

## Transition metals

Element Name and Symbol	Atomic Number	Common Oxidation States	Electron Configuration	
Scandium (Sc)	21	+3	Sc: [Ar] 4s <sup>2</sup> 3d <sup>1</sup>	Sc: [Ar] $\frac{1\downarrow}{4s}$ $\frac{1}{3d}$
Titanium (Ti)	22	+4	Ti: [Ar] 4s <sup>2</sup> 3d <sup>2</sup>	Ti: [Ar] $\frac{1\downarrow}{4s}$ $\frac{1\downarrow}{3d}$ $\frac{1}{3d}$
Vanadium (V)	23	+2, +3, +4, +5	V: [Ar] 4s <sup>2</sup> 3d <sup>3</sup>	V: [Ar] $\frac{1\downarrow}{4s}$ $\frac{1\downarrow}{3d}$ $\frac{1}{3d}$ $\frac{1}{3d}$
Chromium (Cr)	24	+2, +3, +6	Cr: [Ar] 4s <sup>1</sup> 3d <sup>5</sup>	Cr: [Ar] $\frac{1}{4s}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$
Manganese (Mn)	25	+2, +3, +4, +6, +7	Mn: [Ar] 4s <sup>2</sup> 3d <sup>5</sup>	Mn: [Ar] $\frac{1\downarrow}{4s}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$
Iron (Fe)	26	+2, +3	Fe: [Ar] 4s <sup>2</sup> 3d <sup>6</sup>	Fe: [Ar] $\frac{1\downarrow}{4s}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$
Cobalt (Co)	27	+2, +3	Co: [Ar] 4s <sup>2</sup> 3d <sup>6</sup>	Co: [Ar] $\frac{1\downarrow}{4s}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$
Nickel (Ni)	28	+2	Ni: [Ar] 4s <sup>2</sup> 3d <sup>7</sup>	Ni: [Ar] $\frac{1\downarrow}{4s}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$
Copper (Cu)	29	+2	Cu: [Ar] 4s <sup>1</sup> 3d <sup>10</sup>	Cu: [Ar] $\frac{1}{4s}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$
Zinc (Zn)	30	+2	Zn: [Ar] 4s <sup>2</sup> 3d <sup>10</sup>	Zn: [Ar] $\frac{1\downarrow}{4s}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$ $\frac{1\downarrow}{3d}$

A transition metal is a metal that can have at least one oxidation state with an electron in the d-orbital. According to this definition Scandium and Zinc are not transition metals because both Sc<sup>3+</sup> and Zn<sup>2+</sup> do not have electrons in their d orbitals.

### General Properties

#### Size

The size of the transition metals is more or less the same. This is due to the fact that the last shell is actually the s orbital which is being shielded by the d electrons. The increase in the electron count in the d-orbital is counteracted with an increase in the positive charge in the nucleus, and therefore the effect of the nuclear charge on the outer electrons found in the s-orbital would be more or less constant. This would result in a similar atomic radius for all the transition metals in the same period.

		Group									
		3	4	5	6	7	8	9	10	11	12
Period	4	164	147	135	129	137	126	125	125	128	137
	5	182	160	147	140	135	134	134	137	144	152
	6	172	159	147	141	137	135	136	139	144	155

- Atomic radius remains more or less constant, increasing in size only slightly.
- Ionisation energy decreases by a marginally small amount due to the constant atomic radius.

### *Metallic character*

Transition metals are metals, and therefore they have metallic bonding, giving them similar properties as all metals:

- High melting point and boiling point
- High densities
- High electrical and thermal conductivities
- High tensile strength
- Good mechanical properties

### *Oxidation state*

All transition elements lose the 2 electrons found in the s orbitals to produce a 2+ ion. Apart from this they can also lose electrons from the d-orbital, since these would have similar ionisation energies. Normally it is easier to lose unpaired electrons, although the most stable oxidation states are the +3 for the transition metals on the left and the +2 for the transition elements on the right. Manganese can have a maximum oxidation state of +7 since it is  $4s^2 3d^5$  and therefore it has 5 unpaired electrons in the d-orbitals.

