# Ionic equlibrium, Buffer solution, pH and Henderson Equation, Common Ion effect, Solubility and Solubility Product

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## Electrolytes

Substances like acid, base and salts which can conduct electricity in their aqueous solution due to ionization are known as **electrolytes**. The presence of ions in solution is responsible for the conduction and when the current passes through the wires, it shows the movement of free electrons through wire. This process of dissociation and flow of ions in aqueous solution of electrolyte is known as **electrolysis**.

**Example:** the electrolysis of sodium chloride in its aqueous solution gives sodium and chloride ion in solution.



There are many electrolytes present in the human body such as sodium, magnesium, calcium, potassium, chloride ion and bicarbonate ions. These electrolytes involve in various functions of our cells and our organs.



## Examples of Electrolytes

Hydrochloric acid (EICl)
Nitria acid (EECl)
Sulfuric acid (EECl)
Hydroindia acid (EECl)
Hydroindia acid (HClC)
Perchloric acid (ECCl)
Acetic acid (ECCC)
Acetic acid (EECC)
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Rubidiure Hydroxide (RhOH) Cesium Hydroxide (C4OH) Calcium Hydroxide (Ca(OH)<sub>2</sub>) Scrontiure Hydroxide (Sr(OH)<sub>2</sub>) Barium Hydroxide (Sa(OH)<sub>2</sub>)

## Strong Electrolytes

Some electrolytes such as potassium chloride, sodium hydroxide, sodium nitrate are completely ionized in their constituent ions in solution. These electrolytes are known as **strong electrolytes**. In other words strong electrolytes are 100% dissociated in their aqueous solution.

For example,

KCl(aq) → K(aq) + Cl(aq) NH,NO<sub>1ai</sub> → NH;(aq) + NO;(aq) HNO<sub>3ai</sub> → H·(aq) + NO;(aq) NaOH(aq) → Na'(aq) + OH(aq)



Strong electrolytes like salts are composed of oppositely charged ions. In solid state, these ions are held by strong electrostatic forces of attractions. When theses electrolytes are dissolved in water, the attraction forces between ions are highly weakened due to high dielectric constant of water that is 80.

## LIST OF STRONG ELECTROLYTES

Chloric acid	HC10
Sulphurous acid	H2\$O3
Hydriodic acid HI	HI
Hydrobromic acid HBr	HBr
Hydrochloric acid HCl	HC1
Sulphurie acid	H2\$04
Nitrie acid	HNO <sub>1</sub>

# Weak Electrolytes

Electrolytes which are weakly ionized in their aqueous solution are called as **weak electrolytes**. In the aqueous solution of weak electrolytes, the constituent ions are in equilibrium with un-dissociated molecules of electrolytes.

Due to partial dissociation of electrolytes, the number of ions in a given volume of the solution is almost remains same for a large change in the concentration.

Example:

CH,COOH + H,O + H,O + CH,COO

Such type of equilibrium exists between ions and unionized molecule is called ionic 35 equilibrium. The fraction of molecules dissociates can be represents by degree using of dissociation. Generally weak acids and bases are good examples of weak electrolytes.



## LIST OF WEAK ELECTROLYTES

	Weak electrobre	Chemical formula
1	Acebic acid	CH/COOH
2.	Hydrocyanic acid	HCN
3.	Ammonium hydroxide	NHJOH
- 4.	Ammonia	NH
\$.	Hydrofluoric acid	HF
б.	Carbonie acid	H,CO;
7.	Mercuric chloride	HgCl <sub>1</sub>
8.	Chromic acid	H.C.O.
9.	Borie acid	H,BO,
10.	Nitrous acid	HNO2

# DISSOCIATION Vs. IONIZATION



# Dissociation

- Dissociation is the breaking up of a compound into simpler constituent that are usually capable of recombining under other conditions.
- Dissociation is usually Reversible.

 $\begin{array}{c} AB \rightarrow A + B \\ H_2O \rightarrow H^+ + OH \end{array}$ 



- Ionization is any Process by which electrically neutral atoms or molecules are converted to electrically charged atoms or molecules ions.
- Ionization is usually Irreversible.

 $HCI + H_2O \rightarrow H_3O^+ + CI^-$ 

## Difference

 Dissociation is Break Down of a compound (Ionic Compound)

• But Ionization is the formation of charged particles by the gain or loss of electron

## BUFFER SOLUTION

A buffer solution is one which maintains its pH fairly constant even upon the addition of small amounts of acid or base.

