CLASS 11 CHEMISTRY NCERT SOLUTIONS

NCERT Solutions for Class 11 Chemistry Chapter 6: Equilibrium – Free PDF Download

Equilibrium is one of the most conceptually important chapters in the Class 11 Chemistry syllabus, explaining why reversible physical and chemical processes settle into a stable, dynamic balance rather than running to completion. It builds the foundation for the law of chemical equilibrium, equilibrium constants Kc ...

Contents

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Class 11 Chemistry Chapter 6 – Notes and Extra Questions

Equilibrium is one of the most conceptually important chapters in the Class 11 Chemistry syllabus, explaining why reversible physical and chemical processes settle into a stable, dynamic balance rather than running to completion. It builds the foundation for the law of chemical equilibrium, equilibrium constants K_c and K_p, Le Chatelier's principle, and the vast topic of ionic equilibrium — acids, bases, pH, ionization constants, and the common-ion effect. Below you will find complete, independently worked NCERT Solutions for every exercise question of Class 11 Chemistry Chapter 6: Equilibrium, with full step-by-step derivations rather than just final answers.

NCERT Solutions for Class 11 Chemistry Chapter 6: Equilibrium

Q6.1: Effect of Sudden Volume Increase on Liquid–Vapour Equilibrium

A liquid is in equilibrium with its vapour in a sealed container at a fixed temperature. The volume of the container is suddenly increased. (a) What is the initial effect of the change on vapour pressure? (b) How do the rates of evaporation and condensation change initially? (c) What happens when equilibrium is restored finally, and what will be the final vapour pressure?

(a) The moment the volume is increased, the same number of vapour molecules now occupy a larger space, so the vapour pressure initially *decreases*. (b) The rate of evaporation depends only on the temperature and the exposed liquid surface, both of which are unchanged, so it initially stays the same. The rate of condensation, however, depends on how often vapour molecules strike the liquid surface; since the vapour is now more spread out (lower density), collisions with the surface become less frequent, so the rate of condensation initially decreases. (c) Because the rate of evaporation now exceeds the rate of condensation, more liquid evaporates until the two rates become equal again and a new dynamic equilibrium is established. Since vapour pressure is a function of temperature alone (not volume) for a pure liquid, the system returns to exactly the same vapour pressure as before the volume change once equilibrium is restored.

Final answer: vapour pressure initially drops, condensation rate initially drops while evaporation rate is unchanged, and the final equilibrium vapour pressure equals the original vapour pressure because it depends only on temperature.

Q6.2: Calculating K_c for the SO_2–O_2–SO_3 Equilibrium

What is K_c for the following equilibrium when the equilibrium concentration of each substance is $[SO_2]$ = 0.60 M, $[O_2]$ = 0.82 M, and $[SO_3]$ = 1.90 M? $2SO_2(g) + O_2(g) \leftrightarrow 2SO_3(g)$

By the law of mass action, $K_c = [SO_3]^2 / ([SO_2]^2[O_2])$. Substituting the given values: $K_c = (1.90)^2 / [(0.60)^2 \times 0.82] = 3.61 / 0.2952$.

$K_c \approx 12.2\ M^{-1}$.

Q6.3: K_p for Iodine Vapour Dissociation

At a certain temperature and total pressure of 10^5 Pa, iodine vapour contains 40% by volume of I atoms. $I_2(g) \leftrightarrow 2I(g)$. Calculate K_p for the equilibrium.

Since I atoms make up 40% of the mixture by volume (equivalent to mole/pressure fraction), the partial

pressure of I is 40% of 10^5 Pa = 4×10^4 Pa, and the partial pressure of I_2 is the remaining 60%: 6×10^4 Pa. Then $K_p = (p_I)^2 / p_{I2} = (4 \times 10^4)^2 / (6 \times 10^4) = 16 \times 10^8 / 6 \times 10^4$.

$K_p \approx 2.67 \times 10^4$ Pa.

Q6.4: Writing K_c Expressions for Five Reactions

Write the expression for the equilibrium constant, K_c, for each of the following reactions: (i) $2NOCl(g) \leftrightarrow 2NO(g) + Cl_2(g)$ (ii) $2Cu(NO_3)_2(s) \leftrightarrow 2CuO(s) + 4NO_2(g) + O_2(g)$ (iii) $CH_3COOC_2H_5(aq) + H_2O(l) \leftrightarrow CH_3COOH(aq) + C_2H_5OH(aq)$ (iv) $Fe^{3+}(aq) + 3OH^-(aq) \leftrightarrow Fe(OH)_3(s)$ (v) $I_2(s) + 5F_2(g) \leftrightarrow 2IF_5(g)$

Pure solids and pure liquids are omitted from equilibrium-constant expressions because their "concentration" (density ÷ molar mass) is a fixed constant already absorbed into K itself. Applying the law of mass action to only the gaseous/aqueous species that can genuinely vary in concentration: (i) $K_c = [NO]^2[Cl_2] / [NOCl]^2$. (ii) Since $Cu(NO_3)_2$ and CuO are both solids, $K_c = [NO_2]^4[O_2]$. (iii) Here water is a genuine reactant (not present in large excess as a solvent), so it is included: $K_c = [CH_3COOH][C_2H_5OH] / ([CH_3COOC_2H_5][H_2O])$. (iv) $Fe(OH)_3$ is a solid and is omitted, and OH^- carries a stoichiometric coefficient of 3, so $K_c = 1 / ([Fe^{3+}][OH^-]^3)$. (v) I_2 is a solid and is omitted: $K_c = [IF_5]^2 / [F_2]^5$.

The five K_c expressions are as derived above, with all pure solids correctly excluded.

Q6.5: Converting K_p to K_c for Two Equilibria

Find out the value of K_c for each of the following equilibria from the value of K_p: (i) $2NOCl(g) \leftrightarrow 2NO(g) + Cl_2(g)$; $K_p = 1.8 \times 10^{-2}$ at 500 K (ii) $CaCO_3(s) \leftrightarrow CaO(s) + CO_2(g)$; $K_p = 167$ at 1073 K

Using $K_p = K_c(RT)^{\Delta n}$, so $K_c = K_p / (RT)^{\Delta n}$, with R = 0.0831 L bar K^{-1} mol^{-1}. For (i), $\Delta n = 3 - 2 = 1$, so $K_c = 1.8 \times 10^{-2} / (0.0831 \times 500) = 1.8 \times 10^{-2} / 41.55$. For (ii), $\Delta n = 1 - 0 = 1$, so $K_c = 167 / (0.0831 \times 1073) = 167 / 89.17$.

