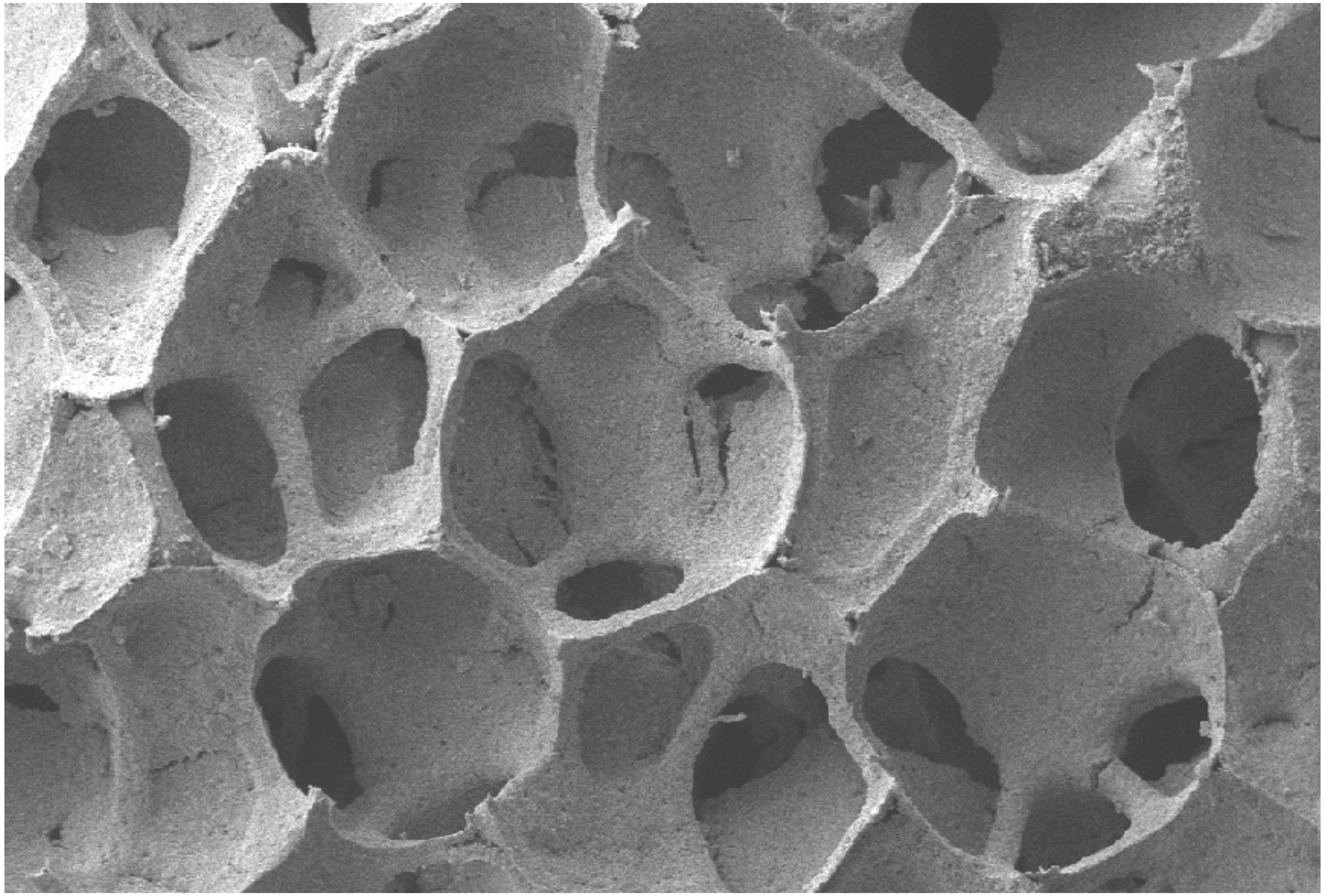
Salt solubility and heterogeneous equilibria



Hydroxyapatite : $Ca_5(PO_4)_3OH$

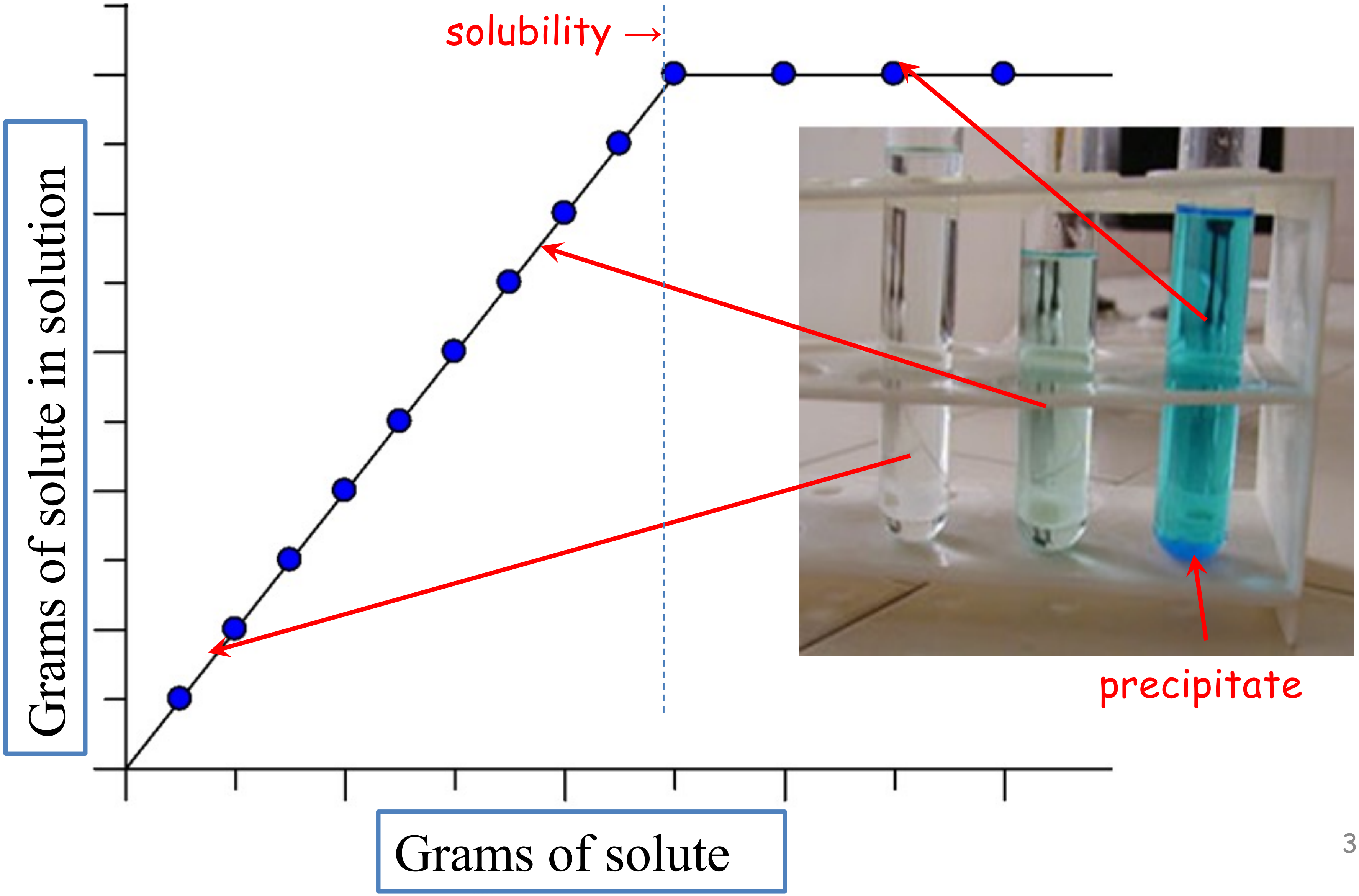
Alcune immagine sono state prese e modificate da “Chimica” di Kotz, Treichel & Weaver, Edises 2007, III edizione

- The solubility of a substance in a solvent is determined by the balance of intermolecular forces between solvent and solutes and by the change in entropy that accompanies solvation.