There are two types of buffer solution.

 Acidic Buffer

2. Basic Buffer

## Acidic Buffer:

A weak acid together with a salt of the same acid with a strong base. These are called Acidic Buffer.

For Example.

## CH3COOH CH3COONa.

## **Basic Buffers:**

A weak base and its salt with a strong acid. These are called Basic Buffers.

## For Example. NH4OH+NH4CL

Let us illustrate buffer action by taking example of a common buffer system consisting of solution of acetic acid and sodium acetate.

CH3COOH  $\rightleftharpoons$  H<sup>+</sup>+ CH3COO<sup>-</sup> CH3COONa  $\rightarrow$  Na<sup>+</sup> + CH3COO<sup>-</sup>

Since the salt is comletely ionized, it provides the common ion CH3COO in excess.

# The Buffer solution has a higher pH than the acid itself;



## The pH of Buffer changes only slightly upon addition of an acid or base;



## HOW A BUFFER OPERATES?

We have already stated that a buffer solution containing equimolar amounts (0.10M) of acetic acid and sodium acetate has pH 4.74.Now we proceed to discuss how the addition of a small amount of HC1 or NaOH to the buffer solution affects its Ph.

The pH of the buffer is governed by the equillibrium

CH3COOH ≓ CH3COO + H

The buffer solution has a large excess of CH3COO ions produced by complete ionization of sodium acetate,

#### CH3COONa CH3COO+Na Addition of HCl:

Upon the addition of HCl, the increase of H ions is counteracted by association with the excess of acetate ions to form unionised CH3COOH. Thus the added H ions are neutralised and the Ph of the buffer solution remains virtually unchanged. However owing to the increased concentration of CH3COOH, the equilibrium shifts slightly to the right to increase H ions. This explains the marginal increase of Ph of the buffer solution of HCl.





## Addition of NaOH:

When NaOH is added to the buffer solution, the additional OH ions combine with H ions of the buffer to form water molecules. As a result the equillibrium shifts to the right to produce more and more H ions till practically all the excess OH ions are neutralized and the original buffer Ph restored. However, a new equilibrium system is set up in which CH3COOH is lower than it was in the original buffer. Consequently H is also slightly less and pH slightly higher than the buffer pH values

Operation of a Basic buffer as NH4OH NH4Cl can also be explained on the same lines as of an acid buffer upon addition of HCl the H ions combine with OH ions of the buffer to form water molecules.The equillibrium,



#### NH4OH NH4 OH

Is shifted to the right till all the additional H ions are neutralized and the original buffer Ph restored When NaOH is added to the buffer solution,OH ions associate with excess of NH4 ions to form unassociated NH4OH. Thus the Ph of the buffer is maintained approximately constant.



## BUFFER CAPACITY:

It can be defined as the ability of a buffer to resist the change of potential of hydrogen. Mathematically it can be expressed as;

How it can be determined?

Buffer capacity of a buffer solution is determined by the size of actual molarities.

How do you calculate the Buffer Capacity? Buffer Capacity refers to the maximum amount of a strong acid or a strong base that is added so that a significant change in pH occur.

## **IMPORTANCE:**

It is important in chemical as well as in biological term.

## **Buffer Capacity of water:**

As the purity of fresh water increases, the buffer capacity decreases and 100 percent pure fresh water is no buffer at all.

What factors determined buffer capacity ?

The concentration of the buffer (the higher the concentration, the larger the buffering capacity) and how close the pka of buffer is compared to the pH of solution.



Water has a slight propensity to dissociate into hydroxide ions and protons.

• The acidity of the aqueous solution is usually reported using the logarithmic p<sup>H</sup> scale.

Buffers usually maintain the p<sup>H</sup> of the extracellular fluid between 7.35-7.45.

PH values in the cell and in the extracellular fluid are kept constant within narrow limits.

- In the blood, the p<sup>H</sup> value normally ranges only between 7.35 and 7.45.
- The p<sup>H</sup> value of cytoplasm is slightly lower than that of blood at 7.0–7.3. In lysosomes pH 4.5– 5.5
- Extreme values are found in the stomach p<sup>H</sup> 2 and in the small bowel > 8.
- Since the kidney can excrete either acids or bases, depending on the state of the metabolism, the p<sup>H</sup> of urine has a particularly wide range of variation 4.8–7.5.









