(g) \rightleftharpoons CH_4(g) + H_2O(g)$

Answer:

$$(a) \Delta G = \Delta G^\circ + RT \ln Q$$

ΔG° = Change in free energy as the reaction proceeds,

ΔG = Standard free energy change,

Q = Reaction quotient,

R = Gas constant,

T = Absolute temperature.

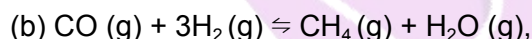
$$\text{Since } \Delta G^\circ = -RT \ln K$$

$$\therefore \Delta G = -RT \ln K + RT \ln Q$$

$$\Delta G = RT \ln Q / K$$

If $Q < K$, ΔG will be negative. The reaction will proceed in forward direction.

If $Q = K$, $\Delta G = 0$, the reaction is in equilibrium, and no net reaction is there.



According to the Le Chatiliers principle, when the pressure increases, the equilibrium will shift in that direction where there are fewer gas molecules.

As per the available information,

The number of moles of reactants = 1 + 3 = 4

The number of moles of product = 1 + 1 = 2

Hence, the equilibrium will shift forward as the number of moles of product is less.