- Factors like temperature and pressure can alter this balance, varying solubility.

$$CuSO_4\,(s) \rightleftarrows Cu^{2+}\,(aq) + SO_4^{2-}\,(aq)$$

Unsaturated Solution ← → Saturated Solution



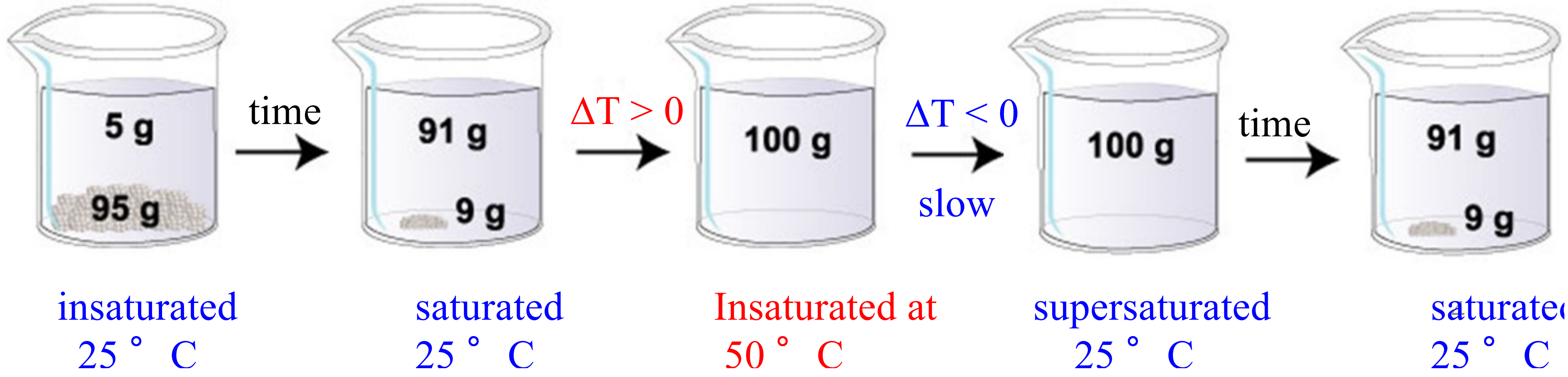
Solubility

It is defined as the maximum amount of solute that dissolves in a specified amount of solvent at a given temperature (and pressure for gases)

Saturated if, in a certain amount of solvent, it is not possible to dissolve further solute (each subsequent addition of solute forms a precipitate)

Unsaturated if the amount of solute dissolved is less than the amount needed to have a saturated solution

Supersaturated if the dissolved solute quantity is greater than the amount needed for a saturated one. Supersaturated solutions are unstable and tend to to reach the state of saturated solutions by precipitaton.



The solubility of a certain solute in a certain solvent depends not only on the characteristics of the two substances, but also on temperature and pressure.

Tahiti

Patagonia

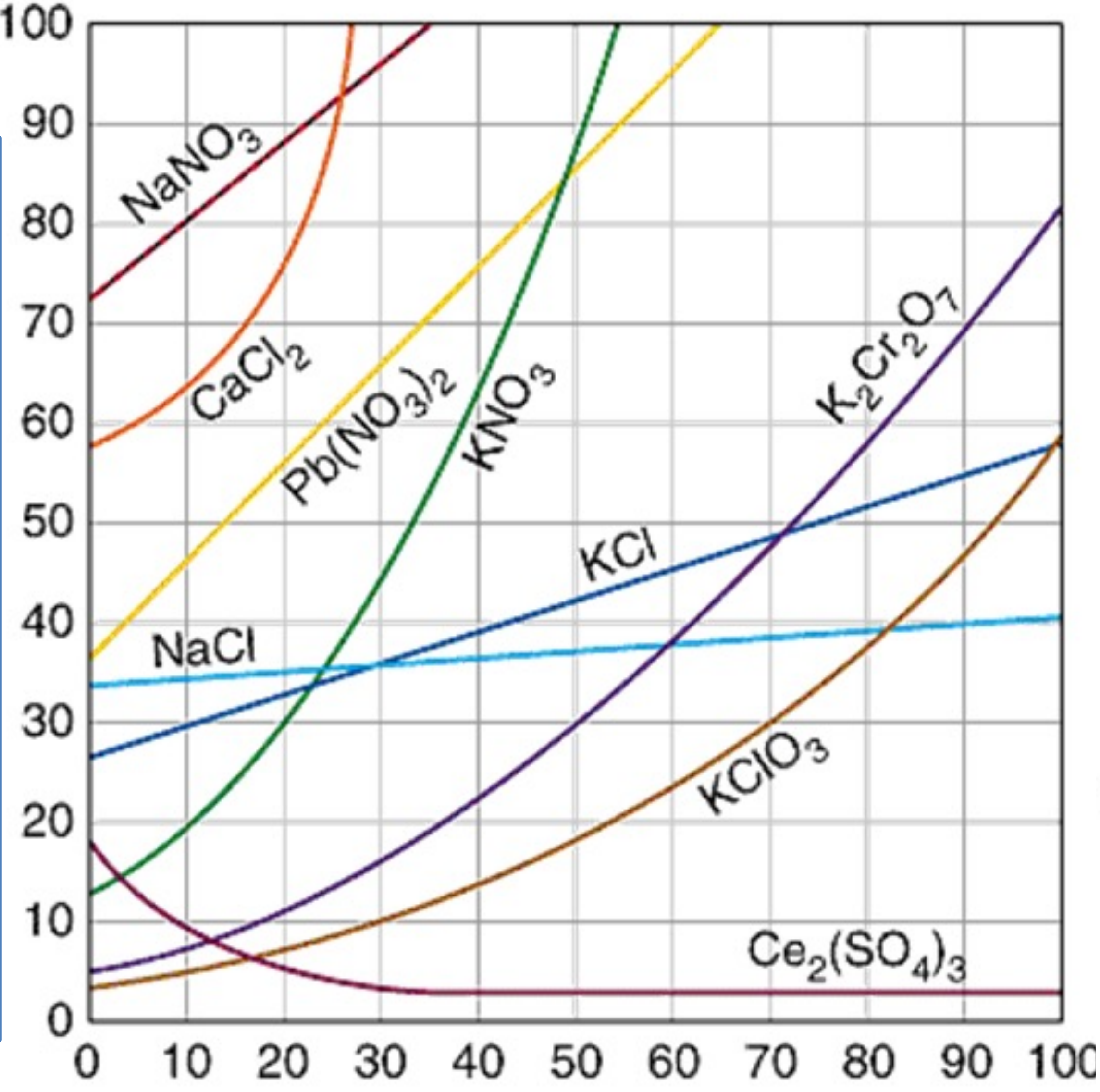
Generally, an increase in temperature increases the solubility of solid substances, while that of the gaseous substances decreases.

Eg tropical seas, warmer, are more "salty" than glacial ones.