(i) $K_c \approx 4.33 \times 10^{-4}$ mol L^{-1}; (ii) $K_c \approx 1.87$ mol L^{-1}.

Q6.6: K_c for the Reverse Reaction

For the equilibrium $NO(g) + O_3(g) \leftrightarrow NO_2(g) + O_2(g)$, $K_c = 6.3 \times 10^{14}$ at 1000 K. Both the forward and reverse reactions are elementary bimolecular reactions. What is K_c for the reverse reaction?

The equilibrium constant of the reverse reaction is simply the reciprocal of the forward one: $K_c' = 1 / K_c = 1 / (6.3 \times 10^{14})$.

$K_c' \approx 1.59 \times 10^{-15}$.

Q6.7: Why Pure Liquids and Solids Are Omitted from K

Explain why pure liquids and solids can be ignored while writing the equilibrium constant expression.

The "concentration" of a pure solid or pure liquid, defined as its density divided by its molar mass, is a fixed physical property at a given temperature — it does not change no matter how much of the solid or liquid is

present, because adding more of it only increases the amount of substance, not its concentration per unit volume. Since this quantity is already a constant, it is mathematically absorbed into the equilibrium constant itself rather than written out separately.

Pure solids and liquids have fixed, unchanging molar concentrations, so they are merged into the value of K rather than appearing explicitly in its expression.

Q6.8: Equilibrium Composition for N_2–O_2–N_2O with a Tiny K_c

$2N_2(g) + O_2(g) \leftrightarrow 2N_2O(g)$. If a mixture of 0.482 mol N_2 and 0.933 mol O_2 is placed in a 10 L vessel and allowed to reach equilibrium at a temperature where $K_c = 2.0 \times 10^{-37}$, determine the composition of the equilibrium mixture.

Initial concentrations are $[N_2] = 0.0482$ M and $[O_2] = 0.0933$ M. Because K_c is astronomically small, only a minuscule amount of N_2O forms, so $[N_2]$ and $[O_2]$ remain essentially unchanged at equilibrium. Let $[N_2O] = x$. Then $K_c = x^2 / ([N_2]^2[O_2])$, so $x^2 = 2.0 \times 10^{-37} \times (0.0482)^2 \times 0.0933 = 2.0 \times 10^{-37} \times 2.324 \times 10^{-3} \times 0.0933 \approx 4.33 \times 10^{-41}$. Taking the square root, $x = [N_2O] \approx 6.6 \times 10^{-21}$ M.

At equilibrium, $[N_2] \approx 0.0482$ M, $[O_2] \approx 0.0933$ M, and $[N_2O] \approx 6.6 \times 10^{-21}$ M — confirming the reaction barely proceeds forward.

Q6.9: Equilibrium Amounts of NO and Br_2 in the NOBr Reaction

$2NO(g) + Br_2(g) \leftrightarrow 2NOBr(g)$. When 0.087 mol NO and 0.0437 mol Br_2 are mixed in a closed container at constant temperature, 0.0518 mol of NOBr is obtained at equilibrium. Calculate the equilibrium amounts of NO and Br_2.

Since 2 mol NOBr forms from 2 mol NO, the moles of NO consumed equal the moles of NOBr formed: 0.0518 mol. Since 2 mol NOBr forms from 1 mol Br_2, the Br_2 consumed is 0.0518 / 2 = 0.0259 mol. Subtracting from the initial amounts: NO remaining = 0.087 − 0.0518 = 0.0352 mol; Br_2 remaining = 0.0437 − 0.0259 = 0.0178 mol.

At equilibrium, NO = 0.0352 mol and Br_2 = 0.0178 mol.

Q6.10: K_c from K_p for the SO_2 Oxidation Equilibrium

$2SO_2(g) + O_2(g) \leftrightarrow 2SO_3(g)$. At 450 K, $K_p = 2.0 \times 10^{10}$ bar^{-1}. What is K_c at this temperature?

Here $\Delta n = 2 - 3 = -1$, so $K_p = K_c/(RT)$, which rearranges to $K_c = K_p \times RT$. Substituting: $K_c = 2.0 \times 10^{10} \times 0.0831 \times 450 = 2.0 \times 10^{10} \times 37.4$.

$K_c \approx 7.48 \times 10^{11}$ mol^{-1} L. (Note: because Δn is negative, K_c must be obtained by *multiplying* K_p by RT, not dividing — a common source of error.)

Q6.11: K_p for the Dissociation of HI

A sample of HI(g) is placed in a flask at a pressure of 0.2 atm. At equilibrium, the partial pressure of HI(g) is 0.04 atm. What is K_p for $2HI(g) \leftrightarrow H_2(g) + I_2(g)$?

The drop in HI pressure is 0.2 – 0.04 = 0.16 atm. Since 2 mol HI produces 1 mol each of H_2 and I_2, the partial pressures of H_2 and I_2 at equilibrium are each half of that drop: 0.16 / 2 = 0.08 atm. Then $K_p = (p_{H2} \times p_{I2}) / (p_{HI})^2 = (0.08 \times 0.08) / (0.04)^2 = 0.0064 / 0.0016$.

$K_p = 4$.

Q6.12: Checking Whether an NH_3 Mixture Is at Equilibrium

A mixture of 1.57 mol N_2, 1.92 mol H_2, and 8.13 mol NH_3 is placed in a 20 L vessel at 500 K, where K_c for $N_2(g) + 3H_2(g) \leftrightarrow 2NH_3(g)$ is 1.7×10^2. Is the mixture at equilibrium? If not, in which direction does the reaction proceed?

Concentrations: $[N_2] = 1.57/20 = 0.0785$ M, $[H_2] = 1.92/20 = 0.096$ M, $[NH_3] = 8.13/20 = 0.4065$ M. The reaction quotient is $Q_c = [NH_3]^2 / ([N_2][H_2]^3) = (0.4065)^2 / (0.0785 \times (0.096)^3) = 0.1652 / (0.0785 \times 8.847 \times 10^{-4}) = 0.1652 / 6.945 \times 10^{-5}$.

$Q_c \approx 2.4 \times 10^3$, which is far greater than $K_c = 1.7 \times 10^2$. Since $Q_c > K_c$, the mixture is not at equilibrium and the reaction will proceed in the reverse (backward) direction, converting NH_3 back into N_2 and H_2.

