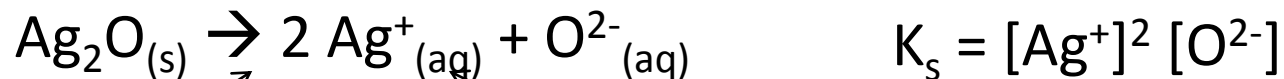


Do now:

The solubility of Ag_2O in pure water is $0.00108 \text{ mol.L}^{-1}$, calculate the solubility constant, K_s .



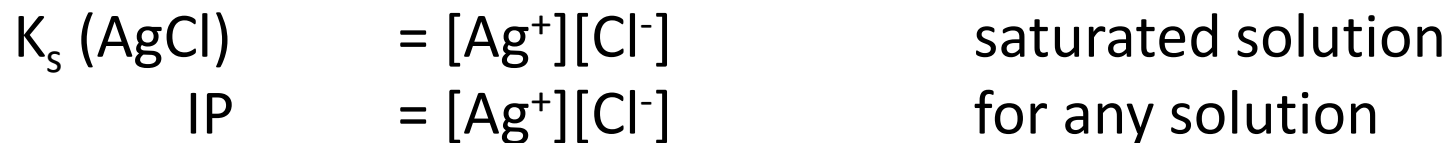
$$[\text{O}^{2-}] = 0.00108 \text{ mol.L}^{-1}$$

$$[\text{Ag}^+] = 2 \times 0.00108 \text{ mol.L}^{-1} = 0.00216 \text{ mol.L}^{-1} \quad K_s = 0.00216^2 \times 0.00108$$
$$= 5.04 \times 10^{-9}$$

Predicting precipitation

We can use K_s values to predict if a solution will form a precipitate or what concentration of an ion is needed to form a precipitate.

K_s is for **saturated** solutions in **equilibrium**, but the same expression can be used to calculate the ionic product (IP) of any solution.



If the IP is greater than K_s a precipitate will form.

Predicting precipitation

If solid sodium chloride is added to a 0.01 mol.L^{-1} solution of AgNO_3 solution what is the minimum concentration of Cl^- needed to give a precipitate of AgCl ? $K_s(\text{AgCl}) = 2 \times 10^{-10}$.



What information do we have?

$$[\text{Ag}^+] = 0.01 \text{ mol.L}^{-1} \quad [\text{Cl}^-] = ? \quad K_s = 2 \times 10^{-10}$$

If $IP > K_s$ then precipitation will occur.

$$[\text{Cl}^-] > \frac{K_s}{[\text{Ag}^+]} \quad [\text{Cl}^-] > \frac{2 \times 10^{-10}}{0.01} \quad [\text{Cl}^-] > 2 \times 10^{-8} \text{ mol.L}^{-1}$$

Predicting precipitation

If solid copper sulfate is added to a 0.025 mol.L^{-1} solution of sodium carbonate solution what is the maximum amount of Cu^{2+} that can be added to the sodium carbonate solution so a precipitate of CuCO_3 will not form? $K_s(\text{CuCO}_3) = 7.8 \times 10^{-9}$.

Workbook pg 198 Q1

Predicting precipitation

Will a precipitate form when 75 mL of $4.0 \times 10^{-3} \text{ mol.L}^{-1}$ NaCl solution and 25 mL of $6.0 \times 10^{-5} \text{ mol.L}^{-1}$ AgNO₃ solution are mixed?
 $K_s(\text{AgCl}) = 1.8 \times 10^{-10}$.



What do we know? What should we calculate?

Calculate IP with [Ag⁺] and [Cl⁻]

Remember we have mixed two solutions together. Total volume is now 100 mL.

$$\text{So: } [\text{Ag}^+] = 6.0 \times 10^{-5} \times (25/100) = 1.5 \times 10^{-5} \text{ mol.L}^{-1}$$

$$[\text{Cl}^-] = 4.0 \times 10^{-3} \times (75/100) = 3.0 \times 10^{-3} \text{ mol.L}^{-1}$$

$$\text{IP} = 1.5 \times 10^{-5} \times 3.0 \times 10^{-3} = 4.5 \times 10^{-8}$$

Is this bigger than K_s ?

YES! Precipitate will form.

Predicting precipitation

10.0 mL of 0.001 mol.L⁻¹ solution of CaCl₂ was mixed with 10.0 mL of 0.001 mol.L⁻¹ solution of Na₂SO₄. Will a precipitate form of calcium sulfate form?

$$K_s(\text{CaSO}_4) = 2 \times 10^{-5}.$$

Workbook pg 199 Q2

Workbook pg 200 Q5

2012 Exam

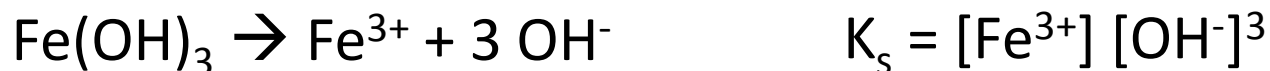
- (c) (i) Determine whether a precipitate of iron(III) hydroxide, $\text{Fe}(\text{OH})_3$, will form when $\text{Fe}(\text{NO}_3)_3$ is dissolved in water. $[\text{Fe}(\text{NO}_3)_3] = 1.05 \times 10^{-4} \text{ mol L}^{-1}$.