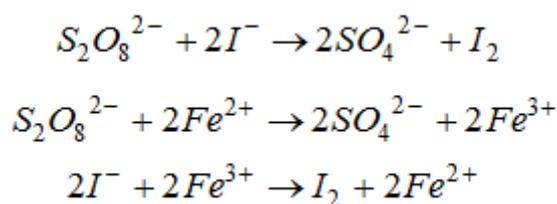
### *Catalytic properties*

#### Heterogeneous catalysts

A heterogenous catalyst is a catalyst that is in a different state to the reactants, for example a solid in a solution mixture. This normally offers an adsorption surface where the reaction can take place, just like Iron in the preparation of ammonia.

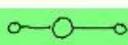
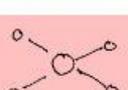
### Homogenous catalysts

A homogenous catalyst is one which is in the same state as the reaction mixture. In this case transition metals can use their variable oxidation states. One such reaction is the reaction between iodide ions and thiosulfate.



### Complexes

Complex compounds are produced between transition metals or ions by receiving electrons from a ligand via a dative bond. The ligand can be a neutral molecule or a negatively charged ion.

Coordination Number	Shape	Diff ligands + diff Metal ions Diff complex ion	Complex ion	oxidation state on metal ion	overall charge on complex ion	Ligand (charge)
2	Linear 	$[CuCl_2]^-$	$[CuCl_2]^-$	+1	-1	Cl (-)
		$[Ag(NH_3)_2]^+$	$[Ag(NH_3)_2]^+$	+1	+1	NH <sub>3</sub> (0)
		$[Ag(CN)_2]^-$	$[Ag(CN)_2]^-$	+1	-1	CN (-)
4	Square Planar 	$[CuCl_4]^{2-}$	$[CuCl_4]^{2-}$	+2	-2	Cl (-)
		$[Cu(NH_3)_4]^{2+}$	$[Cu(NH_3)_4]^{2+}$	+2	+2	NH <sub>3</sub> (0)
		$[CoCl_4]^{2-}$	$[CoCl_4]^{2-}$	+2	-2	Cl (-)
		$[NiCl_4]^{2-}$	$[NiCl_4]^{2-}$	+2	-2	Cl (-)
4	Tetrahedral 	$[Zn(NH_3)_4]^{2+}$	$[Zn(NH_3)_4]^{2+}$	+2	+2	NH <sub>3</sub> (0)
		$[MnCl_4]^{2-}$	$[MnCl_4]^{2-}$	+2	-2	Cl (-)
6	Octahedral 	$[Cu(H_2O)_6]^{2+}$	$[Cu(H_2O)_6]^{2+}$	+2	+2	H <sub>2</sub> O (0)
		$[Fe(OH)_3(H_2O)_3]$	$[Fe(OH)_3(H_2O)_3]$	+3	0	OH (-) H <sub>2</sub> O (0)
		$[Fe(CN)_6]^{3-}$	$[Fe(CN)_6]^{3-}$	+3	-3	CN (-)
		$[Cr(NH_3)_4Cl_2]^+$	$[Cr(NH_3)_4Cl_2]^+$	+3	+1	NH <sub>3</sub> (0) Cl (-)

There are two types of ligands:

Monodentate: can only form 1 dative bond with the metal ion

Polydentate: can form more than one dative bond with the metal ion

### Writing the formula

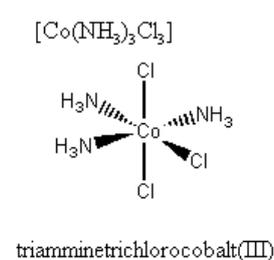
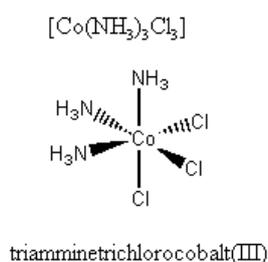
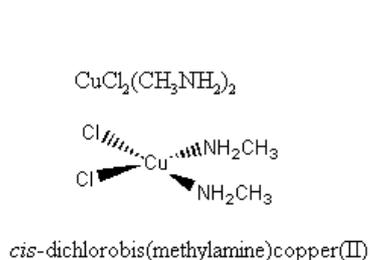
The formula is written in a specific order:

1. Symbol of metal ion
2. Formula of negatively charged ligands
3. Formula of neutral ligands.
4. If there is an overall charge the complex is written inside a square bracket.

