GENERAL CHEMISTRY.

UNI. CODE - US01CCHE21.

CREDIT - 2. {2 hr. Exam}
MARKS - 70 Marks. External
30 Marks, Internal.

UNIT - 3

IONIC EQULIBRIUM.

In. Aquecus Solution

Ref. Book - Mahan B.H. Uni. chem.
3rd Addition.

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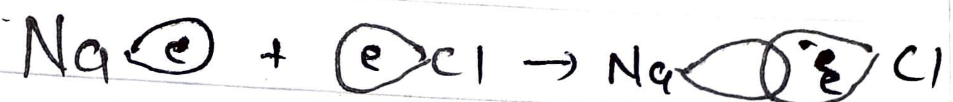
Ionic Equilibrium.

Introduction.

When One of atom donate its Outer electron to another atom. via overlaping of thair orbitals and form a bond is know as ionic BOND

e.g



$Na^+ + Cl^- \rightarrow NaCl.$

* The substence made through ionic bond. such type of substences are known as ionic SUB. means inorgenic substence.

* The deteil study of ionic Sub. is called InOrgenic chemistry.

Generally ionic substence are soluble in aqueous, or ionic solvent via ionization of substence and form a positive & Negative ions. these are known as electrolytes.

e.g $NaCl_{(s)} + H_2O \longrightarrow Na^+_{(aq)} + Cl^-_{(aq)}$

Ionic substances are classifyed on base of their ionization factor.

* Strong electrolyte :- Such type of ionic substances are mostly ionized in aqueous solution e.g strong Acid + strong base salt, NaCl, Na_2SO_4 KBr. KCl .... etc.

* Weak electrolyte :- Such electrolyte substance has lower ionized in solvent Aproximeltly 1 to 20 % ionized.

* Speringly soluble salt :- Such type of electrolyte are very low ionized ~~an~~. so very little amount are soluble in solvent. e.g $PbSO_4$, AgCl, CaF_2, Ag_2CrO_4....

[Speringly soluble salt are ~~soluble~~ in less then 1 gm soluble in 1 lit solvents]

Solubility :- Maximum soluble amount of electrolytic substance per liter solvent is called solubility.

Solybility is denoted by 'S'

Thus we can say that

strong electrolyte has highest solubility while weak electrolytes has lower solubility and spearingly soluble salt has lowest solubility

Ionic Equilibrium and Ksp.

→ Consider AB is a speringly soluble salt, saturated solution of 'AB' in water,' form an ionic equilibrium such as

$$H_2O + AB \rightleftharpoons A^+ + B^-$$

where as

$$AB + H_2O \xrightarrow{Forward} \text{ionization Reaction}$$

$$A^+ + B^- \xrightarrow{\text{Association}} [AB]$$

* thus Forward Reaction is cleavege of ionic bond and Backward Reaction is formation of ionic bond.

We know that

Equilibrium constant K is define as

$$K_a = \frac{[A^+][B^-]}{[AB]}$$

$$\therefore \boxed{K_a[AB] = K_{sp}} \quad \text{—(I)}$$

$$\therefore \boxed{K_{sp} = [A^+][B^-]} \quad \text{—(II)}$$

Eq.(I) & Eq(II) are methemetical statement of 'KSP'

Defination of KSP :- it is methemetical statement It is define as:

" Jonic equilibrium constant ksp is multiplication q concentration q their ions."

$$K_{sp} = K_a \cdot [AB]$$

$$K_{sp} = [A^+][B^-]$$

Ksp. value q is an internal Property q. speringly soluble salt. each speringly soluble salt has Perticular value at 25°C temp and 1 atm. Pressure

Ksp value is very importante. we can determine → solubility q salt

→ common ions effect.

→ selective precipitation....etc.

Ex:1 Calculate solubility q AgCl in a water at 25°C and 1. atm. pressure.

