

## 2.6. Reactions of Inorganic Compounds in Aqueous Solution

### Lewis acids and bases

Definitions: **Lewis acid**: electron pair acceptor  
**Lewis base**: electron pair donor

In the formation of complex ions the ligand is the Lewis base because it is donating a pair of electrons in the dative covalent bond and the metal ion is the Lewis acid.

### Metal-aqua ions

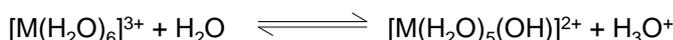
Metal aqua ions are formed in aqueous solution.

$[M(H_2O)_6]^{2+}$ , limited to M = Fe (green) and Cu (blue);  
 $[M(H_2O)_6]^{3+}$ , limited to M = Al (colourless), and Fe (violet)

In solution and Fe(III) appears yellow/brown due to hydrolysis reactions. The violet colour is only really seen in solid hydrated salts that contain these complexes.

### Acidity or hydrolysis reactions

The following equilibria happen in aqueous solutions of metal ions.



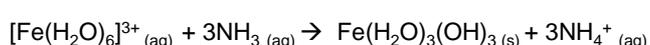
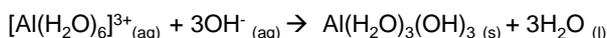
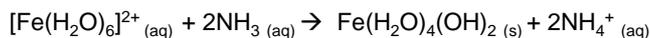
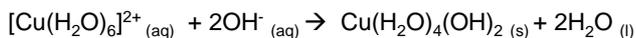
The equilibria lead to generation of acidic solutions with  $M^{3+}$  ions, and very weakly acidic solutions with  $M^{2+}$  ions. The  $3^+$  ions are noticeably more acidic.

The acidity of  $[M(H_2O)_6]^{3+}$  is greater than that of  $[M(H_2O)_6]^{2+}$  in terms of the greater polarising power (charge/size ratio) of the  $3^+$  metal ion. The greater the polarising power, the more strongly it attracts the water molecule. This weakens the O-H bond so it breaks more easily.

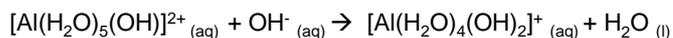
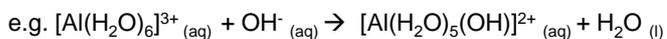
### Reaction with limited $OH^-$ and limited $NH_3$

The bases  $OH^-$  and ammonia when in limited amounts form the same hydroxide precipitates. They form in **deprotonation acid base reactions**

$M(OH)_2(H_2O)_4(s)$  : Cu blue ppt, Fe (II) green ppt  
 $M(OH)_3(H_2O)_3(s)$  : Fe(III) brown ppt, Al white ppt



This process can happen step wise removing one proton at a time. Be able to write equations for this too.

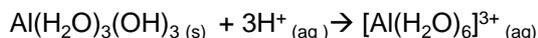
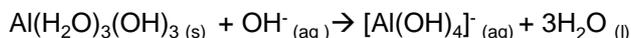


Here the  $NH_3$  and  $OH^-$  ions are acting as **Bronsted-Lowry bases** accepting a proton

### Reaction with excess $OH^-$

With excess NaOH the Al hydroxide dissolves. Al becomes  $[Al(OH)_4]^-$  (aq) colourless solution.

This hydroxides is classed as **amphoteric** because it reacts and dissolves in both acids and bases.

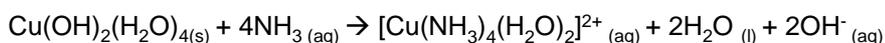


### Reaction with excess $NH_3$

With **excess  $NH_3$**  a **ligand substitution** reaction occurs with Cu and its precipitate dissolve

This substitution is incomplete with Cu.

Cu becomes  $[Cu(NH_3)_4(H_2O)_2]^{2+}$  deep blue solution

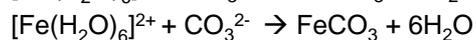
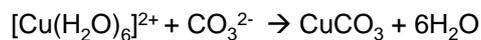
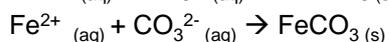
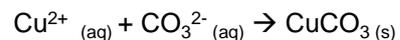


In this reactions  $NH_3$  is acting as a **Lewis base** donating an electron pair.

## Reactions with Carbonate solution

The 2+ ions react differently to the 3+ ions with carbonate solutions.

The **2+** ions with **carbonate** solution results in **MCO<sub>3</sub> ppt** being formed (Cu blue/green, Fe(II) green )



These are  
**precipitation**  
reactions

The **3+** ions with **carbonate** solution form a **M(OH)<sub>3</sub> ppt** and **CO<sub>2</sub>** gas is evolved.

Al forms white ppt of  $\text{Al}(\text{OH})_3(\text{H}_2\text{O})_3 + \text{CO}_2$

Fe(III) forms brown ppt of  $\text{Fe}(\text{OH})_3(\text{H}_2\text{O})_3 + \text{CO}_2$



These are classed as **acidity reactions**.

MCO<sub>3</sub> is formed with 2+ ions but M<sub>2</sub>(CO<sub>3</sub>)<sub>3</sub> is not formed with 3+ ions. The difference is explained by the greater polarising power of the 3+ ion due to its higher charge density.