### Naming complexes

A complex should be named as follows:

1. Ligands in alphabetical order
2. Name of metal
  - a. Normal name if complex is positive or neutral
  - b. Latin name ending in -ate if compound is negative
3. Oxidation number of the metal

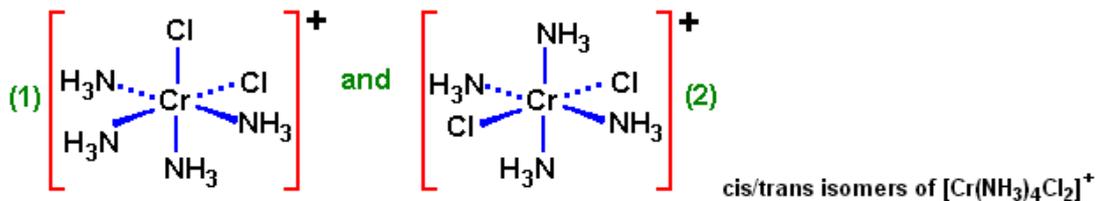
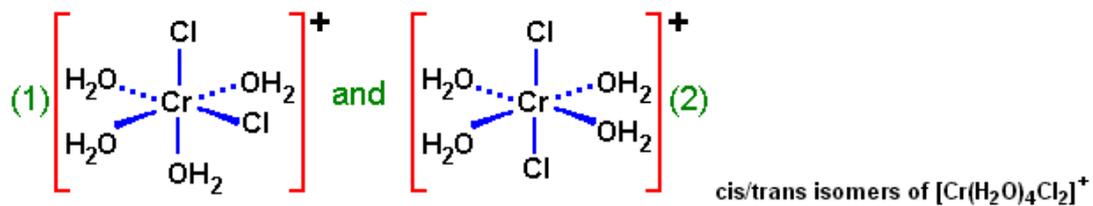
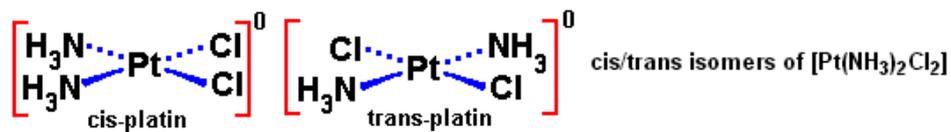


### Shapes

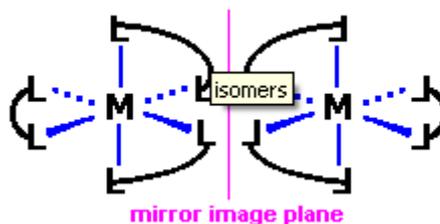
Complexes can have a co-ordination number of 2, 4 or 6. A co-ordination number of 2 would give rise to a linear molecule, a co-ordination number of 4 can be either tetrahedral or square planar while co-ordination number of 6 gives rise to an octahedral complex.

Isomerism

Cis/trans isomerism



Chirality



Ligand groups and anions

Chromium(III) chlorides display the somewhat unusual property of existing in a number of distinct hydrates, forming a series of  $[\text{CrCl}_{3-n}(\text{H}_2\text{O})_n]^{2+}$ . The main hexahydrate can be more precisely described as  $[\text{CrCl}_2(\text{H}_2\text{O})_4]\text{Cl} \cdot 2\text{H}_2\text{O}$ . It consists of the cation  $[\text{CrCl}_2(\text{H}_2\text{O})_4]^+$  and additional molecules of water and a chloride anion in the lattice. Two other hydrates are known, pale green  $[\text{CrCl}(\text{H}_2\text{O})_5]\text{Cl}_2 \cdot \text{H}_2\text{O}$  and violet  $[\text{Cr}(\text{H}_2\text{O})_6]\text{Cl}_3$ .