The dead sea, at 480 m below sea level

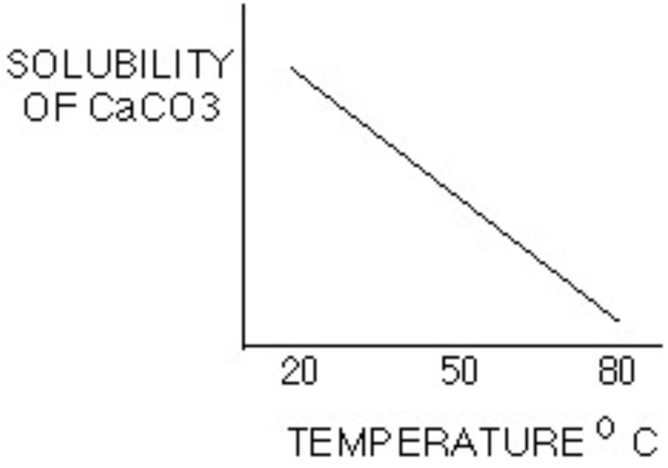
Dead sea: [salt] = 340 g/L (salts of Mg^{2+}, Na^{+}, Ca^{2+}, K^{+}, e Cl^{-} e Br^{-}).
On average in the sea: [NaCl] =35 g/L

Solubility of some salts as a function of temperature at 1 Atm



Solubility (g of salt in 100 ml of water

Temperature ° C



Solubilty of salts

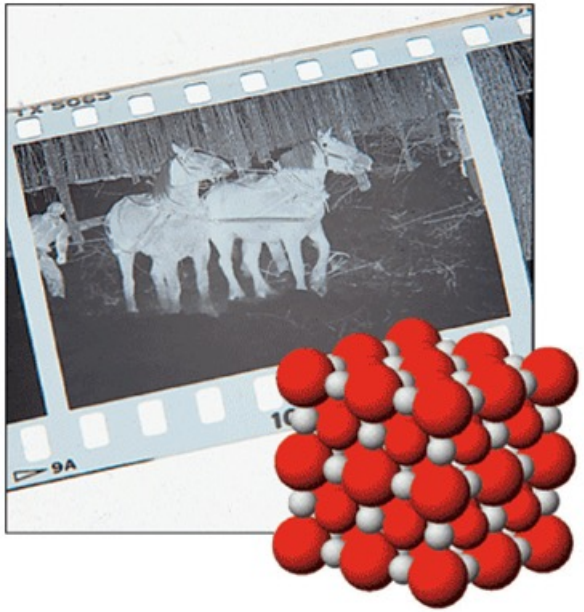
Precipitation reactions are exchange reactions in which one of the products is an insoluble compound in water.

$$CaCl_2\ (aq) + Na_2CO_3\ (aq) \rightleftarrows CaCO_3\ (s) + 2\ NaCl\ (aq)$$

$$AgNO_3\ (aq) + HBr\ (aq) \rightleftarrows AgBr\ (s) + H_3O^+\ (aq) + NO_3^-\ (aq)$$

The solubility of $CaCO_3$ is ~58 mM a 25 ° C and for AgBr it is ~ 0.7 mM at 25 ° C.

Stalactites $CaCO_3$



A b/W film is covered with Ag Br, insoluble in water

The Solubility Product K_{sp}

If we add AgBr to water, a sall amount is dissolved and it establishes this equilibrium:

$$AgBr\ (s) \rightleftarrows Ag^+\ (aq) + Br^-\ (aq)$$

The concentrations of are Ag^+ e Br^- in water are:
$[Ag^+] = [Br^-] = 7.35 \times 10^{-7}$ M at 25 ° C.

The amount of an scarcely soluble salt in water can be expressed in terms of a constant of equilibrium for the solubilization process:

$$K = \frac{[Ag^+]\cdot[Br^-]}{[AgBr]_{solid}} \rightarrow K\cdot[AgBr]_{solid} = [Ag^+]\cdot[Br^-] = K_{PS}$$

$$\mathrm{K_{PS}} = [\mathrm{Ag}^+]\cdot[\mathrm{Br}^-] = 5.4\cdot 10^{-13}\ \mathrm{M}^2$$

The value of the equilibrium constant reflects the solubility of a compound and it is defined as solubility product:

For a given salt, C_xA_y (C = cation and A = anion) the solubilty product has this form:

$$C_xA_y\,(s) \rightleftarrows x\,C^{y+}\,(aq) + y\,A^{x-}\,(aq)$$

$$\boxed{K_{PS} = [C^{y+}]^x \cdot [A^{x-}]^y}$$

dimensions $M^{(x+y)}$

Esemples:

$CaF_2\,(s) \rightleftarrows Ca^{2+}\,(aq) + 2\,F^-\,(aq)$ (x=1 e y=2)

$K_{PS} = [Ca^{2+}]\,[F^-]^2 = 5.3\times10^{-11}\ M^3$

$Ag_2SO_4\,(s) \rightleftarrows 2\,Ag^+\,(aq) + SO_4^{2-}\,(aq)$ (x=2 e y=1)

$K_{PS} = [Ag^+]^2\,[SO_4^{2-}] = 1.2\times10^{-5}\ M^3$

$Ca_3(PO_4)_2\,(s) \rightleftarrows 3\,Ca^{2+}\,(aq) + 2\,PO_4^{3-}\,(aq)$ (x=3 e y=2)

$K_{PS} = [Ca^{2+}]^3[PO_4^{3-}]^2 = 2.0\times10^{-29}\ M^5$

NB: K_{PS} is different from the solubility

If C_xA_y achieves the following equilibrium

$$C_xA_y\,(s) \rightleftarrows x\,C^{y+}\,(aq) + y\,A^{x-}\,(aq)$$

If for S moles of C_xA_y we have xS moles dof C^{y+} and yS moles of A^{x-} then:

$$[C^{y+}] = xS \text{ e } [A^{x-}] = yS$$

Substituting these 2 equations in the formula for the solubility product:

$$K_{PS} = (x \cdot S)^x \cdot (y \cdot S)^y = x^x \cdot y^y \cdot S^{x+y}$$

From which we deduce:

$$\boxed{S = \left(\frac{K_{PS}}{x^x \cdot y^y} \right)^{\frac{1}{x+y}}}$$

The relationship between solubility and the solubility priduct depends on the stoichiometry of the compound

Examples:

$CaF_2\ (s) \rightleftarrows Ca^{2+}\ (aq) + 2\ F^-\ (aq)$ (x=1 e y=2)

$K_{PS} = [Ca^{2+}]\cdot[F^-]^2 = 5.3\cdot10^{-11}\ M^3$

$$S = \left(\frac{5.3\cdot10^{-11}}{1^1\cdot2^2}\right)^{\frac{1}{1+2}} = \left(\frac{5.3\cdot10^{-11}}{4}\right)^{\frac{1}{3}} = 2.4\cdot10^{-4}\ M$$

$Ag_2SO_4\ (s) \rightleftarrows 2\ Ag^+\ (aq) + SO_4^{2-}\ (aq)$ (x=2 e y=1)

$K_{PS} = [Ag^+]^2\cdot[SO_4^{2-}] = 1.2\cdot10^{-5}\ M^3$

$$S = \left(\frac{1.2\cdot10^{-5}}{2^2\cdot1^1}\right)^{\frac{1}{2+1}} = \left(\frac{1.2\cdot10^{-5}}{4}\right)^{\frac{1}{3}} = 1.4\cdot10^{-2}\ M$$

$Ca_3(PO_4)_2\ (s) \rightleftarrows 3\ Ca^{2+}\ (aq) + 2\ PO_4^{3-}\ (aq)$ (x=3 e y=2)

$K_{PS} = [Ca^{2+}]^3\cdot[PO_4^{3-}]^2 = 2.0\cdot10^{-29}\ M^5$

$$S = \left(\frac{2.0\cdot10^{-29}}{3^3\cdot2^2}\right)^{\frac{1}{3+2}} = \left(\frac{2.0\cdot10^{-29}}{108}\right)^{\frac{1}{5}} = 7.1\cdot10^{-7}\ M$$

Some compounds with low solubility and their K_{PS}

formula	mane	K_{PS} (25 ° C)	solubility S (M)
$CaCO_3$	Calcium carbonate	$3.4 \cdot 10^{-9}$ M^2	$5.8 \cdot 10^{-5}$
$MnCO_3$	Manganes II carbonate(II)	$2.3 \cdot 10^{-11}$ M^2	$4.8 \cdot 10^{-6}$
$FeCO_3$	Iron II carbonate(II)	$3.1 \cdot 10^{-11}$ M^2	$5.6 \cdot 10^{-6}$
CaF_2	Calcium fluoride	$5.3 \cdot 10^{-11}$ M^3	$2.4 \cdot 10^{-4}$
$AgCl$	Silver chloride	$1.8 \cdot 10^{-10}$ M^2	$1.3 \cdot 10^{-5}$
$AgBr$	Silver bromide	$5.4 \cdot 10^{-13}$ M^2	$7.3 \cdot 10^{-7}$
$CaSO_4$	Calcium sulphate	$4.9 \cdot 10^{-5}$ M^2	$7.0 \cdot 10^{-3}$
$BaSO_4$	Barium sulphate	$1.1 \cdot 10^{-10}$ M^2	$1.0 \cdot 10^{-5}$
$SrSO_4$	Strontium sulphate	$3.4 \cdot 10^{-7}$ M^2	$5.8 \cdot 10^{-4}$
$Ca(OH)_2$	Calcium hydroxide	$5.5 \cdot 10^{-5}$ M^3	$2.4 \cdot 10^{-2}$
$Mg(OH)_2$	Magnesium hydroxide	$5.6 \cdot 10^{-12}$ M^3	$1.1 \cdot 10^{-4}$
CaC_2O_4	Calcium oxalate	$1.3 \cdot 10^{-9}$ M^2	$3.6 \cdot 10^{-5}$
$Ca_3(PO_4)_2$	Calcium phosphate	$2.0 \cdot 10^{-29}$ M^5	$7.1 \cdot 10^{-7}$