Assume the pH of the water is 7.

$$K_s(\text{Fe}(\text{OH})_3) = 2.00 \times 10^{-39}$$

What do we know?

$$[\text{Fe}^{3+}] = 1.05 \times 10^{-4} \text{ mol.L}^{-1} \quad [\text{OH}^-] = 1.00 \times 10^{-7} \text{ mol.L}^{-1}$$



$$\begin{aligned} \text{IP} &= 1.05 \times 10^{-4} \times (1.00 \times 10^{-7})^3 \\ &= 1.05 \times 10^{-25} \end{aligned}$$

$\text{IP} > K_s$
so precipitation will occur

- (ii) Discuss the effect of decreasing the pH of the water on the solubility of $\text{Fe}(\text{OH})_3$.

2012 Exam

Assessment schedule

(c) (i)	$\text{Fe}(\text{OH})_3(s) \rightleftharpoons \text{Fe}^{3+}(aq) + 3\text{OH}^-(aq)$ <p>Ion Product (IP) = $[\text{Fe}^{3+}][\text{OH}^-]^3$</p> <p>At pH 7, $[\text{OH}^-] = 1 \times 10^{-7} \text{ mol L}^{-1}$</p> $\text{IP} = [1.05 \times 10^{-4}][1 \times 10^{-7}]^3 = 1.05 \times 10^{-25}$ <p>Since $\text{IP} > K_s$, $\text{Fe}(\text{OH})_3$ will form a precipitate</p>
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For A

- Either:
Correct IP expression
OR
Compares IP and K_s to make a valid conclusion.

For M

Method uses correct IP expression but has one calculation error AND
Compares IP and K_s to make a valid conclusion
(3 significant figures).

For E

- Answer correct with supporting calculation (3 significant figures).

2012 Exam

- (ii) When the pH is decreased, $[\text{H}_3\text{O}^+]$ will increase. The H_3O^+ will react with the OH^- and therefore remove them from the equilibrium. This will cause the reaction to replace some of the removed OH^- . As a result more $\text{Fe}(\text{OH})_3$ will dissolve, so decreasing the pH will increase the solubility of $\text{Fe}(\text{OH})_3$.

For A

- Either:
Writes an equilibrium expression AND identifies direction it shifts in.

OR

States $[\text{OH}^-]$ decreases / $[\text{H}_3\text{O}^+]$ increases causing $\text{Fe}(\text{OH})_3$ to be more soluble.

For M

- Either:
States the change in $[\text{OH}^-]$, its impact on the equilibrium position and therefore more $\text{Fe}(\text{OH})_3$ dissolves.

OR

Discussion of effect of decreasing pH on $\text{Fe}(\text{OH})_3$ dissolving in terms of $[\text{H}_3\text{O}^+]$ / $[\text{OH}^-]$ changing.

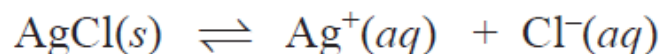
For E

- Complete discussion of effect of decreasing pH on $\text{Fe}(\text{OH})_3$ solubility, including role of H_3O^+ (reacting with OH^-).

2009 Exam Q 2 (b)

QUESTION TWO

Addition of chloride ions to a solution of silver nitrate often results in the formation of a white precipitate of silver chloride (AgCl).



$$K_s(\text{AgCl}) = 1.56 \times 10^{-10}$$

(b) Solid sodium chloride is added to 5.00 L of 0.100 mol L⁻¹ silver nitrate solution.

Calculate the minimum mass of sodium chloride that would be needed to produce a saturated solution of AgCl. Assume that there is no change in volume when the sodium chloride is added.

$$M(\text{NaCl}) = 58.5 \text{ g mol}^{-1}$$

2009 Exam Q 2 (b)

(b)

$$K_s = [\text{Ag}^+][\text{Cl}^-]$$
$$1.56 \times 10^{-10} = [0.100][\text{Cl}^-]$$
$$[\text{Cl}^-] = 1.56 \times 10^{-9}$$

$$n = c \times V$$
$$= 1.56 \times 10^{-9} \times 5.00 \text{ mol}$$
$$= 7.80 \times 10^{-9} \text{ mol}$$

$$m = n \times M$$
$$= 7.80 \times 10^{-9} \text{ mol} \times 58.5 \text{ g mol}^{-1}$$
$$= 4.56 \times 10^{-7} \text{ g}$$

- Correct answer for mass of NaCl
Sensible rounding and units are required.

2014 Exam

- (b) A sample of seawater has a chloride ion concentration of 0.440 mol L^{-1} .

Determine whether a precipitate of lead(II) chloride will form when a 2.00 g sample of lead(II) nitrate is added to 500 mL of the seawater.

$$K_s(\text{PbCl}_2) = 1.70 \times 10^{-5} \quad M(\text{Pb}(\text{NO}_3)_2) = 331 \text{ g mol}^{-1}$$

$$\begin{aligned} n(\text{Pb}(\text{NO}_3)_2) &= \frac{2.00 \text{ g}}{331 \text{ g mol}^{-1}} \\ &= 6.04 \times 10^{-3} \text{ mol} \end{aligned}$$

$$\begin{aligned} \therefore [\text{Pb}^{2+}] &= 6.04 \times 10^{-3} \text{ mol} / 0.500 \text{ L} \\ &= 1.21 \times 10^{-2} \text{ mol L}^{-1} \end{aligned}$$

$$\begin{aligned} Q &= (1.21 \times 10^{-2}) \times (0.440)^2 \\ &= 2.34 \times 10^{-3} \end{aligned}$$

As $Q > K_s$, a precipitate will form.