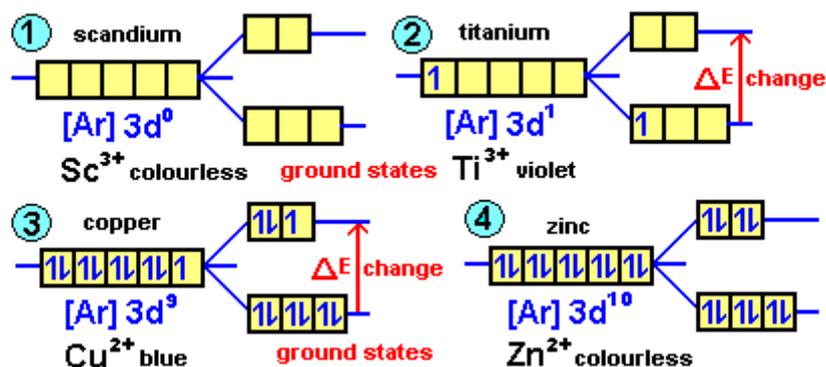
*Coloured compounds*

When transition metals form ligands the overlapping of the orbitals do not remain degenerate, producing 2 different energy states. This is due to the interaction between the orbitals of the ligands with the d-orbital of the transition metal.

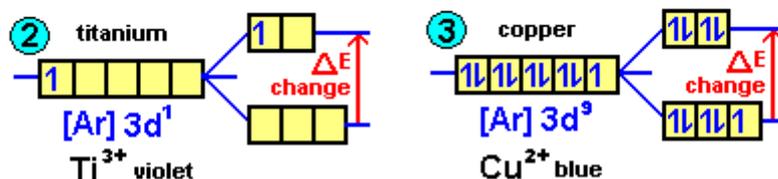
Electrons can be excited from the lower energy state to the higher energy state, and when the electron de-excites itself a wavelength is given off which corresponds to the wavelengths of visible

light. The colour given off depends on the energy difference between the two states, which in turn depends on:

- The central metal or ion
- The ligands.



The electronically excited states of titanium(III) and copper(II) are illustrated below.



## Chromium

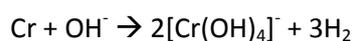
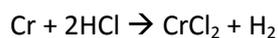
### General Properties

Electronic configuration: [Ar] 4s<sup>1</sup> 3d<sup>5</sup>

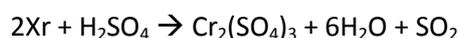
Oxidation states: +2, +3, +6

### Amphoteric properties

Chromium can react with both acids and bases



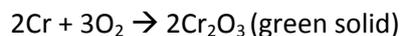
The reaction of Chromium with sulphuric acid is a redox reaction in which the sulphuric acid produces SO<sub>2</sub> while the chromium is oxidised to the +3 oxidation state.



On the other hand it will not react with nitric acid since this will passivate the metal, making it unreactive.



Chromium burns in air to produce the oxide.



This oxide dissolves in hydrochloric acid to give the hexaqua chromium (III) with the chloride ions as the balancing anions (giving rise to the isomerism as discussed previously)

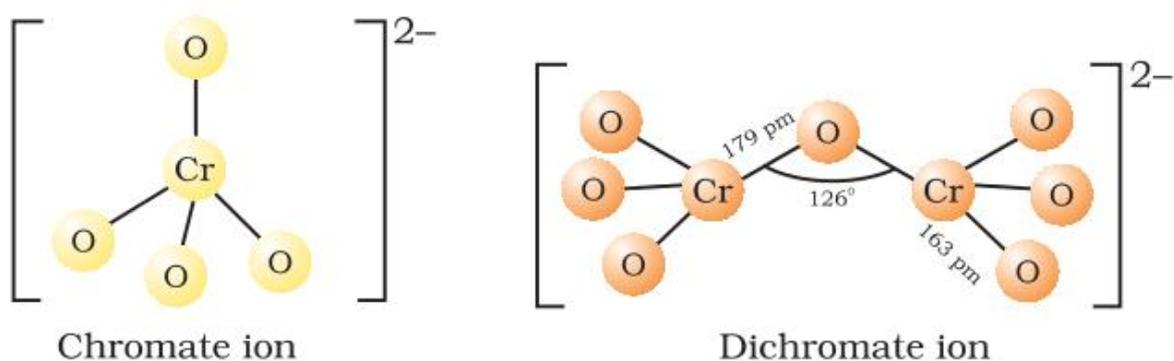
These compounds can be differentiated by the amount of AgCl that is precipitated on the addition of AgNO<sub>3</sub>. The Cl<sup>-</sup> in the complex will not be precipitated.