$K_{sp}(AgCl) = 2.8 \times 10^{-10}$ mol/lit

Ans-1 Consider, saturated solution q. AgCl forming an ionic equilibrium such as

$$AgCl + H_2O \rightleftharpoons Ag^+_{(aq)} + Cl^-_{(aq)}$$

	AgCl	H₂O	$Ag^+_{(aq)}$	$Cl^-_{(aq)}$
			0 mol	0 mol
At initial stage	1 mol/lit			
At Equi. stage	1-S	—	S mol/lit	S mol/lit

we know that

$$K_a = \frac{[Ag^+][Cl^-]}{[AgNO_3]} \quad — (1)$$

$\therefore\ K_a[AgCl] = [\cancel{A^+}][\cancel{B^-}]\ [Ag^+][Cl^-]$

we know that $K_a[AgCl] = K_{sp}$

$\therefore\ K_{sp} = [Ag^+][Cl^-]$

$= [s][s]$.

where 's' is solubility of $AgCl$

$\therefore\ s^2 = K_{sp} = 2.8 \times 10^{-10}$ M.

$\therefore\ s = \sqrt{2.8 \times 10^{-10}}$

$s = 1.7 \times 10^{-5}$ mol/lit

$\therefore\ S = 1.7 \times 10^{-5} \times$ Mol. Wt of $AgCl$ (gm/lit)

$= 1.7 \times 10^{-5} \times 143$ gm/lit

$= 243.1 \times 10^{-5}$

$S = 2.431 \times 10^{-3}$ gm/lit

Thus. $AgCl$ is only 2.431×10^{-3} gm soluble in one liter solution [less than 1.0 gm/lit]

<u>Ex = 2</u>

Calculate solubility of sparingly soluble salt $PbSO_4$ in water at 25°c Temp and 1.0 atm pressure [$K_{sp}(PbSO_4) = 1.8 \times 10^{-8}$ M]

Consider saturated aqueous solⁿ of $PbSO_4$ form a following ionic equilibrium such as.

$$PbSO_4 + H_2O \rightleftharpoons Pb^{+2} + SO_4^{-2}$$

At initial state $\quad$ 1 mol/lit $\quad$ 0 mol/lit $\quad$ 0.M.

At equilibrium $\quad$ 1-S $\quad$ [S] Mol/lit $\quad$ [S] Mol/lit

we know that

$$Ka = \frac{[Pb^{+2}][SO_4^{-2}]}{[PbSO_4]}$$

$$\therefore Ka[PbSO_4] = Ksp = [Pb^{+2}][SO_4^{-2}]$$

$$1.8 \times 10^{-8} M = [S][S]$$

$$\therefore [S]^2 = 1.8 \times 10^{-8}$$

$$\therefore S = \sqrt{1.8 \times 10^{-8}} \text{ mol/lit}$$

$$= 1.34 \times 10^{-4} \text{ mol/lit}$$

Thus the only 1.34×10^{-4} mol $PbSO_4$ soluble in a one lit solution it is conform that $PbSO_4$ is speringly soluble salt

EXA = 03

calculate solubility of speringly soluble salt Ag_2CRO_4 in aqueous solution at 25°C Temp and 1. atm. Pressure.

$$Ksp[Ag_2CRO_4] = 1.9 \times 10^{-12} M.$$

Ans-3. → Consider saturated aqueous solution of Ag_2CrO_4 forming a ionic equilibrium such as.

$$Ag_2CrO_4 + H_2O \rightleftharpoons 2Ag^+ + CrO_4^{-2}$$

At initial stage 1 mol/lit 0 mol/lit 0 mol/lit

At equilibrium 1-S 2S + S

$$\therefore Ka = \frac{[2Ag^+]^2 [CrO_4^-]}{[Ag_2CrO_4]}$$

$$\therefore Ka[Ag_2CrO_4] = Ksp = [2Ag^+]^2 [CrO_4^{-2}]$$

$$= [2S]^2 [S]$$

$$= 4S^3$$

$$\therefore 4S^3 = 1.9 \times 10^{-12}$$

$$S^3 = \frac{1.9 \times 10^{-12}}{4} = 0.475 \times 10^{-12}$$

$$\therefore S = [0.475 \times 10^{-12}]^{1/3}$$

$$= 7.8 \times 10^{-5} \text{ mol/lit}$$

Thus only 7.8×10^{-5} mol Ag_2CrO_4 is dissolve in one lit solⁿ.