Relationship between solubility and K_{PS}

Cubic crystals of fluorite

$$CaF_2\ (s) \rightleftarrows Ca^{2+}\ (aq) + 2\ F^-\ (aq)$$

If at 25 ° C we measure $[Ca^{2+}] = 2.4 \times 10^{-4}$ M, which is K_{PS} ?

$$[F^-] = 2\ [Ca^{2+}] = 2 \cdot 2.4 \cdot 10^{-4} = 4.8 \cdot 10^{-4}\ M$$

therefore:

$$K_{PS} = [Ca^{2+}] \cdot [F^-]^2 = 2.4 \cdot 10^{-4} \cdot 4.8 \cdot 10^{-4} = 5.3 \cdot 10^{-11}\ M^3 \text{ a } 25\ °\ C.$$

The solubility is $S = [Ca^{2+}] = [F^-] / 2 = 2.4 \times 10^{-4}$ M

Barium sulphate is opaque to x-rays andit is used in to examine the digestive tract

barium sulphate is insoluble in water and therefore in organic liquids can not be absorbed by the body and can not exert toxic or lethal action.

$BaSO_4$ (s) ⇄ Ba^{2+} (aq) + SO_4^{2-} (aq)

$K_{PS} = [Ba^{2+}]\ [SO_4^{2-}] = 1.1 \times 10^{-10}\ M^2$

$\text{solubility} = [Ba^{2+}] = [SO_4^{2-}] = 1.0 \times 10^{-5}\ M$

Barium is characterized, as all positive contrast agents, by a high atomic number (Z = 56)

Relationship between solubility and K_{PS}

K_{PS} values for low soluble salts can be used to calculate the solubility of a salt or to determine if a precipitate will form when solutions of its cation and its anion are mixed.

Example 1: for Barium sulphate at 25 ° C is $K_{PS} = 1.1 \cdot 10^{-10}\ M^2$.
Calculate its solubilty: a) in moles /L e b) g/L ($M_{BaSO_4} = 233$ g/mol)

$$BaSO_4\,(s) \rightleftarrows Ba^{2+}\,(aq) + SO_4^{2-}\,(aq) \qquad K_{PS} = [Ba^{2+}] \cdot [SO_4^{2-}] = 1.1 \cdot 10^{-10}\ M^2$$

Equation $BaSO_4\,(s) \rightleftarrows$	$Ba^{2+}\,(aq)$ +	$SO_4^{2-}\,(aq)$
Initial amount(M)	0	0
Amount dissolved(M)	+x	+x
At equilibrium (M)	x	x

$$K_{PS} = [Ba^{2+}] \cdot [SO_4^{2-}] = x^2 \qquad x = \sqrt{K_{PS}} = \sqrt{1.1 \cdot 10^{-10}} = 1.0 \cdot 10^{-5}\ M$$

a) Solubility in mol/L è $1.0 \cdot 10^{-5}$ M
b) Solubility in g/L è $1.0 \cdot 10^{-5} \cdot 233 = 0.0024$ g/L

Solubility and the "common ion effect"

What happens if to a saturated solution of an insoluble salt one adds its cation or anion?

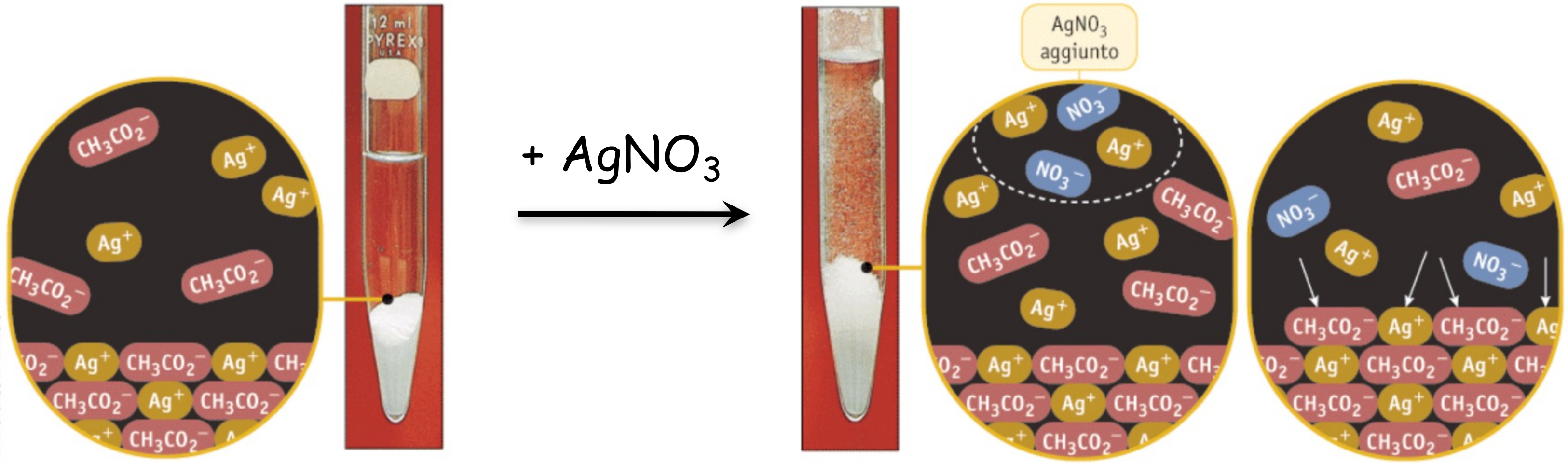
$$\mathrm{Ag\,CH_3COO\,(s)} \rightleftarrows \mathrm{Ag^+\,(aq) + CH_3COO^-\,(aq)}$$

$$K_{PS} = [\mathrm{Ag^+}] \times [\mathrm{CH_3COO^-}] = 2.0\times10^{-3}\ \mathrm{M}^2$$

In the saturated solution we have $[\mathrm{Ag^+}] = [\mathrm{CH_3COO^-}] = (K_{PS})^{\frac{1}{2}} = 0.045$ M.

If we add 0.1 M $\mathrm{AgNO_3}$ to this solution

$$Q = [\mathrm{Ag^+}]\,[\mathrm{CH_3COO^-}] = 0.145\times0.045 = 6.5\times10^{-3}\ \mathrm{M}^2 > K_{PS}$$



+ $\mathrm{AgNO_3}$

Accordng to Le Châtelier principle, some Silver acetate will precipitate .

How much silver acetate will precipitate if we cosided 1 L of solution(M_{AgCH_3COO} = 277.26 g/mol)?