The hexaqua chromium (III) is acidic since the Chromium will pull the electrons towards it making one of the waters to lose an H<sup>+</sup>



Chromium (III) can be oxidised to Chromium (VI) using a strong oxidising agent like the peroxide. Chromate are mostly known for their ability to act as oxidising agents, in which they are reduced to chromium (III).

Chromium (VI) can exist in two forms, CrO<sub>4</sub><sup>2-</sup> or Cr<sub>2</sub>O<sub>7</sub><sup>2-</sup> with the former being found in alkaline conditions and the latter being found in acidic conditions.

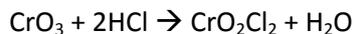


Oxide

The oxide is prepared by the action of concentrated sulfuric acid on the salt.



This oxide can then react with hydrochloric acid to create  $\text{CrO}_2\text{Cl}_2$  which is an oxidising agent for the reaction of methylbenzene to benzaldehyde.



## Manganese

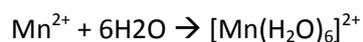
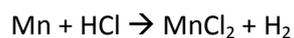
### General Properties

Electronic configuration:  $[\text{Ar}] 4s^2 3d^5$

Oxidation states: +2, +3, +6

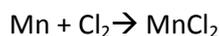
### Reaction with acids

The metal reacts with acids to form the  $\text{Mn}^{2+}$  ion. The ion produced then forms very stable hexaaqua complexes, which are slightly pink in colour. This is a very stable complex.



### Reaction with halogens

Reacting Manganese with halogens gives the  $\text{Mn}^{2+}$  ion.



### Oxide

The oxide of manganese produced in combustion with air depends on the amount of Oxygen produced.

The reaction pathway goes like this:

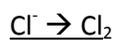
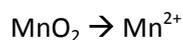


With the Mn (IV) being the most stable)

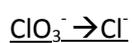
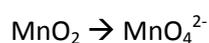
### $\text{Mn}^{4+}$

MnO<sub>2</sub> can act as both an oxidising agent and a reducing agent, since it can be oxidised to Manganese (VII) while it can also be reduced to Mn<sup>2+</sup>.

*Oxidising agent:*



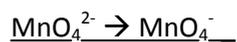
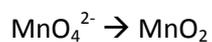
*Reducing agent:*



MnO<sub>2</sub> is also oxidised by heating in alkali in contact with air.

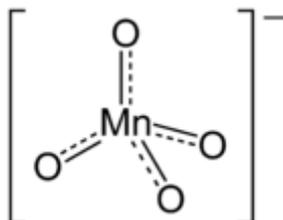
**Mn<sup>6+</sup>**

The manganite (VI) ion disproportionates to give Manganese (VII) and Manganese (IV)



**Mn<sup>7+</sup>**

Manganese (VII) is a very good oxidising agent being itself reduced to Manganese (II)



**Test for Mn<sup>2+</sup>**

Mn<sup>2+</sup> can be oxidised to MnO<sub>4</sub><sup>-</sup>

## Iron

### General Properties

Electronic configuration: [Ar] 4s<sup>2</sup> 3d<sup>6</sup>

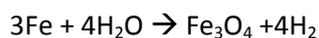
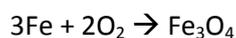
Oxidation states: +2, +3

### Rust

Rust is the reaction of iron, water and oxygen. All 3 are needed in order for rust to take place.