EXA = 04

Calculate solubility of sperіngly soluble salt CaF_2 in aqueous solⁿ. at 25 °C temp and 1 · Atm pressure [$K_{sp}[CaF_2] = 1.7 \times 10^{-10}$ M

Ans. 4 → Consider saturated aqueous solution of CaF_2 forming a following ionic Equilibrium.

$$H_2O + CaF_2 \rightleftharpoons Ca^{+2} + 2F^-$$

At initial steag	1 mol/lit	0.0 M	0.0 M
At equilibrium steag	1 − S	S +	2S

$$\therefore K_a = \frac{[Ca^{+2}][2F^-]^2}{[CaF_2]}$$

$$\therefore K_a[CaF_2] = K_{sp} = [Ca^{+2}][2F^-]^2$$

$$= [S] \cdot [2S]^2$$

$$\therefore 4S^3 = K_{sp} = 1.7 \times 10^{-10} \text{ M.}$$

$$S^3 = \frac{1.7 \times 10^{-10}}{4}$$

$$\therefore S = [0.425 \times 10^{-10}]^{1/3}$$

$$= 3.4 \times 10^{-4} \text{ mol/lit (M)}$$

EXA :-05 → Calculate solubility q spearingly soluble salt $Mg(OH)_2$ in aqueous solution at 25°C Temp and 1 atm pressure

$$K_{sp}[Mg(OH)_2] = 1.4 \times 10^{-12} M.$$

ANS-05 :- Consider, Saturated aqueous solution q $Mg(OH)_2$ is forming a following ionic equilibrium such as

$$Mg(OH)_2 + H_2O \longrightarrow Mg^{+}_{(aq)} + 2OH^{-}_{(aq)}$$

At initial stage $\quad 1\ M \qquad 0\ M + 0.0\ M$

At equilibrium $\quad 1 - S \qquad S \quad + 2S$

$$\therefore Ka = \frac{[Mg^{+2}][2OH^{-}]}{[Mg(OH)_2]}$$

$$\therefore Ka[Mg(OH)_2] = K_{sp} = [Mg^{+2}][2OH^{-}]$$

$$= [S][2S]^2$$

$$= 4S^3$$

$$\therefore S^3 = \frac{K_{sp}}{4} = 1.4 \times 10^{-12} M$$

$$S = [0.35 \times 10^{-12}]^{1/3}$$

$$S = 1.4 \times 10^{-4} \text{ mol/lit (M)}$$

Thus solubility q $Mg(OH)_2$ is 1.4×10^{-4} mol/lit

COMMON IONS EFFCT

In 1884, French chemist Henrey le-chatelier introduce principle for all system in equilibrium. He stated as " When a stress is applied on a system in equilibrium. The system tends to adjust itself so as to reduce the stress" .

In case of ionic equilibrium, supre-ssion of the dissociation of weak electrolyt [~~backward Reaction~~] done by addition of owns ions.- is called common ions effects.

Mean Presence of common ions is increase rate of reaction of back ward reaction. is called common ion effect.