Q6.13: Deriving the Balanced Equation from a K_c Expression

The equilibrium constant expression for a gas reaction is $K_c = [NH_3]^4[O_2]^5 / ([NO]^4[H_2O]^6)$. Write the balanced chemical equation corresponding to this expression.

Since NH_3 and O_2 appear in the numerator, they are the products, each with the stoichiometric coefficient shown as its exponent (4 and 5); NO and H_2O, in the denominator, are the reactants with coefficients 4 and 6. Checking atom balance: N: 4 (from NO) = 4 (in NH_3) ✓; H: 6 × 2 = 12 (from H_2O) = 4 × 3 = 12 (in NH_3) ✓; O: 4 + 6 = 10 (from NO + H_2O) = 5 × 2 = 10 (in O_2) ✓.

$4NO(g) + 6H_2O(g) \leftrightarrow 4NH_3(g) + 5O_2(g)$.

Q6.14: K_c for the Water-Gas Shift Reaction

One mole of H_2O and one mole of CO are taken in a 10 L vessel and heated to 725 K. At equilibrium, 40% of the water (by mass) reacts with CO according to $H_2O(g) + CO(g) \leftrightarrow H_2(g) + CO_2(g)$. Calculate the equilibrium constant for the reaction.

Initial concentrations: $[H_2O] = [CO] = 1/10 = 0.1$ M. Since 40% reacts, the amount converted is $0.4 \times 0.1 = 0.04$ M. At equilibrium: $[H_2O] = [CO] = 0.1 - 0.04 = 0.06$ M, and $[H_2] = [CO_2] = 0.04$ M. Then $K_c = ([H_2][CO_2]) / ([H_2O][CO]) = (0.04 \times 0.04) / (0.06 \times 0.06) = 0.0016 / 0.0036$.

$K_c \approx 0.44$.

Q6.15: Finding $[H_2]$ and $[I_2]$ at Equilibrium from HI Decomposition

At 700 K, K_c for $H_2(g) + I_2(g) \leftrightarrow 2HI(g)$ is 54.8. If 0.5 mol L^{-1} of HI is present at equilibrium, starting purely from HI, what are the equilibrium concentrations of H_2 and I_2?

Since the system was built up starting only from HI, equal amounts of H_2 and I_2 must have formed; let each equal x. For the decomposition direction, $K_c' = 1/K_c = 1/54.8$, and $K_c' = ([H_2][I_2]) / [HI]^2 = x^2 / (0.5)^2$. So $x^2 = 0.25/54.8 = 4.562 \times 10^{-3}$, giving $x = \sqrt{(4.562 \times 10^{-3})}$.

$[H_2] = [I_2] \approx 0.0676$ mol L^{-1}.

Q6.16: Equilibrium Concentrations for the ICl Decomposition

$2ICl(g) \leftrightarrow I_2(g) + Cl_2(g)$; $K_c = 0.14$. What is the equilibrium concentration of each substance when the initial concentration of ICl was 0.78 M?

Let $x = [I_2] = [Cl_2]$ formed at equilibrium; since 2 mol ICl is consumed per mole of I_2 (or Cl_2) formed, [ICl] at equilibrium = 0.78 – 2x. Then $K_c = x^2 / (0.78 - 2x)^2 = 0.14$, so $\sqrt{0.14} = 0.374 = x / (0.78 - 2x)$. Solving: $x = 0.374(0.78 - 2x) = 0.2917 - 0.748x$, giving $1.748x = 0.2917$, so $x \approx 0.167$.

$[I_2] = [Cl_2] \approx 0.167$ M, and $[ICl] = 0.78 - 2(0.167) \approx 0.446$ M.

Q6.17: Equilibrium Pressure of C_2H_6 from Ethane Cracking

$K_p = 0.04$ atm at 899 K for $C_2H_6(g) \leftrightarrow C_2H_4(g) + H_2(g)$. What is the equilibrium pressure of C_2H_6 when it is placed in a flask at 4.0 atm and allowed to reach equilibrium?

Let p = the pressure of C_2H_4 (= pressure of H_2) formed at equilibrium. Then $K_p = p^2 / (4 - p) = 0.04$, giving $p^2 + 0.04p - 0.16 = 0$. Using the quadratic formula, $p = [-0.04 \pm \sqrt{(0.0016 + 0.64)}] / 2 = [-0.04 \pm 0.801] / 2$, taking the positive root, $p \approx 0.38$ atm.

p(C_2H_6) at equilibrium = 4 – 0.38 = 3.62 atm.

Q6.18: Reaction Quotient and Equilibrium Constant for Ethyl Acetate Formation

Ethyl acetate is formed by the reaction of ethanol and acetic acid: $CH_3COOH(l) + C_2H_5OH(l) \leftrightarrow CH_3COOC_2H_5(l) + H_2O(l)$, where water is not present in excess. (i) Write Q_c for this reaction. (ii) At 293 K, starting with 1.00 mol acetic acid and 0.18 mol ethanol, 0.171 mol ethyl acetate is present at equilibrium; calculate K_c. (iii) Starting instead with 0.5 mol ethanol and 1.0 mol acetic acid at 293 K, 0.214 mol ethyl acetate is found after some time. Has equilibrium been reached?

(i) Since water is treated as a genuine participant here, $Q_c = ([CH_3COOC_2H_5][H_2O]) / ([CH_3COOH][C_2H_5OH])$. (ii) Because the numbers of moles are the same on both sides of the reaction, the common volume V cancels out of the K_c expression. At equilibrium: acetic acid = 1 – 0.171 = 0.829 mol, ethanol = 0.18 – 0.171 = 0.009 mol, ester = water = 0.171 mol. So $K_c = (0.171 \times 0.171) / (0.829 \times 0.009) = 0.02924 / 0.007461$. (iii) At this snapshot: acetic acid = 1 – 0.214 = 0.786 mol, ethanol = 0.5 – 0.214 = 0.286 mol, ester = water = 0.214 mol. So $Q_c = (0.214 \times 0.214) / (0.786 \times 0.286) = 0.04580 / 0.22480$.

(ii) $K_c \approx 3.92$. (iii) $Q_c \approx 0.204$, which is much less than $K_c = 3.92$, so equilibrium has NOT yet been reached — the reaction will continue to move forward.

