Chemistry 225 Semester 04-2016
Homework #4b - Key

1) The solubility product of magnesium hydroxide is 1.8×10^{-11} @25°C.

a) Write the equilibrium equation relating to this solubility product. (2)

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

b) Write the equilibrium expression relating to it in the form K_{sp}=. (2)

$$K_{sp} = [Mg^{2+}][OH^-]^2$$

c) Calculate the molar solubility of magnesium hydroxide in water at 25°C, assuming no hydroxide ions are contributed by water. (4)

	$Mg(OH)_2$	Mg^{2+}	OH^-
I /M	--	0	0[1]
C /M	--	+x	+2x
E /M	--	x	2x

[1]Assuming the OH^- contributed by water is negligible.

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

$$x = \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{1.8\times10^{-11}}{4}} == 1.6510\times10^{-4}$$

$$\therefore S(Mg(OH)_2) = [Mg^{2+}] = x = \underline{\underline{1.7\times10^{-4}\ M}} \text{ to 2 s.f. at 25°C}$$

d) What effect would you expect the addition of sodium hydroxide solution to a mixture of magnesium hydroxide and water would have on the solubility of the magnesium hydroxide? Explain with reference to Le Chatelier's principle. (3)

The addition of sodium hydroxide would be expected to decrease the solubility of magnesium hydroxide. It would increase the concentration of OH^- and so the equilibrium of 1(a) would respond to try to decrease that concentration again. This implies a shift to the left[2] in favour of undissolved magnesium hydroxide.

e) $35cm^3$ of 0.010M magnesium sulfate solution is added to $65cm^3$ of 0.010M sodium hydroxide solution. Determine whether precipitation occurs or not. (4)

We calculate the concentration of those ions involved in the equilibrium of 1(a)

[2]This is an example of the common ion effect.

immediately after mixing, before any reaction occurs. The sulfate and sodium ions are irrelevant and must not be included. Note that the total volume becomes 35 + 65 = 100 cm^3

$$[Mg^{2+}] = \frac{35}{100} \times .01 = 3.5 \times 10^{-3} M$$

$$[OH^-] = \frac{65}{100} \times .01 = 6.5 \times 10^{-3} M$$

Since the system in not at equilibrium, $[Mg^{2+}][OH^-]^2$ = Q, not K

$$\therefore Q = 3.5 \times 10^{-3} \times (6.5 \times 10^{-3})^2 = 1.47875 \times 10^{-7}$$

Since K = 1.8×10^{-11}, Q > K and so the reaction as in 1(a) will proceed to the left, causing precipitation.

2) Chlorine gas may be prepared by the action of potassium permanganate on concentrated hydrochloric acid in aqueous solution. The permanganate ions are reduced to manganese(II), and the chloride ions oxidised to chlorine gas. Derive a balanced ionic equation to show this process, showing the essential steps of the derivation. (6)

NB: the potassium ions play no part in the reaction and must not be included.

Reduction

The unbalanced reduction reaction is: $MnO_4^- \rightarrow Mn^{2+}$

Add H_2O to balance O: $MnO_4^- \rightarrow Mn^{2+} + 4H_2O$

Add H^+ to balance H: $MnO_4^- + 8H^+ \rightarrow Mn^{2+} + 4H_2O$

Add electrons to balance charge: $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ ...(1)

Oxidation

The unbalanced oxidation reaction is: $Cl^- \rightarrow Cl_2$

Balance non-O, non-H element: $2Cl^- \rightarrow Cl_2$

No O to balance.

No H to balance.

Add electrons to balance charge: $2Cl^- \rightarrow Cl_2 + 2e^-$ …(2)

Unite the half-equations

Multiply (1) by 2 to give: $2MnO_4^- + 16H^+ + 10e^- \rightarrow 2Mn^{2+} + 8H_2O$

and (2) by 5 to give: $10Cl^- \rightarrow 5Cl_2 + 10e^-$

This equalises the number of electrons so that they will cancel.

Add the equations:

$2MnO_4^- + 16H^+ + 10e^- + 10Cl^- \rightarrow 2Mn^{2+} + 8H_2O + 5Cl_2 + 10e^-$

Cancel electrons:

$2MnO_4^- + 16H^+ + 10Cl^- \rightarrow 2Mn^{2+} + 8H_2O + 5Cl_2$

Put in states:

$\underline{\underline{2MnO_4^-(aq) + 16H^+(aq) + 10Cl^-(aq) \rightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 5Cl_2(g)}}$