* Solubility of speringly soluble salt. is always decreses. by effect of common ions present in ionic Equilibrium system. Thus According to le-chatelier principle. the reverce process is fevered to reach the new state of equi. by common ions effect

We can experimentally derive common ion effect by following illustration

EXA ·06

* Calculele solubility q AgCl in pure water and 0.01 M NaCl solution and justfy he-chatelier principle

$$[K_{sp}[AgCl] = 2.8 \times 10^{-10} M$$

Ans-06

Consider, saturated aqueous solution q AgCl form a following ~~Eqli~~ Eqli.

$$H_2O + AgCl_{(s)} \rightleftharpoons Ag^+_{(aq)} + Cl^-_{(aq)}$$

At initial stage	1 M	0	0
At Equi. stage	1 - S	S	S

$$\therefore \text{Ionic Equi } K_a = \frac{[Ag^+][Cl^-]}{[AgCl]}$$

$$\therefore K_a[AgCl] = K_{sp} = [Ag^+][Cl^-]$$

$$\therefore 2.8 \times 10^{-10} M = [S][S]$$

$$\therefore S^2 = 2.8 \times 10^{-10}$$

$$S = (2.8 \times 10^{-10})^{1/2} = 1.7 \times 10^{-5} M \text{ —(1)}$$

* Thus solubility of AgCl in pure wate is $1{\cdot}7 \times 10^{-5}$ M

* solubility of AgCl in 0.1 M NaCl.

We know that NaCl is strong acid·base salt so it is complitely ionized and them soluble in dilute solution

-So id AgCl form a such ionic Equilibrium

At equi	$NaCl_{(s)} \rightleftharpoons$	$Na^+_{(aq)} +$	$Cl^-_{(aq)}$
	0.01	0.01	0.01 M
	$AgCl \rightleftharpoons$	$Ag^+_{(aq)} +$	$Cl^-_{(aq)}$
At initial state	1	0	0
At equi	1-s	s	s

Total = s + [s+0.01]

∴ [s (Cl⁻) is very low] so

$s + 0.01 \cong 0.01$ M

we know

$Ka \cdot [AgCl] = Ksp = [Ag^+][Cl^-]$

$= [s][0.01]$

∴ $[s] = Ksp/0.01$

∴ Solubility of AgCl in 0.01 M NaCl solution is

$$[S] = \frac{Ksp}{0.01} = \frac{2.8 \times 10^{-10}}{0.01}$$

$$= 2.8 \times 10^{-8} \text{ M.}$$

Thus

Solubility of AgCl in pure water 1.7×10^{-5} M $>$ Solubility of AgCl in 0.01 M NaCl is 2.8×10^{-8} M

Thus it is true that due to presence of common ion $[Cl^-]$ the solubility of salt is decrease. This illustration ~~follow~~ ~~justly~~ fi [illegible] justify hachetelier common ions effect.

*** EXA - 07**

Calculate solubility of speringly soluble salt $PbSO_4$ (i) in aqueous solution
(ii) in 0.1 M $Pb(NO_3)_2$ solⁿ
(iii) in 1.0×10^{-3} M Na_2SO_4 solⁿ.
$Ksp[PbSO_4] = 1.8 \times 10^{-8}$ Mol/lit

Ans-07 :- Consider Saturated solution of $PbSO_4$ in pure water, it form a ionic Equilibrium such as {consider solubility of $PbSO_4$ is 's'

$$PbSO_{4(s)} + H_2O \rightleftharpoons Pb^{+2} + SO_4^{-2}$$

At initial stage   1 mol/lit   0. mol/lit   0. mol/

At Equi $1-S$ S S

$$\therefore \quad Ka = \frac{[Pb^{+2}][SO_4^{-2}]}{[PbSO_4]}$$

$$\therefore Ka[PbSO_4] = Ksp = [Pb^{+2}][SO_4^{-2}]$$

$$\therefore 1.8 \times 10^{-8} M = [S][S]$$

$$\therefore S = \sqrt{1.8 \times 10^{-8} M} = 1.34 \times 10^{-4} M$$

Solubility of $PbSO_4$ in water is $\underline{1.34 \times 10^{-4}}$ M.

ii) In 0.1 M $Pb(NO_3)$ solⁿ.