Q6.19: Equilibrium Concentrations of PCl_3 and Cl_2 from PCl_5 Decomposition

$PCl_5(g) \leftrightarrow PCl_3(g) + Cl_2(g)$. A sample of pure PCl_5 was placed in an evacuated vessel at 473 K; at equilibrium, $[PCl_5] = 0.05$ mol L^{-1} and $K_c = 8.3 \times 10^{-3}$. What are the equilibrium concentrations of PCl_3 and Cl_2?

Let $x = [PCl_3] = [Cl_2]$ at equilibrium. Then $K_c = x^2 / [PCl_5] = x^2 / 0.05 = 8.3 \times 10^{-3}$, so $x^2 = 4.15 \times 10^{-4}$, giving $x = \sqrt{(4.15 \times 10^{-4})}$.

$[PCl_3] = [Cl_2] \approx 0.0204$ mol L^{-1}.

Q6.20: Equilibrium Partial Pressures in Iron Ore Reduction

$FeO(s) + CO(g) \leftrightarrow Fe(s) + CO_2(g)$; $K_p = 0.265$ at 1050 K. What are the equilibrium partial pressures of CO and CO_2 if the initial partial pressures are $p_{CO} = 1.4$ atm and $p_{CO2} = 0.80$ atm?

First check the direction: $Q_p = p_{CO2}/p_{CO} = 0.80/1.4 = 0.571$, which is greater than $K_p = 0.265$, so the reaction shifts backward, converting some CO_2 back into CO. Let p = the decrease in p_{CO2} (equal to the increase in p_{CO}). Then $K_p = (0.80 - p) / (1.4 + p) = 0.265$, so $0.80 - p = 0.371 + 0.265p$, giving $0.429 = 1.265p$, so $p \approx 0.339$ atm.

At equilibrium, $p_{CO2} = 0.80 - 0.339 \approx 0.461$ atm and $p_{CO} = 1.4 + 0.339 \approx 1.739$ atm.

Q6.21: Direction of Approach to Equilibrium for Ammonia Synthesis

K_c for $N_2(g) + 3H_2(g) \leftrightarrow 2NH_3(g)$ at 500 K is 0.061. A reaction mixture contains 3.0 mol L^{-1} N_2, 2.0 mol L^{-1} H_2, and 0.5 mol L^{-1} NH_3. Is the mixture at equilibrium? If not, in which direction does it proceed?

$Q_c = [NH_3]^2 / ([N_2][H_2]^3) = (0.5)^2 / (3.0 \times (2.0)^3) = 0.25 / 24$.

$Q_c \approx 0.0104$, which is less than $K_c = 0.061$. Since $Q_c < K_c$, the mixture is not at equilibrium, and the reaction proceeds in the forward direction to form more NH_3.

Q6.22: Equilibrium Concentration of BrCl from Decomposition

$2BrCl(g) \leftrightarrow Br_2(g) + Cl_2(g)$; $K_c = 32$ at 500 K. If pure BrCl is initially present at 3.3×10^{-3} mol L^{-1}, what is its molar concentration at equilibrium?

Let $x = [Br_2] = [Cl_2]$ formed. Then [BrCl] at equilibrium $= 3.3 \times 10^{-3} - 2x$, and $K_c = x^2 / (3.3 \times 10^{-3} - 2x)^2 = 32$, so $\sqrt{32} = 5.657 = x / (3.3 \times 10^{-3} - 2x)$. Solving: $x = 5.657(3.3 \times 10^{-3} - 2x) = 0.01867 - 11.314x$, so $12.314x = 0.01867$, giving $x \approx 1.516 \times 10^{-3}$ M.

[BrCl] at equilibrium $= 3.3 \times 10^{-3} - 2(1.516 \times 10^{-3}) \approx 2.7 \times 10^{-4}$ mol L^{-1}.

Q6.23: K_c for the Boudouard Equilibrium from Percentage Composition

At 1127 K and 1 atm, a gaseous mixture of CO and CO_2 in equilibrium with solid carbon contains 90.55% CO by mass: $C(s) + CO_2(g) \leftrightarrow 2CO(g)$. Calculate K_c at this temperature.

Assume 100 g of gas mixture: CO = 90.55 g, CO^2 = 9.45 g. Moles: CO = 90.55/28 = 3.234 mol, CO_2 = 9.45/44 = 0.215 mol; total = 3.449 mol. Mole (= pressure) fractions give p_{CO} = (3.234/3.449) × 1 ≈ 0.938 atm and p_{CO2} = (0.215/3.449) × 1 ≈ 0.062 atm. Then $K_p = (p_{CO})^2 / p_{CO2} = (0.938)^2 / 0.062 \approx 14.2$. Converting with Δn = 2 – 1 = 1: $K_c = K_p / (RT) = 14.2 / (0.0831 \times 1127) = 14.2 / 93.65$.

$K_c \approx 0.15$ mol L^{-1}.

Q6.24: ΔG° and K for the Formation of NO_2 from NO

Calculate (a) $\Delta_rG°$ and (b) the equilibrium constant for $NO(g) + ½O_2(g) \leftrightarrow NO_2(g)$ at 298 K, given $\Delta_fG°[NO_2]$ = 52.0 kJ/mol, $\Delta_fG°[NO]$ = 87.0 kJ/mol, $\Delta_fG°[O_2]$ = 0 kJ/mol.

(a) $\Delta_rG°$ = $\Delta_fG°$(products) – $\Delta_fG°$(reactants) = 52.0 – (87.0 + 0) = –35.0 kJ/mol. (b) Using ΔG° = –RT ln K: ln K = 35000 / (8.314 × 298) = 35000 / 2477.6 ≈ 14.13, so $K = e^{14.13}$.

(a) $\Delta_rG°$ = –35.0 kJ/mol. (b) K ≈ 1.37×10^6, a large value consistent with the strongly negative ΔG°.

Q6.25: Effect of Decreasing Pressure on Three Equilibria

Does the number of moles of reaction products increase, decrease, or remain the same when each of the following equilibria is subjected to a decrease in pressure (by increasing volume)? (a) $PCl_5(g) \leftrightarrow PCl_3(g) + Cl_2(g)$ (b) $CaO(s) + CO_2(g) \leftrightarrow CaCO_3(s)$ (c) $3Fe(s) + 4H_2O(g) \leftrightarrow Fe_3O_4(s) + 4H_2(g)$

By Le Chatelier's principle, decreasing pressure shifts an equilibrium toward the side with the greater number of gas moles. (a) Reactant side has 1 gas mole, product side has 2, so the equilibrium shifts forward. (b) Reactant side has 1 gas mole (CO_2), product side has 0 ($CaCO_3$ is solid), so it shifts backward, away from the product. (c) Both sides have 4 gas moles ($4H_2O$ vs $4H_2$), so pressure has no effect.