Equation $AgCH_3COO$ (s) ⇄	Ag^+ (aq) +	CH_3COO^- (aq)
Initial amount (M)	0.045	0.045
Added amount (M)	0.045+0.1	0
Reacting amount (M)	−x	−x
At equilibrium (M)	0.145−x	0.045−x

$$K_{PS} = [Ag^+]\cdot[CH_3COO^-] = 2.0\text{x}10^{-3}\ M^2$$

$$2.0\text{x}10^{-3} = (0.145-x)(0.045-x) \quad x = 0.0276\ M$$

0.0276 mol/L will precipitate, that is 0.0276x277.26 = 7.6 g

K_{PS}, reaction quotient and precipitation

Q allows one to predict if:

i) a precipitate will be formed, if the ionic concentrations are known
ii) what are the ionic concentrations necessary to induce the precipitation of a salt that has a low solubility

$AgCl\ (s) \rightleftarrows Ag^{+}\ (aq) + Cl^{-}\ (aq)$ $K_{PS} = [Ag^{+}]\ [Cl^{-}] = 1.8\text{x}10^{-10}\ M^{2}$

If in a solution we have Ag^{+} e Cl^{-} at given concentrations $Q = [Ag^{+}]\ [Cl^{-}]$

if $Q = K_{PS}$, the solution is saturated and the system is at equilibrium

if $Q < K_{PS}$, the solution is not saturated

if $Q > K_{PS}$, the solution is beyond saturation and the system is not at equilibrium, we will observe precipitation.

Example 3: The solubility product for magnesium hydroxide at 25 ° C è K_{PS} = $5.6 \cdot 10^{-12}$ M^3. Initially $[Mg^{2+}] = 1.5 \cdot 10^{-6}$ M. a) if we add enough NaOH to achieve $[OH^-] = 1.0 \cdot 10^{-4}$ M, wiil Mg OH_2 precipitate? b) if it does not precipitate, will it do so if OH^- is increased up to $1.0 \cdot 10^{-2}$ M?

$$Mg(OH)_2\ (s) \rightleftarrows Mg^{2+}\ (aq) + 2\ OH^-\ (aq)$$

$$K_{PS} = [Mg^{2+}] \cdot [OH^-]^2 = 5.6 \cdot 10^{-12}\ M^3$$

a) $Q = [Mg^{2+}] \cdot [OH^-]^2 = (1.5 \cdot 10^{-6}) \cdot (1.0 \cdot 10^{-4})^2 = 1.5 \cdot 10^{-14} < K_{PS}$

The solution is not saturated, and there will not be precipitation

b) $Q = [Mg^{2+}] \cdot [OH^-]^2 = (1.5 \cdot 10^{-6}) \cdot (1.0 \cdot 10^{-2})^2 = 1.5 \cdot 10^{-10} > K_{PS}$

$Mg(OH)_2$ will precipitate until the Mg^{2+} e OH^- decrease to reach the values for which $[Mg^{2+}] \cdot [OH^-]^2 = K_{PS}$

Alkali with low solubilty

Acid and base solutions are tolerated by our body (eg in food) only if their acidity or basicity is modest. Concentrated solutions of acids or bases are caustic and, by skin contact or if ingested, they produce severe tissue damage. It is therefore not possible to administer concentrated solutions of these compounds for therapeutic purposes.

Some commonly used compounds used in the symptomatic treatment of gastric hyperacidity are strong bases with low solubility, which have great potential basicity (because they can release large amounts of the hydroxide ion if they dissolve), but yield a weakly basic pH. An example used for this purpose is milk of magnesia.

$Mg(OH)_2\ (s) \rightleftarrows Mg^{2+}\ (aq) + 2\ OH^-\ (aq)$
$Al(OH)_3\ (s) \rightleftarrows Al^{3+}\ (aq) + 3\ OH^-\ (aq)$

A suspension of $Mg(OH)_2$ e $Al(OH)_3$

Magnesium hydroxide has a low solubility

$$Mg(OH)_2\ (s) \rightleftarrows Mg^{2+}\ (aq) + 2\ OH^-\ (aq)$$

$$K_{PS} = [Mg^{2+}] \cdot [OH^-]^2 = 5.6 \cdot 10^{-12}\ M^3$$

The pH of the saturated solution is ($[Mg^{2+}] = S$ e $[OH^-] = 2 \cdot S$)

$$[OH^-] = 2 \cdot S = 2 \cdot \sqrt[3]{\frac{K_{PS}}{4}} = 2 \cdot \sqrt[3]{\frac{5.6 \cdot 10^{-12}}{4}} = 2.2 \cdot 10^{-4}\ M$$

$$pOH = -\log_{10} 2.2 \cdot 10^{-4} = 3.65 \rightarrow pH = 10.35$$

Excess of acid in the gastric envirnment is neutralized:

$$Mg(OH)_2\ (s) \rightleftarrows Mg^{2+}\ (aq) + \cancel{2\ OH^-}\ (aq)$$

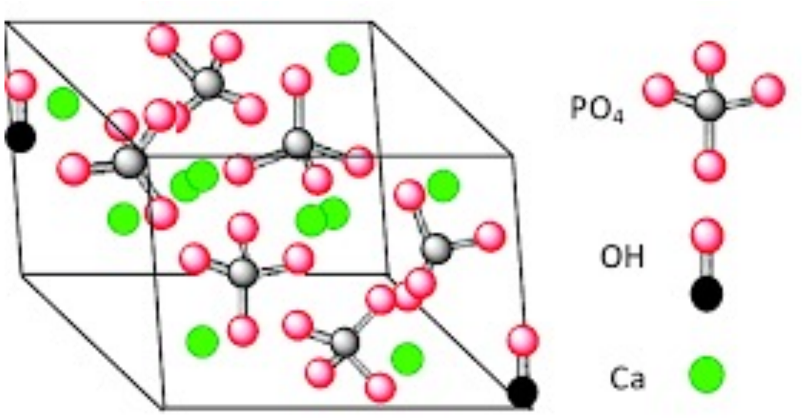
$$2\ H_3O^+\ (aq) + \cancel{2\ OH^-}\ (aq) \rightleftarrows Mg^{2+}\ (aq) + 4\ H_2O\ (l)$$

$$Mg(OH)_2\ (s) + 2\ H_3O^+\ (aq) \rightleftarrows Mg^{2+}\ (aq) + 4\ H_2O\ (l)$$

Net reaction

Low solubility salts from weak acids

1) The bone inorganic matrix is mad of calcium phosphate (actually hydroxyapatite)