$Pb(NO_3)_2$ is a strong-acid-base salt it is complitely soluble in dilute solⁿ via ionization, $PbSO_4$ add in this solⁿ. Form a Ionic Eq. such as.

$$Pb(NO_3)_2 + H_2O \rightleftharpoons Pb^{+2} + 2NO_3^{-}$$

At Eq.   0.1 M   0.1 M + 0.1 M

$$PbSO_4 \rightleftharpoons Pb^{+2} + SO_4^{-2}$$

At Equi $1-S$ $S + S$

Total conc. of $Pb^{+2} = 0.1 + S \simeq 0.1$

becose 's' of Pb^{+2} is very low.

$$Ka = \frac{[Pb^{+2}][SO_4^{-2}]}{[PbSO_4]}$$

$$\therefore Ka[PbSO_4] = Ksp = [Pb^{+2}][SO_4^{-2}]$$

$$\therefore \quad = [0.1 M][S]$$

$$\therefore S = \frac{Ksp}{0.1} = \frac{1.8 \times 10^{-8} M}{0.1}$$

$$= 1.8 \times 10^{-7} M$$

iii) in 1.0×10^{-3} Na_2SO_4 Sol^n

Na_2SO_4 is strong acid·base salt it is complitely ionized and soluble in ~~ac~~ dilute Sol^n (0 1.0×10^{-3} M)

when ~~PbSO~~ $PbSO_4$ dissolve in it They form a following Ionic Equi

$$H_2O + Na_2SO_4 \rightleftharpoons 2Na^{+}_{(aq)} + SO^{-2}_{4(aq)}$$

$$1.0 \times 10^{-3} \qquad 1.0 \times 10^{-3} M$$

At equi

$$PbSO_4 \rightleftharpoons Pb^{+2} + SO_4^{-2}$$

$$1 - S \qquad S \qquad S$$

Total conc^n $(SO_4^{-2}]$ is $= [S + 1.0 \times 10^{-3} M]$

$$\cong [1.0 \times 10^{-3} M]$$

$\therefore$ ~~S = K~~

$$Ka = \frac{[Pb^{+2}][SO_4^{-2}]}{[PbSO_4]}$$

$K_{sp} = [Pb^{+2}][SO_4^{-2}]$

$= [S][1.0 \times 10^{-3} M]$

$\therefore S = K_{sp} / 1.0 \times 10^{-3} M$

$= \frac{1.8 \times 10^{-8}}{1.0 \times 10^{-3}}$

$S = 1.8 \times 10^{-5}$ Mol/lit

∴ Thus Solublity q $PbSO_4$ in 0.01M Na_2SO_4 is <u>1.8×10^{-5}</u> mol/lit that is less than Pure water. means it is justify la-chetalier & common ions effect principle.

<u>EXA - 08</u> : Calculate ~~soluble~~ solubility q speringly soluble salt CaF_2 in
(I) Pure water (II) in 0.1M $Ca(NO_3)_2$
(III) 0.1 M NaF sol^n and justify la-chetelier principle.

<u>Ans-08</u> :- Consider, saturated aqueous solution q CaF_2 form a following Ionic equilibrium.

$$CaF_2 + H_2O \rightleftharpoons Ca^{+2} + 2F^-$$

At initial Stage	1 mol/lit	0 M	0.0 M.
At Equilibrium Stage	1 - S	S	2S

(where as 'S' is solubility q CaF_2

$$\therefore K_a = \frac{[Ca^{+2}][2F^-]^2}{[CaF_2]}$$

$$\therefore K_a[CaF_2] = K_{sp} = [Ca^{+2}][2F^-]^2$$

$$= [S][2S]^2$$

$$\therefore 4S^3 = K_{sp} = 1.7 \times 10^{-10} M$$

$$\therefore S^3 = \frac{1.7 \times 10^{-10}}{4} = 0.425 \times 10^{-10} M$$

$$= (0.425 \times 10^{-10})^{1/3}$$

$$\therefore S = 3.5 \times 10^{-4} \text{ mol/lit}$$

Thus solubility of CaF_2 in pure water is 3.5 × 10^{-4} mol/lit

* (ii) solubility of CaF_2 in 0.1 M $Ca(NO_3)_2$ solⁿ

we know that $Ca(NO_3)_2$ is strong salt so it is complitely ionized and dissolve in dilute aqueous solⁿ.