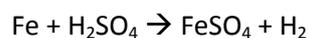
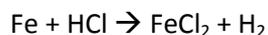
### Mixed oxide

When iron is reacted with oxygen on heating or steam the mixed oxide is formed.



### Reaction with acid

Iron reacts with acids to form Iron(II) compounds since these are not very strong oxidising agents.



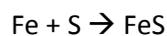
Both of these salts are hydrated. Preparation of the anhydrous chloride by heating is not possible, while preparation of the sulphate is. This is due to the fact that the chloride would actually react with the hydroxide.



### Fe<sup>2+</sup>

#### FeS

Prepared by direct synthesis



This happens because S<sup>2-</sup> is a strong reducing agent and it will reduce Fe<sup>3+</sup> to Fe<sup>2+</sup>.

#### Fe(OH)<sub>2</sub>

Prepared by the addition of  $\text{OH}^-$  to any Iron(II) salt.

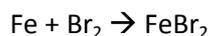


*FeX<sub>2</sub>*

All the halides can be produced by the reaction of the iron with the acid.



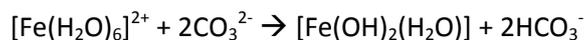
The bromide and the iodide can also be produced direct combination.



The iron should be in excess to ensure that no  $\text{FeBr}_3$  is formed.

*FeCO<sub>3</sub>*

Iron is acidic and dissolution of iron carbonate in water would give rise to iron hydroxide and the hydrogencarbonate.



*FeSO<sub>4</sub>*

Iron(II) sulphate is prepared by the reaction of iron with sulfuric acid.

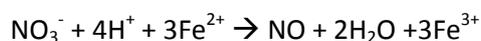
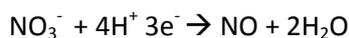
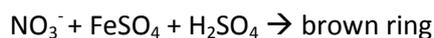
The hydrated crystals are green and turn white when the water of crystallisation is removed.

On strong heating iron (III) oxide is produced together with a mixture of  $\text{SO}_2$  and  $\text{SO}_3$



*Brown ring test*

This is a test for nitrates.



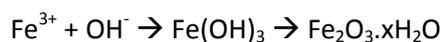
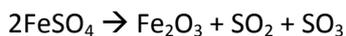
NO can act as a ligand and it will displace one of the water ligands on the iron, and this would form the brown ring visible in the test.



## **Fe<sup>3+</sup>**

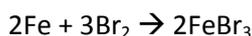
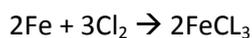
### *Oxide*

The oxide is prepared by the decomposition of Iron (II) sulphate or by the reaction of iron (III) with hydroxide ions.



### *Halides*

The chloride and bromide are prepared by direct combination with the halogens. These are both collected and purified by sublimation.



It is noted that Bromine has to be in excess in order to ensure that FeBr<sub>3</sub> is formed.

### *Thiocyanate*

Thiocyanate and Fe<sup>3+</sup> can form a complex with a deep blood red colouring. The thiocyanate will displace a water molecule to form the complex.



## **Copper**

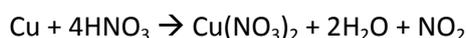
### **General Properties**

Electronic configuration: [Ar] 4s<sup>1</sup> 3d<sup>10</sup>

Oxidation states: +1, +2

### **Reactions of Copper**

It reacts with strong oxidising agents such as HNO<sub>3</sub> and H<sub>2</sub>SO<sub>4</sub> (both concentrated) to give nitrogen dioxide and sulfur dioxide respectively.



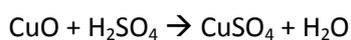


### **Cu<sup>2+</sup>**

Copper (II) oxide can be prepared from the decomposition of copper carbonate, copper nitrate or copper hydroxide.



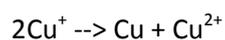
This oxide is basic and therefore it can be used to produce salts by reacting with the respective acids.



### **Cu<sup>+</sup>**

Less stable than copper (II)

Copper (I) disproportionates to give copper and Copper (II)



Iodine and iodide only form the copper (I) salts.