(a) Products increase. (b) Products ($CaCO_3$) decrease. (c) Products remain the same.

Q6.26: Which Reactions Are Affected by Increasing Pressure

Which of the following reactions will be affected by increasing the pressure, and in which direction? (i) $COCl_2(g) \leftrightarrow CO(g) + Cl_2(g)$ (ii) $CH_4(g) + 2S_2(g) \leftrightarrow CS_2(g) + 2H_2S(g)$ (iii) $CO_2(g) + C(s) \leftrightarrow 2CO(g)$ (iv) $2H_2(g) + CO(g) \leftrightarrow CH_3OH(g)$ (v) $CaCO_3(s) \leftrightarrow CaO(s) + CO_2(g)$ (vi) $4NH_3(g) + 5O_2(g) \leftrightarrow 4NO(g) + 6H_2O(g)$

Increasing pressure shifts equilibrium toward the side with fewer gas moles. Counting gas moles on each side: (i) 1 → 2, shifts backward. (ii) 3 → 3, equal, unaffected. (iii) 1 → 2, shifts backward. (iv) 3 → 1, shifts forward. (v) 0 → 1, shifts backward. (vi) 9 → 10, shifts backward.

Reactions (i), (iii), (v), and (vi) shift backward; reaction (iv) shifts forward; reaction (ii) is unaffected because gas moles are equal on both sides.

Q6.27: Equilibrium Pressures from Pure HBr Introduced into a Vessel

The equilibrium constant for $H_2(g) + Br_2(g) \leftrightarrow 2HBr(g)$ is 1.6×10^5 at 1024 K. Find the equilibrium pressure

of all gases if 10.0 bar of HBr is introduced into a sealed container at 1024 K.

Working with the reverse (decomposition) direction, $2HBr(g) \leftrightarrow H_2(g) + Br_2(g)$, $K_p' = 1/(1.6 \times 10^5) = 6.25 \times 10^{-6}$. Let p = pressure of H_2 = pressure of Br_2 formed. Then $K_p' = p^2 / (10 - 2p)^2$, so $\sqrt{(6.25 \times 10^{-6})} = 2.5 \times 10^{-3} = p / (10 - 2p)$. Solving: $p = 2.5 \times 10^{-3}(10 - 2p) = 0.025 - 0.005p$, so $1.005p = 0.025$, giving $p \approx 0.0249$ bar.

$p(H_2) = p(Br_2) \approx 0.025$ bar, and $p(HBr) = 10 - 2(0.025) \approx 9.95$ bar — because K_p strongly favours HBr formation, only a tiny fraction of it decomposes back to H_2 and Br_2.

Q6.28: K_p Expression and Le Chatelier Effects for Steam-Methane Reforming

Dihydrogen is obtained from natural gas by the endothermic reaction $CH_4(g) + H_2O(g) \leftrightarrow CO(g) + 3H_2(g)$. (a) Write an expression for K_p. (b) How will K_p and the composition of the equilibrium mixture be affected by (i) increasing pressure, (ii) increasing temperature, and (iii) using a catalyst?

(a) $K_p = (p_{CO} \times p_{H2}^3) / (p_{CH4} \times p_{H2O})$. (b)(i) The gas moles go from 2 (reactants) to 4 (products), so increasing pressure shifts the composition backward toward CH_4 and H_2O; however, because K_p depends only on temperature, its numerical value is unchanged by pressure. (ii) Since the reaction is endothermic, raising the temperature shifts equilibrium forward and genuinely increases the value of K_p. (iii) A catalyst speeds up the attainment of equilibrium equally in both directions but changes neither K_p nor the equilibrium composition.

$K_p = p_{CO}p_{H2}^3/(p_{CH4}p_{H2O})$; pressure shifts composition backward but leaves K_p unchanged, higher temperature increases K_p and shifts forward, and a catalyst affects neither K_p nor composition.

Q6.29: Le Chatelier Effects on Methanol Synthesis

Describe the effect of (a) addition of H_2, (b) addition of CH_3OH, (c) removal of CO, and (d) removal of CH_3OH on the equilibrium $2H_2(g) + CO(g) \leftrightarrow CH_3OH(g)$.

Le Chatelier's principle predicts that a system disturbed from equilibrium shifts to partially counteract the disturbance. (a) Adding H_2 (a reactant) shifts the equilibrium forward, consuming some of the added H_2. (b) Adding CH_3OH (the product) shifts the equilibrium backward. (c) Removing CO (a reactant) shifts the equilibrium backward, to partially replace the lost CO. (d) Removing CH_3OH (the product) shifts the equilibrium forward, to partially replace it.

(a) Forward shift. (b) Backward shift. (c) Backward shift. (d) Forward shift.

Q6.30: K_c, Reverse K_c, and Temperature Effects for PCl_5 Decomposition

At 473 K, K_c for $PCl_5(g) \leftrightarrow PCl_3(g) + Cl_2(g)$, $\Delta_r H° = 124.0$ kJ mol^{-1}, is 8.3×10^{-3}. (a) Write the K_c expression. (b) What is K_c for the reverse reaction? (c) What is the effect on K_c of (i) adding more PCl_5, (ii) increasing pressure, (iii) increasing temperature?

(a) $K_c = ([PCl_3][Cl_2]) / [PCl_5]$. (b) The reverse equilibrium constant is the reciprocal: $K_c' = 1 / (8.3 \times 10^{-3}) \approx 120.5$. (c) Since K depends only on temperature: (i) adding more PCl_5 shifts the equilibrium position forward but leaves K_c unchanged. (ii) Increasing pressure shifts the composition backward (toward fewer gas

moles) but again leaves K_c unchanged. (iii) Because the reaction is endothermic ($\Delta H°$ positive), raising the temperature genuinely increases K_c.

(a) $K_c = [PCl_3][Cl_2]/[PCl_5]$. (b) $K_c' \approx 1.2 \times 10^2$. (c) K_c is unaffected by adding PCl_5 or by pressure, but increases with rising temperature.