$Ca_3(PO_4)_2\ (s) \rightleftarrows 3\ Ca^{2+}\ (aq) + 2\ PO_4^{3-}\ (aq)$

$K_{PS} = [Ca^{2+}]^3 \cdot [PO_4^{3-}]^2 = 2.0 \times 10^{-29}\ M^5$

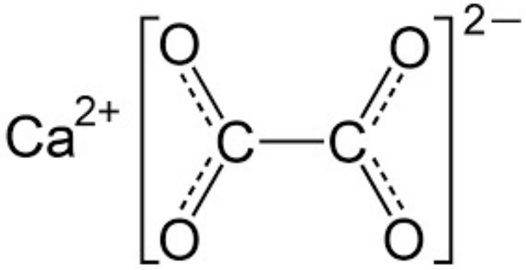
solubility = 7.1×10^{-7} M

2) In 90% of cases kidney stones are made of calcium oxalate

$CaC_2O_4\ (s) \rightleftarrows Ca^{2+}\ (aq) + C_2O_4^{-}\ (aq)$

$K_{PS} = [Ca^{2+}] \cdot [C_2O_4^{2-}] = 1.3 \times 10^{-9}\ M^2$

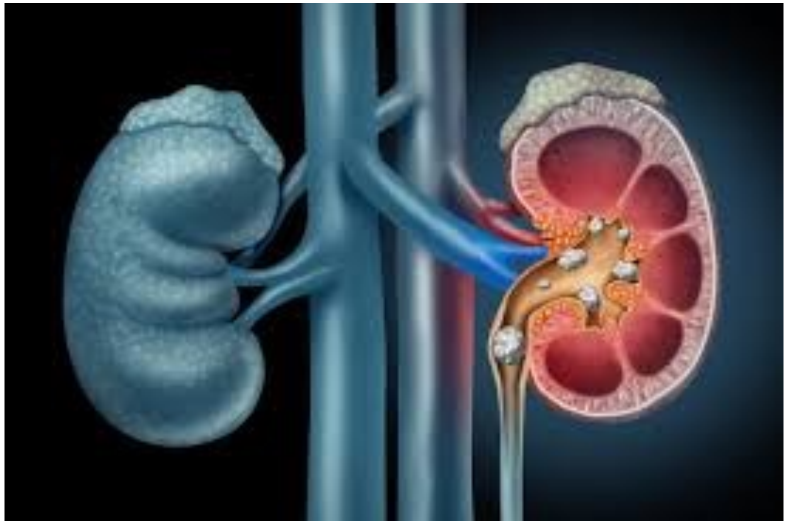
solubility = 3.6×10^{-5} M



Renal colic and the solubility product

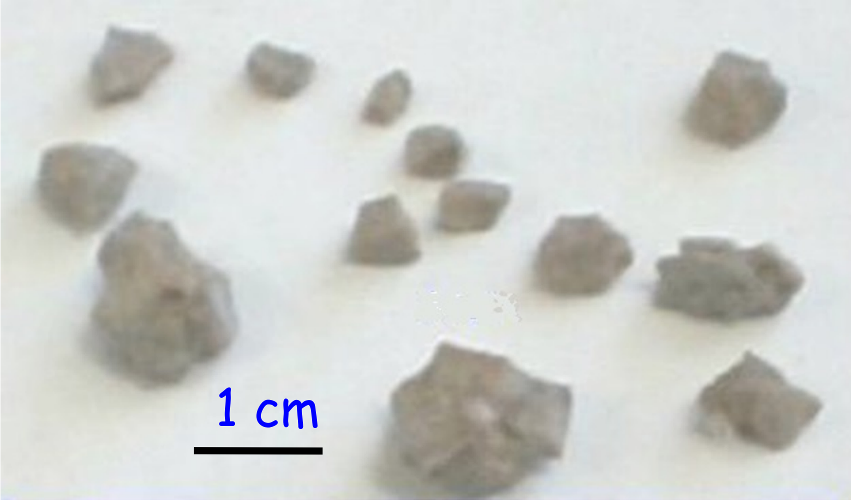
Renal colic affects about 1.2 million people each year, and is about 1% of the causes of hospitalization. Kidney stones, which cause them, are a problem that can arise at any level of the urinary tract, mainly form in the kidney, with a frequency that reaches 12% for men, about double that for women, and a peak of incidence around 35 years.

Renal colic generally results from the mobilization of a kidney stone and usually manifests itself with violent pain in the right or left lumbar region, which radiates to the respective flank and then to the genitals; nausea and vomiting are also often present.

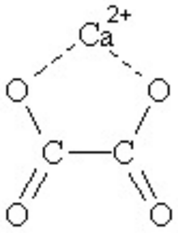


Kidney stones can be visible to X-rays (contain calcium) or invisible to X-rays (if calcium is not present)

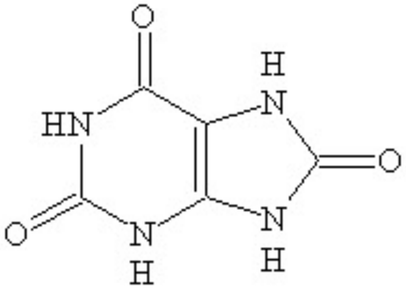
- 60-90%: calcium oxalate or calcium phosphate
- 10-15%: kidney stones made of struvite ($NH_4MgPO_4 \cdot 6\ H_2O$)
- 5-10%: uric acid
- 1%: made by cistine (oxidized dimer of cisteine).



Kidney stonesmade of calcium oxalate (CaC_2O_4)



CaC_2O_4



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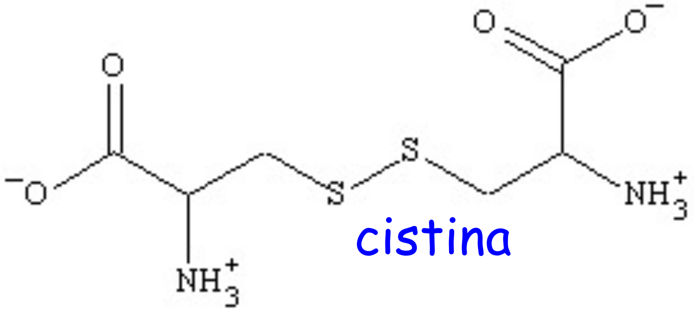
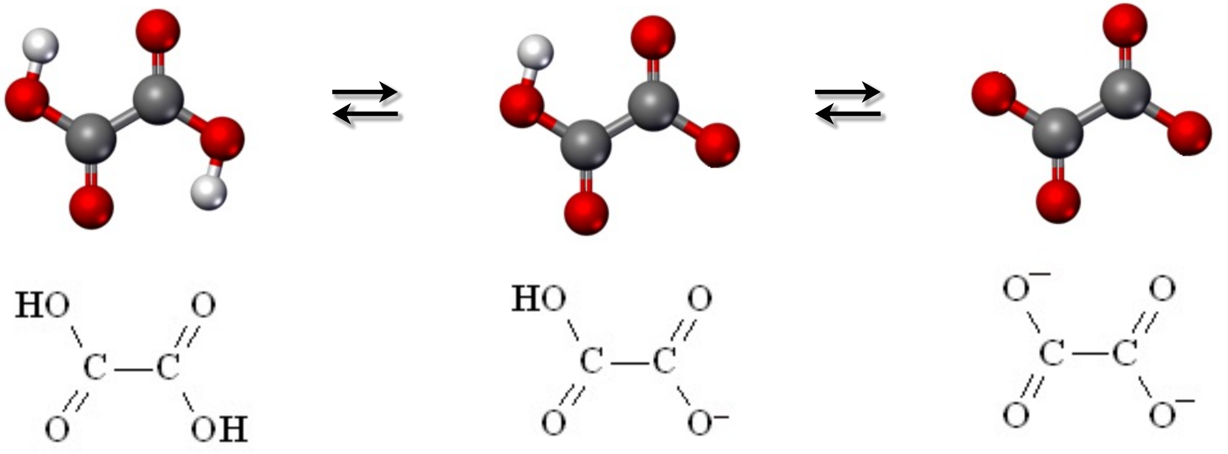


cistina

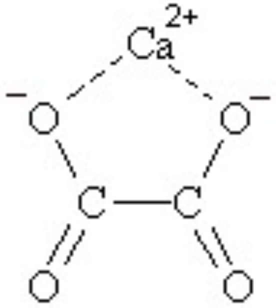
Oxalic acid is a weak diprotic acid

$H_2C_2O_4$ (aq) + H_2O (l) $\rightleftarrows$ $HC_2O_4^-$ (aq) + H_3O^+ (aq) $K_1 = 5.4 \cdot 10^{-2}$ M

$HC_2O_4^-$ (aq) + H_2O (l) $\rightleftarrows$ $C_2O_4^{2-}$ (aq) + H_3O^+ (aq) $K_2 = 5.2 \cdot 10^{-5}$ M



In water its solubility is 143 g/L a 25 ° C (1.59 M).
If Ca^{2+} is added however, the resulting salt (Calcium oxalate) is rather insoluble:



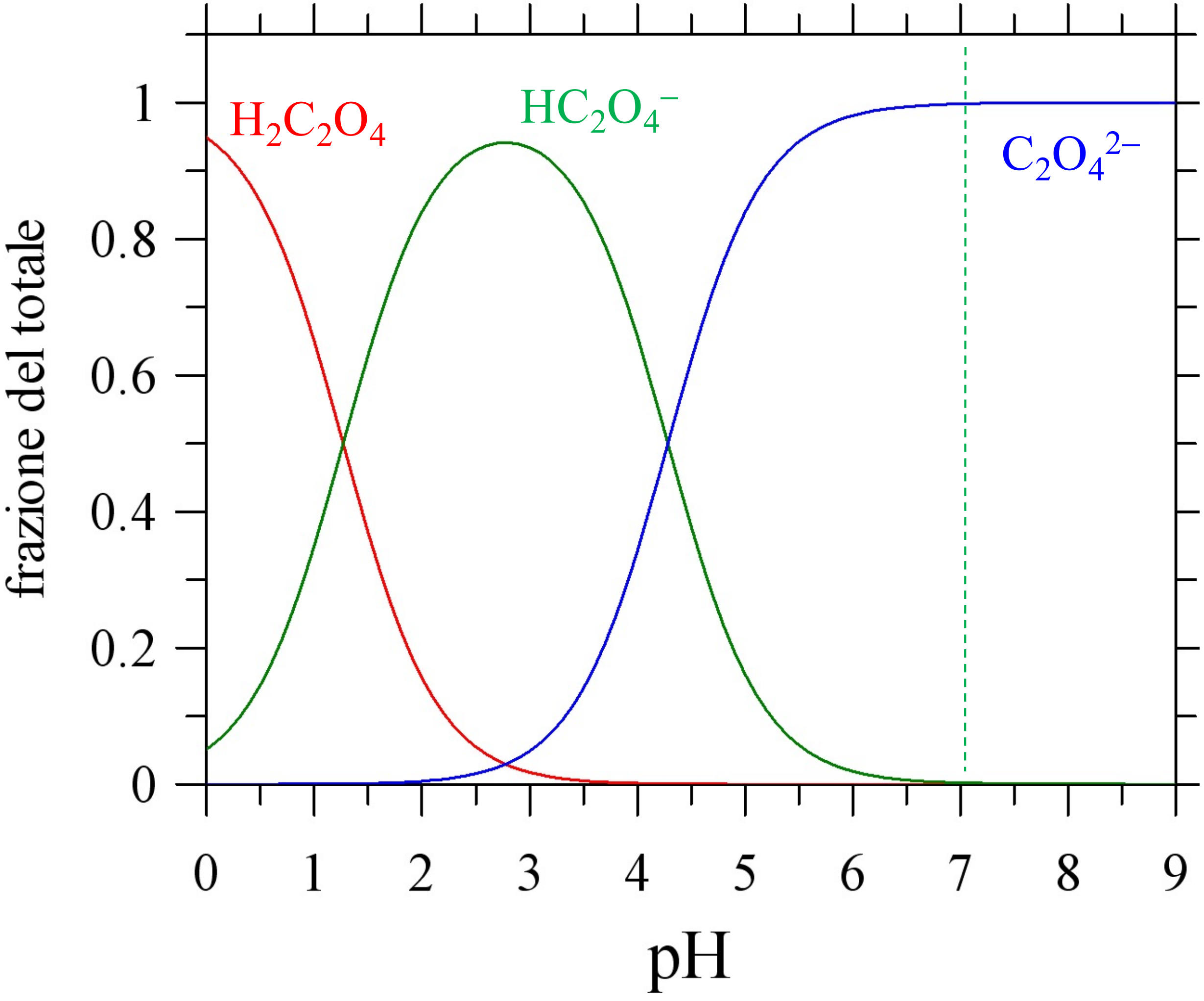
CaC_2O_4 (s) $\rightleftarrows$ Ca^{2+} (aq) + $C_2O_4^-$ (aq)

$K_{PS} = [Ca^{2+}]\,[C_2O_4^{2-}] = 1.3\text{x}10^{-9}\ M^2$ solubility = $3.6\text{x}10^{-5}$ M

What conditions favour calcium oxalate formation??

$H_2C_2O_4\ (aq) + H_2O\ (l) \rightleftarrows HC_2O_4^-\ (aq) + H_3O^+\ (aq) \quad K_1 = 5.4\text{x}10^{-2}\ M$

$HC_2O_4^-\ (aq) + H_2O\ (l) \rightleftarrows C_2O_4^{2-}\ (aq) + H_3O^+\ (aq) \quad K_2 = 5.2\text{x}10^{-5}\ M$



pH of urine: 4.6–8.0

A pH > 7 is 99% present as oxalate $C_2O_4^{2-}$

Data from “Interpretation of Diagnostic Tests” by J. Wallach, Little Brown & Company Eds, 1979

$[Ca^{2+}]$ in urine = 2–8 mM

$K_{PS} = [Ca^{2+}]\,[C_2O_4^{2-}] = 1.3\text{x}10^{-9}\ M^2$

$[C_2O_4^{2-}] = K_{PS} / [Ca^{2+}] = 1.3\text{x}10^{-9} / 0.008 = 1.6\text{x}10^{-7}\ M$

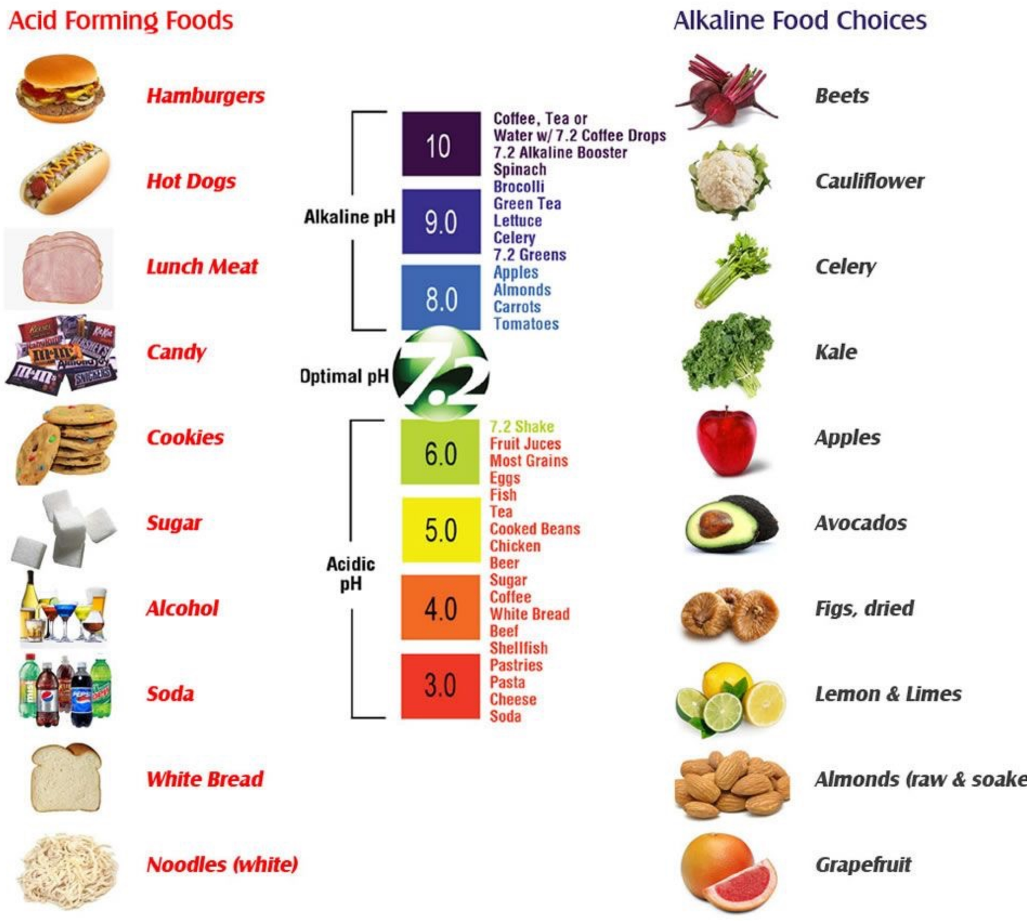
Therefore at pH > 7, a $[Ca^{2+}]$ > 1 mM and oxalate > 10^{-7} M, a precipitate will appear, thac could lead to kidney stone formation.