#### • **PROTONATION** IS THE ADDITION OF A PROTON (H<sup>+</sup>) TO AN ATOM, MOLECULE, OR ION.

 DEPROTONATION IS THE REMOVAL OF A PROTON (H<sup>+</sup>) FROM A MOLECULE, FORMING THE CONJUGATE BASE.

#### THE RELATIVE ABILITY OF A MOLECULE TO GIVE UP A PROTON IS MEASURED BY ITS PK<sub>A</sub> VALUE.

 A LOW Pk<sub>a</sub> VALUE INDICATES THAT THE COMPOUND IS ACIDIC AND WILL EASILY GIVE UP ITS PROTON TO A BASE.

- AT PH VALUES LESS(ACIDIC) THAN PKa, PROTONATED ACID FORM IS MORE (IT WILL GIVE OUT ITS H+ IONS)
- AT PH VALUES(BASIC) GREATER THAN THE PKa, DEPROTONATED BASE FORM IS MORE IN THE SOLUTION (IT WILL TAKE H<sup>+</sup> IONS TO NEUTRALIZE THE PH AND BRING IT DOWN).

#### WHY IS PH IMPORTANT IN BIOLOGY?

- PH AFFECTS SOLUBILITY OF MANY SUBSTANCES.
- P<sup>H</sup> AFFECTS STRUCTURE AND FUNCTION OF MOST PROTEINS INCLUDING ENZYMES.
- MANY CELLS AND ORGANISMS (ESP. PLANTS AND AQUATIC ANIMALS) CAN ONLY SURVIVE IN A SPECIFIC P<sup>H</sup> ENVIRONMENT.

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- IMPORTANT POINT -
  - PH IS DEPENDENT UPON TEMPERATURE







# THE QUANTITATIVE RELATIONSHIP BETWEEN pH and Pka IS DESCRIBED BY HENDERSON-HASSELBALCH EQUATION

HASSELBALCH EQUATION pH = pKa + log [SALT]/[ACID] pOH = pKb + log[SALT]/[BASE] THE EQUATION CAN BE DERIVE AS FOLLOWS -



- MULTIPLY EACH SIDE WITH -log
   -log[H+] =-logKa -log[HA] /[A-}
   i.e, pH = pKa + log[A-]/[HA]
- Due to common ion effect , the concentration of A- is because of salt. So we can write
- pH = pKa + log[salt]/[acid] similarly
- pOH = pKb + log[salt]/[base]

#### IMPORTANCE

- MEASUREMENT OF P<sup>H</sup> IS ONE OF THE MOST IMPORTANT AND FREQUENTLY USED PROCEDURES IN BIOCHEMISTRY. THE P<sup>H</sup> AFFECTS THE STRUCTURE AND ACTIVITY OF BIOLOGICAL MACROMOLECULES;
- 1. FOR EXAMPLE, THE CATALYTIC ACTIVITY OF ENZYMES IS STRONGLY DEPENDENT ON P<sup>H</sup>
- MEASUREMENTS OF THE PH OF BLOOD AND URINE ARE COMMONLY USED IN MEDICAL DIAGNOSES.

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## Solubility

 The solubility of a substance is the amount of that substance that will dissolve in a given amount of solvent. "Solubility" may be considered to be an equilibrium; the equilibrium is between solid and ions in solution. Any ionic solid is 100% ionized in aqueous solution; once it actually dissolves.



Na⁺(aq) + Cl⁻(aq) t↓ NaCl(s)

#### Factors that affect solubility

- Temperature
  - Solubility generally increases with temperature.
- Common ion effect
  - Common ions reduce solubility
- pH of solution
  - pH affects the solubility of ionic compounds.
- Formation of complex ion
  - The formation of complex ion increases solubility

#### Solubility product

In general, solubility product, K<sub>sp</sub>, is the mathematical product of its dissolved ion concentrations raised to the power of their stoichiometric coefficients. This statement is called the solubility product principle



Solubility product constant Molar solubility of the ions

#### **Solubility product**

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Solubility product constant (K<sub>s</sub>)
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Most salts dissociate into ions when they dissolve. For example:

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BaSO_4(s) \Rightarrow Ba^{2*}(aq) + SO_4^{2*}(aq)
```

This equilibrium system may be described by the mass-action expression

 $K_{sp} = [Ba^{2*}][SO_4^{2*}]$ 

$$\mathsf{BaSO}_{4(s)} \rightleftharpoons \mathsf{Ba}^{2+}_{(aq)} + \mathsf{SO}_{4}^{2-}_{(aq)} \qquad \mathsf{Ba}^{2+} \begin{bmatrix} 0 \\ 0 \\ 0 \\ 0 \end{bmatrix}^{2-}_{0}$$

 The Ba<sup>2+</sup> concentration and the SO<sup>2-</sup><sub>4</sub> concentration are equal since for each BaSO<sub>4</sub> unit that dissolves one Ba<sup>2+</sup> and one SO<sup>2-</sup><sub>4</sub> ion form.

**Solubility = s = [Ba^{2+}] = [SO\_4^{2-}]** 

#### From the above equation we know that

 $K_{sp} = [Ba^{2+}] [SO_{4}^{2-}]$  $K_{sp} = s \times s = s^{2}$  $K_{sp} = [Ba^{2+}] [SO_{4}^{2-}]$ 



solubility product

### Problems on solubility product

 Calculate the K<sub>s</sub> value of BaSO<sub>4</sub> which has a solubility of 3.9×10<sup>-5</sup> mol/L at 25°C.

```
BaSO_{4(s)} \rightleftharpoons Ba^{2+}_{(aq)} + SO_{4}^{2-}_{(aq)}
K_{sp} = [Ba^{2+}] [SO_{4}^{2-}]
K_{sp} = s \times s = s^{2}
= 3.9 \times 10^{-5} \times 3.9 \times 10^{-5}
= 1.52 \times 10^{-9}
```

#### **Solubility and Solubility Products**

(1) What is the solubility of AgCl if the  $K_{sp}$  is 1.6 x 10<sup>40</sup>

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\operatorname{AgCl}(s) \rightleftharpoons \operatorname{Ag}^{+}(aq) + \operatorname{Cl}^{+}(aq)
K_{\infty} = [\operatorname{Ag}^{+}][\operatorname{Cl}^{+}]
```

If *s* is the solubility of AgCl, then:

 $[Ag^{+}] = s \text{ and } [Cl^{-}] = s$   $K_{\varphi} = (s)(s) = s^{2} = 1.6 \text{ x } 10^{-10}$   $s = \sqrt{1.6 \text{ x } 10^{-10}}$  $s = 1.3 \text{ x } 10^{-5} \text{ mol/L}$ 



#### **Different types of solution**

Unsaturated solution: More solute can be dissolved in it.

Saturated solution: No more solute can be dissolved in it. Any more of solute you add will not dissolve. It will precipitate out.

Super saturated solution: Has more solute than can be dissolved in it. The solute precipitates out.

#### SOLUBILITY EQUILIBRIUM

- Solubility refers to the ability of a substance to dissolve. In the study of solubility equilibrium we generally deal with low solubility compounds (those which dissolve poorly)
- Low solubility compounds are classified as precipitates. A set of solubility rules indicate which combination of ions generally form precipitates. These rules allow us to identify low solubility substances and potential precipitates. The concentrations of the ions present however are the final determiners of precipitate formation.

#### SOLUBILITY RULES

- (1) Inorganic acids (those not consisting of carbon, hydrogen and oxygen) are soluble. Low molecular weight organic acids are soluble.
- (2) Compound containing alkali metals and ammonium (NH<sub>4</sub><sup>+</sup>) cations are soluble.
- (3) Compounds containing nitrates (NO<sub>3</sub>-), acetates (C<sub>2</sub>H<sub>3</sub>O<sub>2</sub><sup>-</sup>) and perchlorates (ClO<sub>4</sub><sup>-</sup>) are soluble.
- (4) Compounds containing chlorides (Cl<sup>-</sup>) are soluble except for AgCl, PbCl<sub>2</sub> and Hg<sub>2</sub>Cl<sub>2</sub>. Compounds containing bromides (Br<sup>-</sup>) and iodides (I<sup>-</sup>) have solubility similar to chlorides.
- (5) Sulfate containing compounds are soluble except for PbSO<sub>4</sub> and BaSO<sub>4</sub>. CaSO<sub>4</sub> and Ag<sub>2</sub>SO<sub>4</sub> are slightly soluble.
- (6) Hydroxides are insoluble (precipitates) except Alkali Metal (I) and heavier Alkaline Earth Metal (II) hydroxides after Mg.
- (7) Carbonates (CO<sub>3</sub>-<sup>2</sup>), phosphates (PO<sub>4</sub>-<sup>3</sup>) and arsenates (AsO<sub>4</sub>-<sup>3</sup>) are insoluble except for Alkali Metal (I) and Alkaline Earth Metal (II) compounds
- (8) Sulfides (S<sup>-2</sup>) are insoluble except for Alkali Metal (I) and Alkaline Earth Metal (II) compounds.