Q6.31: Partial Pressure of H_2 in the Water-Gas Shift Reaction

$CO(g) + H_2O(g) \leftrightarrow CO_2(g) + H_2(g)$. If a vessel at 400°C is charged with an equimolar mixture such that $p_{CO} = p_{H2O} = 4.0$ bar, what is the partial pressure of H_2 at equilibrium given $K_p = 10.1$?

Let p = the pressure of CO_2 (= pressure of H_2) formed at equilibrium. Then $K_p = p^2 / (4 – p)^2 = 10.1$, so $\sqrt{10.1} = 3.178 = p / (4 – p)$. Solving: $p = 3.178(4 – p) = 12.712 – 3.178p$, giving $4.178p = 12.712$, so $p \approx 3.04$ bar.

The equilibrium partial pressure of H_2 is approximately 3.04 bar.

Q6.32: Identifying the Reaction with Appreciable Reactant and Product Concentrations

Predict which of the following reactions will have appreciable concentrations of both reactants and products at equilibrium: (a) $Cl_2(g) \leftrightarrow 2Cl(g)$; $K_c = 5 \times 10^{-39}$ (b) $Cl_2(g) + 2NO(g) \leftrightarrow 2NOCl(g)$; $K_c = 3.7 \times 10^8$ (c) $Cl_2(g) + 2NO_2(g) \leftrightarrow 2NO_2Cl(g)$; $K_c = 1.8$

As a rule of thumb, an equilibrium contains appreciable amounts of both reactants and products only when K_c is roughly between 10^{-3} and 10^3. Reaction (a) has a vanishingly small K_c, so it barely proceeds beyond reactants. Reaction (b) has an enormous K_c, so it goes essentially to completion. Only reaction (c), with $K_c = 1.8$, falls within the moderate range.

Only reaction (c), $Cl_2 + 2NO_2 \leftrightarrow 2NO_2Cl$, has appreciable concentrations of both reactants and products at equilibrium.

Q6.33: Equilibrium Ozone Concentration in Air

K_c for $3O_2(g) \leftrightarrow 2O_3(g)$ is 2.0×10^{-50} at 25°C. If $[O_2]$ in air is 1.6×10^{-2} M, what is the equilibrium concentration of O_3?

$K_c = [O_3]^2 / [O_2]^3$, so $[O_3]^2 = K_c \times [O_2]^3 = 2.0 \times 10^{-50} \times (1.6 \times 10^{-2})^3 = 2.0 \times 10^{-50} \times 4.096 \times 10^{-6} = 8.19 \times 10^{-56}$. Taking the square root gives $[O_3]$.

$[O_3] \approx 2.86 \times 10^{-28}$ M, showing that even though ozone is vital in the upper atmosphere, its equilibrium concentration relative to O_2 is extraordinarily small.

Q6.34: Equilibrium Methane Concentration in the Sabatier-Type Reaction

$CO(g) + 3H_2(g) \leftrightarrow CH_4(g) + H_2O(g)$ is at equilibrium at 1300 K in a 1 L flask containing 0.30 mol CO, 0.10 mol H_2, 0.02 mol H_2O, and an unknown amount of CH_4. $K_c = 3.90$. Determine $[CH_4]$.

Let $x = [CH_4]$. Then $K_c = ([CH_4][H_2O]) / ([CO][H_2]^3)$, so $3.90 = (x \times 0.02) / (0.30 \times (0.10)^3) = 0.02x / (0.30 \times$

0.001) = 0.02x / 3 × 10^{-4}. Solving, x = 3.90 × 3 × 10^{-4} / 0.02.

[CH_4] ≈ 5.85 × 10^{-2} mol L^{-1}.

Q6.35: Identifying Conjugate Acid–Base Pairs

What is meant by a conjugate acid–base pair? Find the conjugate acid or base for the following species: HNO_2, CN^-, $HClO_4$, F^-, OH^-, CO_3^{2-}, and S^{2-}.

A conjugate acid–base pair consists of two species that differ from each other by exactly one proton (H^+) — for example, HCl and Cl^-, or H_3O^+ and H_2O. Applying this: HNO_2 (an acid) loses a proton to give its conjugate base NO_2^-. CN^- (a base) gains a proton to give its conjugate acid HCN. $HClO_4$ loses a proton to give ClO_4^-. F^- gains a proton to give HF. OH^-, acting as a base, gains a proton to give H_2O. CO_3^{2-} gains a proton to give HCO_3^-. S^{2-} gains a proton to give HS^-.

Conjugate base of HNO_2 is NO_2^-; conjugate acid of CN^- is HCN; conjugate base of $HClO_4$ is ClO_4^-; conjugate acid of F^- is HF; conjugate acid of OH^- is H_2O; conjugate acid of CO_3^{2-} is HCO_3^-; conjugate acid of S^{2-} is HS^-.

Q6.36: Identifying Lewis Acids

Which of the following are Lewis acids? H_2O, BF_3, H^+, and NH_4^+.

A Lewis acid is a species that can accept an electron pair. BF_3 has an incomplete octet on boron (only six electrons), giving it a vacant orbital that readily accepts an electron pair, making it a classic Lewis acid. H^+ has a completely empty 1s orbital and readily accepts a lone pair, making it a Lewis acid as well. H_2O, by contrast, has lone pairs to donate and acts as a Lewis base. NH_4^+, although positively charged, has nitrogen with a complete octet and no accessible vacant orbital, so it does not behave as a simple Lewis acid in the same direct sense.

BF_3 and H^+ are Lewis acids (electron-pair acceptors); H_2O is a Lewis base.

Q6.37: Conjugate Bases of Three Brønsted Acids

What are the conjugate bases for the Brønsted acids HF, H_2SO_4, and HCO_3^-?

Removing one proton from each acid gives its conjugate base: HF loses H^+ to give F^-; H_2SO_4 loses H^+ to give HSO_4^-; HCO_3^- loses H^+ to give CO_3^{2-}.

Conjugate bases: F^- (from HF), HSO_4^- (from H_2SO_4), and CO_3^{2-} (from HCO_3^-).

Q6.38: Conjugate Acids of Three Brønsted Bases

Write the conjugate acids for the Brønsted bases NH_2^-, NH_3, and $HCOO^-$.

Adding one proton to each base gives its conjugate acid: NH_2^- gains H^+ to give NH_3; NH_3 gains H^+ to give NH_4^+; $HCOO^-$ gains H^+ to give HCOOH.