In Hyperoxaluria

uribalry oxalate > 40 mg/die = $1.3\text{x}10^{-9}$ moli/die
volume of urine = 600–2500 mL/die
[oxalate] = $2\text{–}7\text{x}10^{-4}$ moli

$Q = [Ca^{2+}]\,[C_2O_4^{2-}] = 2.0\text{x}10^{-3}\,\text{x}2.0\text{x}10^{-4} = 4.0\text{x}10^{-7} >> K_{PS}$

Hyperoxaluria can be primary or caused by excessive ingestion of foods containing oxalate (eg, rhubarb, spinach, cocoa, peanuts, pepper, tea) or by excessive reabsorption of oxalates due to various intestinal diseases (eg. bacterial overgrowth syndrome, chronic pancreatitis or biliary disease)



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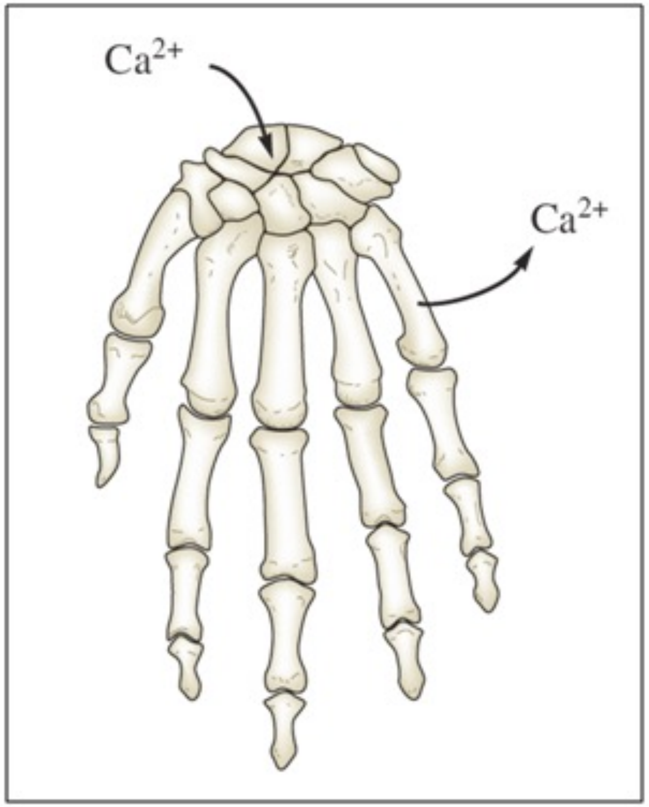
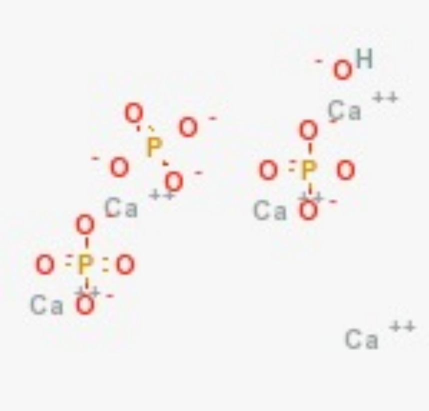
Calcium

Calcium in body fluids is constantly exchanged at millions of construction sites. The reconstruction repairs microfractures and imperfections, allows the growth of the organism and helps in the repair of fractures.



Calcium



- 1 kg / 70 kg body weight
  - 98% in bones as hydroxyapatite $Ca_5(PO_4)_3(OH)$. The OH^- ion can be substitutes by F^-, Cl^-, CO_3^{2-}
  - 1% interstitial fluid
  - 1% in plasma
- Ca^{2+}, free and ionized in plasma is 45% del totale (1.1-1.3 mM)
- Ca^{2+} complexed w/ phosphate (5%), citrate (5%), HCO_3^- and albumine (45%)
  - Ca^{2+} and PO_4^{3-} in plasma are close to the solubility
  - In complex with albumin, free $[Ca^{2+}]$ decreases and the precipitation of calcium phoshate is avoided

$$Ca_3(PO_4)_2\ (s) \rightleftarrows 3\ Ca^{2+}\ (aq) + 2\ PO_4^{3-}\ (aq)$$

$K_{PS} = [Ca^{2+}]^3\ [PO_4^{3-}]^2 = 2.0\times10^{-29}\ M^5$ solubility = 7.1×10^{-7} M

Phosphoric acid

$H_3PO_4\,(aq) + H_2O\,(l) \rightleftarrows H_2PO_4^{-}\,(aq) + H_3O^{+}\,(aq)$ $K_1 = 7.1\text{x}10^{-3}$

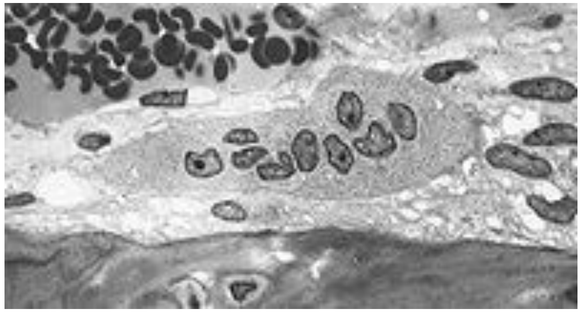
$H_2PO_4^{-}\,(aq) + H_2O\,(l) \rightleftarrows HPO_4^{2-}\,(aq) + H_3O^{+}\,(aq)$ $K_2 = 6.2\text{x}10^{-8}$

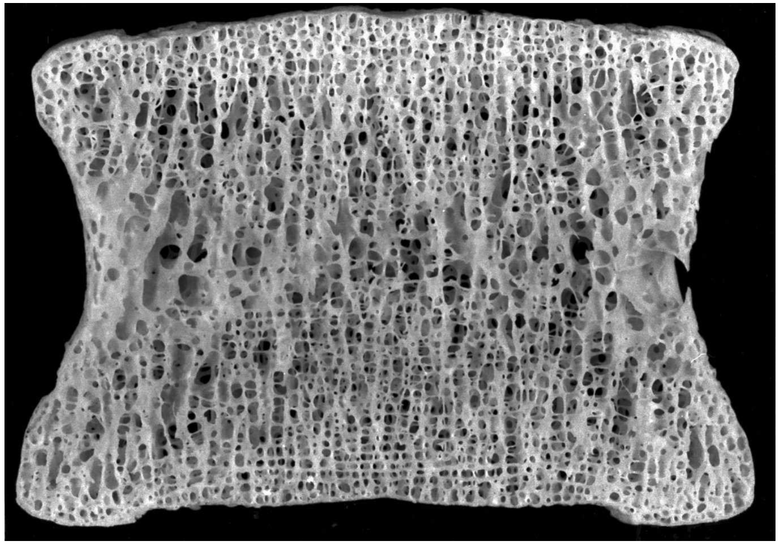
$HPO_4^{2-}\,(aq) + H_2O\,(l) \rightleftarrows PO_4^{3-}\,(aq) + H_3O^{+}\,(aq)$ $K_3 = 4.4\text{x}10^{-13}$

Osteoclasts produce acidic equivalents and enzymes that dissolve the bone matrix.

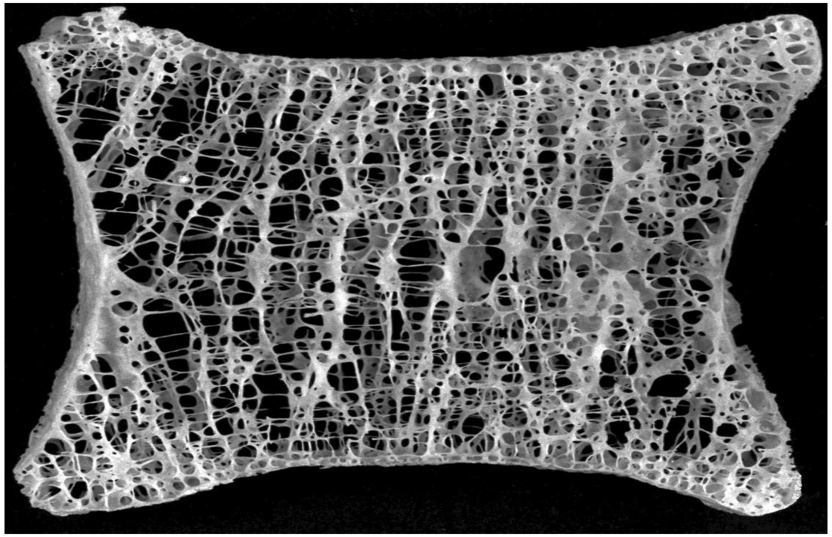
$$\xrightarrow{+\,H_3O^{+}}$$

$$Ca_3(PO_4)_2\,(s) \rightleftarrows 3\,Ca^{2+}\,(aq) + 2\,PO_4^{3-}\,(aq)$$





Normal bone



osteoporosis