#### **Ksp AND SOLUBILITY**

- Since solubility is an equilibrium phenomena, equilibrium concepts can readily be applied.
- If the system is at equilibrium, the solution must be saturated (no more solid can be dissolved). Equilibrium concepts cannot be applied to unsaturated solutions (more solid can still be dissolved) since equilibrium has not yet been reached.
- · For a typical saturated solution of silver chloride,

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

the equilibrium expression is:

 $K = [Ag^+] x [Cl^-] / 1 \text{ or simply } [Ag^+] x [Cl^-]$ 

Since AgCl is and solid it is not included in the equilibrium expression!

• This K value is called the K<sub>sp</sub>. "sp" stands for solubility product. The product of the concentrations of the soluble ions.

	Salt	Ions of Salt	K <sub>sp</sub> (25 "C)
	CaF <sub>2</sub>	$\implies$ Ca <sup>2+</sup> + 2F <sup>-</sup>	$3.9 \times 10^{-11}$
	PbF <sub>2</sub>	$= Pb^{2+} + 2F^{-}$	$3.6 \times 10^{-8}$
	AgCl	= Ag <sup>+</sup> + Cl <sup>-</sup>	$1.8 \times 10^{-10}$
and the second	AgBr	= Ag <sup>+</sup> + Br <sup>-</sup>	$5.0 \times 10^{-13}$
SOLUPH ITY	AgI	$\Rightarrow$ Ag <sup>+</sup> + I <sup>-</sup>	$8.3 \times 10^{-17}$
SOLUBILITY	PbCl <sub>2</sub>	$\implies Pb^{2+} + 2Cl^{-}$	$1.7 \times 10^{-5}$
PRODUCTS	PbBr <sub>9</sub>	$\implies Pb^{2+} + 2Br^{-}$	$2.1 \times 10^{-6}$
IRODUCIS	PbI,	$\implies Pb^{2+} + 21^{-}$	$7.9  imes 10^{-9}$
OF			
	Al(OH) <sub>8</sub>	=== Al3+ + 3OH-	$3 \times 10^{-34}$ (a)
SOME	Ca(OH) <sub>2</sub>	$= Ca^{2+} + 2OH^{-}$	$6.5  imes 10^{-6}$
COMPACIN	Fe(OH) <sub>2</sub>	==== Fe <sup>2+</sup> + 2OH <sup>-</sup>	$7.9  imes 10^{-16}$
COMMON	Fe(OH)3	=== Fe <sup>3+</sup> + 3OH-	$1.6  imes 10^{-39}$
SALTS	Mg(OH) <sub>2</sub>	$\implies$ Mg <sup>2+</sup> + 2OH <sup>-</sup>	$7.1  imes 10^{-12}$
SALIS	Zn(OH) <sub>2</sub>	$\implies Zn^{2+} + 2OH^{-}$	$3.0 \times 10^{-16}$ (b)
AT	ALC:		
	Ag <sub>2</sub> CO <sub>3</sub>	$= 2Ag^{+} + CO_{3}^{2-}$	$8.1  imes 10^{-12}$
25 ° C	CaCO <sub>3</sub>	$= Ca^{2+} + CO_3^{2-}$	$4.5 \times 10^{-9}$ (c)
	SrCO <sub>8</sub>	$\implies$ Sr <sup>2+</sup> + CO <sub>3</sub> <sup>2-</sup>	$9.3 \times 10^{-10}$
	BaCO <sub>3</sub>	$= Ba^{2+} + CO_a^{2-}$	$5.0  imes 10^{-9}$
	CoCO <sub>3</sub>	$\Longrightarrow$ Co <sup>2+</sup> + CO <sub>3</sub> <sup>2-</sup>	$1.0 \times 10^{-10}$
	NiCO <sub>3</sub>	$\implies$ Ni <sup><math>a+</math></sup> + CO <sub>a</sub> <sup><math>a-</math></sup>	$1.3 \times 10^{-7}$
	ZnCO <sub>3</sub>	$\implies$ Zn <sup>2+</sup> + CO <sub>3</sub> <sup>2-</sup>	$1.0  imes 10^{-10}$
	Ag <sub>2</sub> CrO <sub>4</sub>	$\implies 2Ag^+ + CrO_4^{2-}$	$1.2 \times 10^{-12}$
	PbCrO <sub>4</sub>	$\implies Pb^{2+} + CrO_4^{2-}$	$1.8  imes 10^{-14}$ (d)
	CaSO <sub>4</sub>	$\implies$ Ca <sup>2+</sup> + SO <sub>4</sub> <sup>2-</sup>	$2.4 \times 10^{-5}$
	SrSO.	$\implies Sr^{2+} + SO_4^{2-}$	$3.2 \times 10^{-7}$ _
and the second	BaSO.	$= Ba^{2+} + SO_4^{2-}$	$1.1 \times 10^{-10}$
	PbSO4	$\implies Pb^{2+} + SO_*^{2-}$	$6.3 \times 10^{-7}$