Conjugate acids: NH_3 (from NH_2^-), NH_4^+ (from NH_3), and HCOOH (from $HCOO^-$).

Q6.39: Conjugate Acids and Bases of Four Amphoteric Species

The species H_2O, HCO_3^-, HSO_4^-, and NH_3 can act as both Brønsted acids and bases. Give the corresponding conjugate acid and conjugate base for each.

Each species can either donate a proton (acting as an acid, forming its conjugate base) or accept one (acting as a base, forming its conjugate acid). For H_2O: conjugate acid H_3O^+, conjugate base OH^-. For HCO_3^-: conjugate acid H_2CO_3, conjugate base CO_3^{2-}. For HSO_4^-: conjugate acid H_2SO_4, conjugate base SO_4^{2-}. For NH_3: conjugate acid NH_4^+, conjugate base NH_2^-.

$H_2O \rightarrow H_3O^+/OH^-$; $HCO_3^- \rightarrow H_2CO_3/CO_3^{2-}$; $HSO_4^- \rightarrow H_2SO_4/SO_4^{2-}$; $NH_3 \rightarrow NH_4^+/NH_2^-$.

Q6.40: Classifying Four Species as Lewis Acids or Bases

Classify the following species into Lewis acids and Lewis bases, and show how each acts as such: (a) OH^- (b) F^- (c) H^+ (d) BCl_3.

A Lewis base donates an electron pair; a Lewis acid accepts one. (a) OH^- has lone pairs on oxygen available to donate, so it is a Lewis base. (b) F^- similarly has lone pairs to donate, so it is a Lewis base. (c) H^+ has an empty 1s orbital that accepts an electron pair, so it is a Lewis acid. (d) BCl_3 has boron with only six electrons around it (an incomplete octet), giving it a vacant orbital that accepts an electron pair, so it is a Lewis acid.

Lewis bases: OH^- and F^- (electron-pair donors). Lewis acids: H^+ and BCl_3 (electron-pair acceptors).

Q6.41: pH from Hydrogen Ion Concentration in Soft Drink

The concentration of hydrogen ion in a sample of soft drink is 3.8×10^{-3} M. What is its pH?

$pH = -\log[H^+] = -\log(3.8 \times 10^{-3}) = 3 - \log(3.8) = 3 - 0.5798$.

pH ≈ 2.42, confirming the soft drink is fairly acidic.

Q6.42: Hydrogen Ion Concentration in Vinegar from pH

The pH of a sample of vinegar is 3.76. Calculate the hydrogen ion concentration.

Since $pH = -\log[H^+]$, $[H^+] = 10^{-pH} = 10^{-3.76} = 10^{0.24} \times 10^{-4}$.

$[H^+] \approx 1.74 \times 10^{-4}$ mol L^{-1}.

Q6.43: K_b of the Conjugate Bases of Three Weak Acids

The ionization constants of HF, HCOOH, and HCN at 298 K are 6.8×10^{-4}, 1.8×10^{-4}, and 4.8×10^{-9} respectively. Calculate the ionization constants of their conjugate bases.

Using $K_b = K_w/K_a$ with $K_w = 1.0 \times 10^{-14}$: $K_b(F^-) = 10^{-14}/(6.8 \times 10^{-4})$; $K_b(HCOO^-) = 10^{-14}/(1.8 \times 10^{-4})$; $K_b(CN^-) = 10^{-14}/(4.8 \times 10^{-9})$.

$K_b(F^-) \approx 1.47 \times 10^{-11}$; $K_b(HCOO^-) \approx 5.56 \times 10^{-11}$; $K_b(CN^-) \approx 2.08 \times 10^{-6}$ — the weakest acid (HCN) has, as expected, the strongest conjugate base.

Q6.44: Phenolate Ion Concentration and the Common-Ion Effect

The ionization constant of phenol is 1.0×10^{-10}. What is the concentration of phenolate ion in a 0.05 M solution of phenol? What will its degree of ionization be if the solution is also 0.01 M in sodium phenolate?

For pure phenol, $C_6H_5OH + H_2O \leftrightarrow C_6H_5O^- + H_3O^+$. Let $x = [C_6H_5O^-]$. Since ionization is small, $K_a = x^2/0.05 = 1.0 \times 10^{-10}$, so $x^2 = 5.0 \times 10^{-12}$, giving $x \approx 2.24 \times 10^{-6}$ M. When 0.01 M sodium phenolate (a strong electrolyte) is also present, it supplies an initial 0.01 M of $C_6H_5O^-$, suppressing further ionization of phenol by the common-ion effect. Let y = additional ionization. Then $K_a = (0.01 + y)(y)/0.05 \approx (0.01)(y)/0.05 = 1.0 \times 10^{-10}$, so $y = 1.0 \times 10^{-10} \times 0.05/0.01 = 5.0 \times 10^{-10}$. The degree of ionization is then y/0.05.

Without sodium phenolate, $[C_6H_5O^-] \approx 2.24 \times 10^{-6}$ M. With 0.01 M sodium phenolate present, the degree of ionization of phenol drops to about 1×10^{-8} (roughly 10,000 times smaller), a clear demonstration of the common-ion effect.

Q6.45: pH and Degree of Hydrolysis of Sodium Nitrite Solution

The ionization constant of nitrous acid is 4.5×10^{-4}. Calculate the pH of a 0.04 M sodium nitrite solution and its degree of hydrolysis.

$NaNO_2$ is the salt of a weak acid (HNO_2) and a strong base (NaOH), so the nitrite ion hydrolyses: $NO_2^- + H_2O \leftrightarrow HNO_2 + OH^-$. Its hydrolysis constant is $K_b = K_w/K_a = (1.0 \times 10^{-14})/(4.5 \times 10^{-4}) \approx 2.22 \times 10^{-11}$. Using $[OH^-] = \sqrt{(K_b \times C)} = \sqrt{(2.22 \times 10^{-11} \times 0.04)} = \sqrt{(8.89 \times 10^{-13})} \approx 9.43 \times 10^{-7}$ M. Then $pOH = -\log(9.43 \times 10^{-7}) \approx 6.03$, so $pH = 14 - 6.03 \approx 7.97$. The degree of hydrolysis is $h = \sqrt{(K_b/C)} = \sqrt{(2.22 \times 10^{-11}/0.04)} = \sqrt{(5.56 \times 10^{-10})}$.

pH ≈ 7.97 (mildly basic, as expected for the salt of a weak acid and a strong base), and the degree of hydrolysis $h \approx 2.36 \times 10^{-5}$.

