Do now:

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

$$Ag_2O_{(s)} \rightarrow 2\ Ag^+_{(aq)} + O^{2-}_{(aq)}$$

$$K_s = [Ag^+]^2\ [O^{2-}]$$

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

$[Ag^+]$ = 2 x 0.00108 mol.L^{-1} = 0.00216 mol.L^{-1}

$K_s = 0.00216^2 \times 0.00108$

$= 5.04 \times 10^{-9}$

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.

K_s (AgCl)	$= [Ag^+][Cl^-]$	saturated solution
IP	$= [Ag^+][Cl^-]$	for 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? $K_s(AgCl) = 2 \times 10^{-10}$.

$AgCl_{(s)} \rightarrow Ag^+_{(aq)} + Cl^-_{(aq)}$ $\quad\quad$ $K_s = [Ag^+][Cl^-]$

What information do we have?

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

If IP > K_s then precipitation will occur.

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

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? $K_s(CuCO_3) = 7.8 \times 10^{-9}$.

Workbook pg 198 Q1

Predicting precipitation

Will a precipitate form when 75 mL of 4.0×10^{-3} mol.L^{-1} NaCl solution and 25 mL of 6.0×10^{-5} mol.L^{-1} $AgNO_3$ solution are mixed? $K_s(AgCl) = 1.8 \times 10^{-10}$.

$AgCl_{(s)} \rightarrow Ag^+_{(aq)} + Cl^-_{(aq)}$ $\quad$ $K_s = [Ag^+][Cl^-] \quad = 1.8 \times 10^{-10}$

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.

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

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

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

Is this bigger than K_s?

YES! Precipitate will form.

Predicting precipitation

10.0 mL of 0.001 mol.L^{-1} solution of $CaCl_2$ was mixed with 10.0 mL of 0.001 mol.L^{-1} solution of Na_2SO_4. Will a precipitate form of calcium sulfate form?
$K_s(CaSO_4) = 2 \times 10^{-5}$.

Workbook pg 199 Q2

Workbook pg 200 Q5

2012 Exam

(c) (i) Determine whether a precipitate of iron(III) hydroxide, $Fe(OH)_3$, will form when $Fe(NO_3)_3$ is dissolved in water. $[Fe(NO_3)_3] = 1.05 \times 10^{-4}$ mol L^{-1}.

Assume the pH of the water is 7.

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

What do we know?

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

$Fe(OH)_3 \rightarrow Fe^{3+} + 3\,OH^-$ $K_s = [Fe^{3+}]\,[OH^-]^3$

$$IP = 1.05 \times 10^{-4} \times (1.00 \times 10^{-7})^3$$
$$= 1.05 \times 10^{-25}$$

$IP > K_s$
so precipitation will occur

(ii) Discuss the effect of decreasing the pH of the water on the solubility of $Fe(OH)_3$.

2012 Exam

Assessment schedule

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

For A	For M	For E
• Either: Correct IP expression OR Compares IP and K_s to make a valid conclusion.	Method uses correct IP expression but has one calculation error AND Compares IP and K_s to make a valid conclusion (3 significant figures).	• Answer correct with supporting calculation (3 significant figures).

2012 Exam

(ii) When the pH is decreased, $[H_3O^+]$ 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 $Fe(OH)_3$ will dissolve, so decreasing the pH will increase the solubility of $Fe(OH)_3$.

For A

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

  OR

  States $[OH^-]$ decreases / $[H_3O^+]$ increases causing $Fe(OH)_3$ to be more soluble.

For M

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

  OR

  Discussion of effect of decreasing pH on $Fe(OH)_3$ dissolving in terms of $[H_3O^+]$ / $[OH^-]$ changing.

For E

- Complete discussion of effect of decreasing pH on $Fe(OH)_3$ 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).

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

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

(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.

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

2009 Exam Q 2 (b)

(b)	$K_s = [Ag^+][Cl^-]$ $1.56 \times 10^{-10} = [0.100][Cl^-]$ $[Cl^-] = 1.56 \times 10^{-9}$ $n = c \times V$ $= 1.56 \times 10^{-9} \times 5.00$ mol $= 7.80 \times 10^{-9}$ mol $m = n \times M$ $= 7.80 \times 10^{-9}$ mol $\times$ 58.5 g mol^{-1} $= 4.56 \times 10^{-7}$ 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.

$K_s(PbCl_2) = 1.70 \times 10^{-5}$ $\quad M(Pb(NO_3)_2) = 331$ g mol^{-1}

$$n(Pb(NO_3)_2) = \frac{2.00\text{ g}}{331\text{ g mol}^{-1}} = 6.04 \times 10^{-3}\text{ mol}$$

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

$$Q = (1.21 \times 10^{-2}) \text{ x } (0.440)^2 = 2.34 \times 10^{-3}$$

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