~~CaF_2~~ Saturated CaF_2 solution in 0.1 M $Ca(NO_3)_2$ forming following Ionic Equi

$$Ca(NO_3)_2 + H_2O \rightleftharpoons Ca^{+2} + 2(NO_3)^-$$

At eq. 0.1 M 0.1 M + 2(0.1)

$$CaF_2 + H_2O \rightleftharpoons Ca^{+2} + 2F^-$$

At initial stage 1.M 0.0 M 0.0 M

At equi [1-S] [S] + $[2S]^2$

$$Ka = \overset{\text{Total}}{\text{Conc. of}}\ Ca^{+2} = 0.1\,M + S$$

$$\cong 0.1\,M \quad \because S \text{ is very low}$$

$$\therefore Ka = \frac{[Ca^{+2}][2F^-]}{[CaF_2]}$$

$$\therefore Ka[CaF_2] = Ksp = [Ca^{+2}][2F^-]$$

$$= [0.1][2S]^2$$

$$\therefore 4S^2 = \frac{1.7\times10^{-10}}{0.1}$$

$$\therefore S^2 = \frac{1.7\times10^{-10}}{4\times0.1} = 0.425\times10^{\cancel{-10}\,-09}$$

$$\therefore S = 2.05\times10^{-5}\,M.$$

The solubility of CaF_2 is 2.05×10^{-5} M in 0.1 M $Ca(NO_3)_2$ solution that is lower than pure water [3.5×10^{-4} Mol/lit]

(iii) In 0.1 M NaF solⁿ.

NaF is strong salt it is completely ionized and dissolved in dilute 0.1 M solⁿ. Consider CaF_2 ~~[illegible]~~ saturated solution of CaF_2 in 0.1 M NaF forming following ionic Equilibrium.

$$NaF + H_2O \rightleftharpoons Na^+ + F^-$$

At Equil $\qquad 0.1 \qquad 0.1$

$$CaF_2 + H_2O \rightleftharpoons Ca^{+2} + 2F^-$$

At Equil. $\qquad 1-S \qquad S \qquad +2S$

Total conc. of F^- = $2s + 0.1 \cong 0.1\ M$
[Because 's' is very little]

$\therefore K_a = \dfrac{[Ca^{+2}][2F^-]}{[CaF_2]} = K_a[CaF_2] = \cancel{[ca}^{+2}$

$\therefore K_a[CaF_2] = K_{sp} = [Ca^{+2}][2F^-]^2$
$= [S][0.1\ M]^2$

$S = K_{sp} / [0.1]^2$

$= \dfrac{1.7 \times 10^{-10}}{(0.1)^2}$

Solubitity $S = 1.7 \times 10^{-8}$ mol/lit

Thus solubility of CaF_2 in 0.1 M NaF solution is 1.7×10^{-8} M that is very lower than pure water 2.05×10^{-5} M. It is true that the solubility of salt is decreses with presence of common ions.