#### **THE COMMON ION EFFECT**

- The common ion effect is an application of LeChatelier's Principle. Recall that equilibrium shifts are caused be stresses applied to the system.
- If additional products are added to an equilibrium system that system will shift towards the reactants. In a solubility equilibrium adding one of the dissolved ion components will cause a shift left and the solubility of the compound will decrease.
- Reducing the concentration of one of the dissolved ions will cause a shift to the right thereby allowing the compound to dissolve to a greater extend.

# THUS THE SOLUBILITY OF AgCI DECREASED DUE TO COMMON CI- ION AND SOME AgCI GET PRECIPATED OUT.



#### THE COMMON ION EFFECT

- Problem: Calculate the solubility of silver chloride in pure water and in a 0.10 M NaCl solution. Compare the results.
- Solution:  $AgCl_{(s)} \leftrightarrow Ag^{+}_{(aq)} + Cl^{-}_{(aq)}$
- $Ksp = [Ag^+] \times [Cl^-] = 1.8 \times 10^{-10}$
- Each dissolved AgCl gives one Ag<sup>+</sup> and one Cl<sup>-</sup> ion therefore if [Ag<sup>+</sup>] = X, [Cl<sup>-</sup>] = X
- Ksp =  $X^2 = 1.8 \times 10^{-10}$ , X =  $(1.8 \times 10^{-10})^{1/2} = 1.3 \times 10^{-5} M$
- In 0.10 M NaCl the concentration of Cl<sup>-</sup> = 0.10 M and
- Ksp =  $[Ag^+]$  (0.10)\* = 1.8 x 10<sup>-10</sup>,  $[Ag^+]$  = 1.8 x 10<sup>-10</sup> / 0.10
- ]Ag<sup>+</sup>] = 1.8 x 10<sup>-9</sup> M (the presence of the common Cl<sup>-</sup> ion in solution has shifted the equilibrium hard to the left and made the silver chloride much less soluble than in pure water
- \*Note: the Cl<sup>-</sup> from the AgCl in the common ion system is so small as compared to the 0.10 M Cl<sup>-</sup> from the NaCl that is was neglected.

## WILL A PRECIPITATE FORM?



- In order to decide whether a precipitate will form when two solutions are mixed we must first calculate the value of Q (reaction quotient). Q is calculated with the same format as Ksp (an ion concentration product)
- If the value of Q exceeds the Ksp value, a precipitate will form.
- If the value of Q is less than Ksp, a precipitate will not be formed.
- If Q = Ksp then the solution is a saturated solution.

## WILL A PRECIPITATE FORM?

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# WILL A PRECIPITATE FORM?

- Problem: If 100 ml of 0.0001 M AgNO<sub>3</sub> is mixed with 200 ml of 0.0002 M NaCl, will a precipitate form?
- Solution: The potential precipitate is AgCl. (See the solubility rules).
- Q = [Ag<sup>+</sup>] x [Cl<sup>-</sup>], when the two solutions are mixed, the concentrations of both the Ag<sup>+</sup> ion and the Cl<sup>-</sup> ion will change due to dilution of one solution by the other.
- $M_1V_1 = M_2V_2$  (the dilution formula)
- For  $Ag^+$ , 0.0001 (100) =  $[Ag^+]$  (100 + 200)
- [Ag<sup>+</sup>] = 3.33 x 10<sup>-5</sup> M
- For  $Cl^2$ ,  $0.0002(100) = [Cl^2](100 + 200)$
- [Cl<sup>-</sup>] = 6.67 x 10<sup>-5</sup> M
- $Q = (3.33 \text{ x } 10^{-5}) (6.67 \text{ x } 10^{-5}) = 2.22 \text{ x } 10^{-9}$ , Ksp = 1.8 x 10<sup>-10</sup>
- Q > Ksp therefore a precipitate will form