Q6.46: Degree of Dissociation and pH of Acetic Acid

The ionization constant of acetic acid is 1.74×10^{-5}. Calculate the degree of dissociation of acetic acid in its 0.05 M solution and the concentration of acetate ion and pH of the solution.

For a weak acid at low degree of dissociation, $K_a \approx \alpha^2 C$, so $\alpha = \sqrt{(K_a/C)} = \sqrt{(1.74 \times 10^{-5}/0.05)} = \sqrt{(3.48 \times 10^{-4})} \approx 0.01865$, i.e. about 1.87%. The acetate ion concentration is $[CH_3COO^-] = \alpha C = 0.01865 \times 0.05 \approx 9.33 \times 10^{-4}$ M, which equals $[H^+]$. Then $pH = -\log(9.33 \times 10^{-4}) \approx 3.03$.

Degree of dissociation α ≈ 0.0187 (1.87%); $[CH_3COO^-] \approx 9.33 \times 10^{-4}$ M; pH ≈ 3.03.

Q6.47: Finding K_a and pK_a of an Organic Acid from Its pH

The pH of a 0.01 M solution of an organic acid is 4.15. Calculate the concentration of the anion, the ionization constant of the acid, and its pK_a.

$[H^+] = 10^{-4.15} = 10^{0.85} \times 10^{-5} \approx 7.08 \times 10^{-5}$ M. Since $HA \leftrightarrow H^+ + A^-$ dissociate 1:1, $[A^-] = [H^+] \approx 7.08 \times 10^{-5}$ M. The undissociated acid remaining is $[HA] \approx 0.01 - 7.08 \times 10^{-5} \approx 9.93 \times 10^{-3}$ M. Then $K_a = [H^+][A^-]/[HA] = (7.08 \times 10^{-5})^2/(9.93 \times 10^{-3}) \approx 5.05 \times 10^{-7}$, and $pK_a = -\log(5.05 \times 10^{-7}) \approx 6.30$.

$[A^-] \approx 7.08 \times 10^{-5}$ M; $K_a \approx 5.0 \times 10^{-7}$; $pK_a \approx 6.30$.

Q6.48: pH of Four Strong Acid and Base Solutions

Assuming complete dissociation, calculate the pH of the following solutions: (a) 0.003 M HCl (b) 0.005 M NaOH (c) 0.002 M HBr (d) 0.002 M KOH.

Since strong acids and bases dissociate completely, $[H^+]$ or $[OH^-]$ equals the stated molarity directly. (a) $[H^+] = 3 \times 10^{-3}$ M, pH = 3 – log 3 = 3 – 0.477. (b) $[OH^-] = 5 \times 10^{-3}$ M, pOH = 3 – log 5 = 3 – 0.699, pH = 14 – pOH. (c) $[H^+] = 2 \times 10^{-3}$ M, pH = 3 – log 2 = 3 – 0.301. (d) $[OH^-] = 2 \times 10^{-3}$ M, pOH = 3 – log 2 = 2.699, pH = 14 – pOH.

(a) pH ≈ 2.52. (b) pH ≈ 11.70. (c) pH ≈ 2.70. (d) pH ≈ 11.30.

Class 11 Chemistry Chapter 6 – Notes and Extra Questions

Once you have worked through every exercise question above, it is worth reinforcing your understanding of Equilibrium with focused revision notes and additional practice beyond the textbook. Key ideas to revisit include the distinction between K_c and K_p, how to build an ICE (Initial–Change–Equilibrium) table for any reversible reaction, the direction rules of Le Chatelier's principle, and the full ionic-equilibrium toolkit — pH, pOH, K_a, K_b, K_w, the common-ion effect, and salt hydrolysis. Check the Class 11 Chemistry Revision Notes and Extra Questions section on this site for a chapter-wise summary sheet and additional numerical practice problems modelled on this exact exercise pattern, useful for both school exams and competitive exams such as JEE and NEET.

What is the difference between K_c and K_p, and how are they related?

K_c is the equilibrium constant expressed in terms of molar concentrations, while K_p is expressed in terms of partial pressures of gaseous species. They are related by $K_p = K_c(RT)^{\Delta n}$, where Δn is the difference between the moles of gaseous products and gaseous reactants, R is the gas constant (0.0831 L bar K^{-1} mol^{-1}), and T is the absolute temperature. When $\Delta n = 0$, K_p and K_c are numerically equal.

How does Le Chatelier's principle predict the direction of a shift in equilibrium?

Le Chatelier's principle states that if a system at equilibrium is subjected to a change in concentration, pressure, volume, or temperature, the equilibrium shifts in the direction that partially counteracts that change. For example, adding a reactant shifts equilibrium forward; increasing pressure shifts equilibrium toward the side with fewer gas moles; and raising the temperature shifts equilibrium in the endothermic direction. Note that only a genuine temperature change alters the numerical value of K itself — changes in concentration, pressure, or the addition of a catalyst change only the position of equilibrium, not K.

What does the pH scale actually measure, and why does pure water have a pH of 7?

The pH scale measures the concentration of hydrogen ions in a solution on a logarithmic basis: pH = $-\log_{10}[H^+]$. Pure water self-ionizes very slightly, $2H_2O \leftrightarrow H_3O^+ + OH^-$, producing equal concentrations of H^+ and OH^-, each equal to 1.0×10^{-7} M at 25 °C. Taking the negative logarithm of 10^{-7} gives a pH of exactly 7, which is why neutral water sits at the midpoint of the 0–14 pH scale at that temperature.

What is the common-ion effect, and why does it suppress the ionization of a weak acid?

The common-ion effect is the suppression of the ionization (or solubility) of a weak electrolyte when a second source of one of its own ions is added to the solution. For a weak acid like phenol, adding sodium phenolate (which fully dissociates to supply extra phenolate ions) increases $[C_6H_5O^-]$ in the solution. Because the equilibrium constant $K_a = [C_6H_5O^-][H^+]/[C_6H_5OH]$ must stay fixed at a given temperature, this extra phenolate ion forces $[H^+]$ (and hence the degree of ionization of phenol) to decrease sharply, exactly as seen in Q6.44 above, where the degree of ionization dropped roughly 10,000-fold in the presence of the common ion.