EXA-09 :- Calculate solubility of speringly soluble salt Ag_2CrO_4 in
(I) Pure water
(II) 0.05 M $AgNO_3$ solⁿ
(III) 0.05 M K_2CrO_4 solⁿ.
$K_{sp}[Ag_2CrO_4] = 1.99 \times 10^{-12}$ M

Ans-10 → Consider saturated solution of Ag_2CrO_4 in pure water forming a following Ionic Equilibrium

$$Ag_2CrO_4 + H_2O \rightleftharpoons 2Ag^{+} + CrO_4^{-2}$$

At initial stage   1   0   0   0

At Equi $1-s$ $2s$ + s

$$\therefore Ka = \frac{[2Ag^{+}]^{2}[CrO_4^{-2}]}{[Ag_2CrO_4]}$$

$$\therefore Ka[Ag_2CrO_4] = Ksp = [2s]^{2}[s]$$

$$= 4s^{3}$$

$$4s^{-3} = 1.99 \times 10^{-12}$$

$$s^{3} = 0.50 \times 10^{-12}$$

$$s = [0.80 \times 10^{-12}]^{1/3}$$

$$= 7.0 \times 10^{-5} \text{ mol/lit}$$

(ii) In 0.05 M $AgNO_3$ solution

We know that $AgNO_3$ is a strong salt, it is completely ionized and dissolve in dilute aqueous solⁿ.

~~Ag~~ Saturated solution of ~~AgN~~ Ag_2CrO_4 in 0.05 M $AgNO_3$ Solⁿ, form a

Such as

$$AgNO_3 + H_2O \rightleftharpoons Ag^+ + NO_3^-$$

At eq. $\quad 0.05\,M \qquad\qquad 0.05\,M + 0.05\,M$

$$Ag_2CrO_4 + H_2O \rightleftharpoons 2Ag^+ + CrO_4^{-2}$$

At Eq. $\quad 1-S \qquad\qquad 2s + s$

Total conc. of $Ag^+ = (0.05 + 2s)\,M$

$$\simeq 0.05\,M$$

$$\therefore\ Ka = \frac{[2Ag^+]^2[CrO_4^{-2}]}{[Ag_2CrO_4]}$$

$$\therefore\ Ka[Ag_2CrO_4] = Ksp = [\cancel{2\cdot}[0.05]]^2[s]$$

$$\therefore\ s = \frac{1.99\times10^{-12}}{[0.05]^2} = 0.8\times10^{-9}\ \text{mole/litr}$$

(III) in 0.05 M K_2CrO_4 soln.

K_2CrO_4 is strong salt. it is completly ionized and dissolve in dilute aqueous soln. consider saturated soln of Ag_2CrO_4 in 0.05 M K_2CrO_4 forming following ionic eqlibrium.

$$\cancel{A}\quad K_2CrO_4 + H_2O \rightleftharpoons 2K^+ + CrO_4^{-2}$$

At equi $\quad 0.05\,M \qquad\qquad 2(0.05\,M) + 0.05\,M$

$$Ag_2CrO_4 \rightleftharpoons 2Ag^+ + CrO_4^{-2}$$

$1-S \qquad\qquad 2s + s$

Total conc. of $CrO_4^{-2} = S + 0.05$

$= 0.05$ mole/lit

$$\therefore K_a = \frac{[2Ag^{+}]^2[CrO_4^{-2}]}{[Ag_2CrO_4]}$$

$$\therefore K_a[Ag_2CrO_4^{-2}] = K_{sp} = [2Ag^{+}]^2[CrO_4^{-2}].$$

$$\therefore K_{sp} = [2S]^2[0.05]$$

$$\therefore 4S^2 = \frac{1.99 \times 10^{-10}}{[0.05] \times 4} = 9.95 \times 10$$

$$S = \left[\frac{1.99 \times 10^{-12}}{4 \times 0.05}\right]^{1/2}$$

$$= 1.0 \times 10^{-6}$$

$$\therefore S^2 = \cancel{\frac{1.99 \times 10^{-12}}{0.2}} = \cancel{1.99 \times 10^{-9}} M$$

$$= 1.99$$

Thus solubility of Ag_2CrO_4 in 0.05 M K_2CrO_4 aqueous solution is $\frac{\cancel{1.99} \times 10^{-9}}{1.0 \times 10^{-6} \text{ mol/lit}}$ that is lower that pure water