



# Reactions

So Many Reactions  
Such Little Time

# Same as Before, But Different

- Name types of reactions
  - In “regular chemistry”
    - Single Replacement
    - Double Replacement
    - Composition, Synthesis
    - Decomposition
    - Combustion

# Solubility Rules

for Common Ionic Compounds in Water

You need to memorize solubility rules.

1. Most nitrates are soluble.
2. Most salts containing alkali metal ions and ammonium are soluble  
 $\text{Li}^{+1}, \text{Na}^{+1}, \text{K}^{+1}, \text{Cs}^{+1}, \text{NH}_4^{+1}$
3. Most chloride, bromide, and iodide salts are soluble  
Notable exceptions are  $\text{Hg}_2^{+2}, \text{Ag}^{+2}, \text{Pb}^{+2}$
4. Most sulfate ( $\text{SO}_4$ ) salts are soluble  
Notable exceptions are  $\text{Ca}^{+2}, \text{Ba}^{+2}, \text{Pb}^{+2}, \text{Hg}_2^{+2}$
5. Most hydroxides are only slightly soluble.  $\text{Na}^{+}$  and  $\text{K}^{+}$  are soluble.  
 $\text{Ba}^{+2}, \text{Sr}^{+2}$  and  $\text{Ca}^{+2}$  are marginally soluble
6. Most sulfide ( $\text{S}^{-2}$ ), carbonate ( $\text{CO}_3^{-2}$ ), chromate ( $\text{CrO}_4^{-2}$ ) and phosphate ( $\text{PO}_4^{-3}$ ) are only slightly soluble.

# Thank's Luck for the Handout

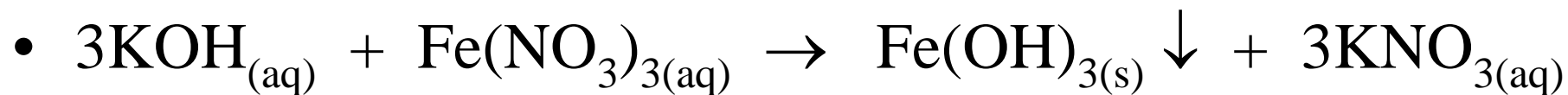


# Reactions in Solution

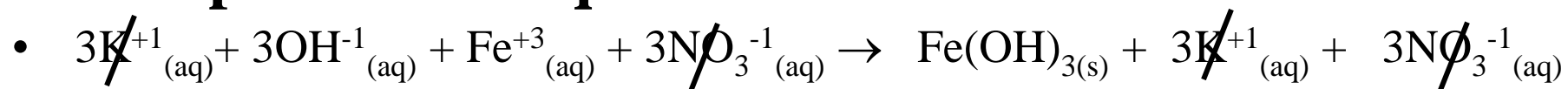
- $\text{KOH}_{(\text{aq})} + \text{Fe}(\text{NO}_3)_3_{(\text{aq})} \rightarrow \text{Fe}(\text{OH})_3_{(\text{s})} + 3\text{KNO}_3_{(\text{aq})}$
- Aq = aqueous or soluble in water.
- S = solid or insoluble, it is the precipitate.

# How do you write equations?

## Molecular Equations

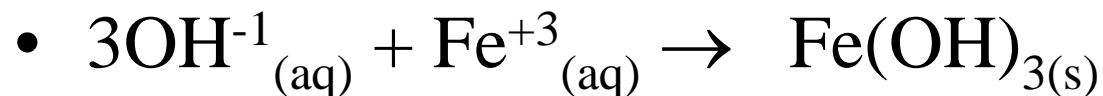


## Complete Ionic Equations



You do not need the (aq) it is understood.

## Net Ionic Equations (Use this form on AP Exam)



- Spectator ions do not react (change)

- Not included in the Net Ionic Equation

# Write the Net Ionic Equation

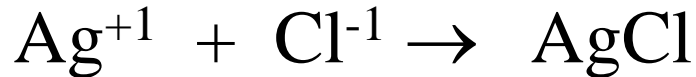
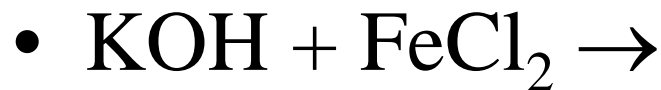
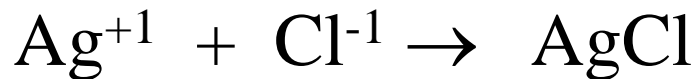
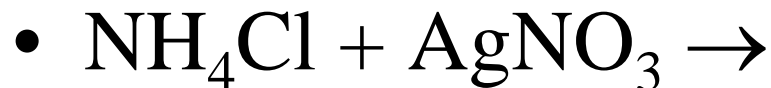
- Get out your periodic table

## **Steps to write Net Ionic Equations**

- Determine the reaction
- Evaluate the spectators
- Write the equation

# Write the Net Ionic Equation

*All Precipitate Reactions*



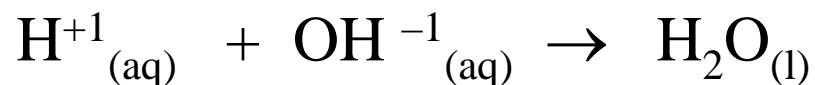


# Net Ionic Equations

## Acid base Reactions

- Acid donates a proton ( $\text{H}^{+1}$ ) and Base accepts a proton.

(Bronsted-Lowry definition of an acid)



- Strong acids are highly dissociated

HCl is written  $\text{H}^{+} + \text{Cl}^{-}$

(assuming both ions react!)

- Weak acids are “associated”

Acetic Acid is written  $\text{HC}_2\text{H}_3\text{O}_2$  ,

# How do you know an acid is strong or weak?

- You have memorized the strong acids
  - $\text{H}_2\text{SO}_4$
  - $\text{HCl}$ ,  $\text{HBr}$ ,  $\text{HI}$
  - $\text{HClO}_4$
  - $\text{HNO}_3$
- All the other acids are weak electrolytes

# Common Misconceptions

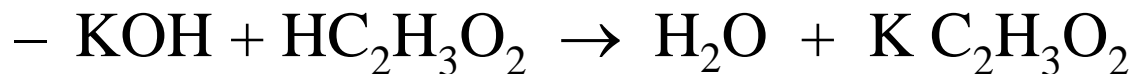
- All weak acids are not soluble in water
  - Acid solubility is not related to solubility
  - It has to do with the basicity of the conjugate base
- Weak acids are dilute solutions
  - Concentration is not related to acid strength
  - There can be 98 % solutions of sulfuric and acetic acid

# Net Ionic Equations

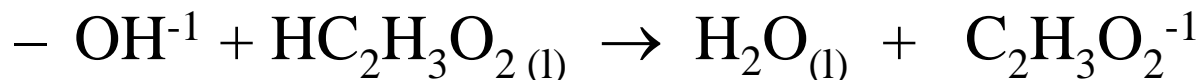
## Acid base Reactions

Acetic acid is reacted with a solution of potassium hydroxide

- **Molecular equation:**



- **Net Ionic Equation**



- **Why is Acetic acid written out?**

- It is a weak electrolyte, so it exists as  $\text{HC}_2\text{H}_3\text{O}_{2(l)}$
  - KOH is a strong electrolyte, it is completely ionized.

# Redox reactions

- Many chemical reactions involve a transfer of electrons between some of the atoms in a reaction. These are called **redox reactions**.
- The reactions must be properly balanced, so that there is no net gain or loss of electrons.

# Introduction to Redox Reactions

- **Oxidation** is the loss of electrons.  
Oxidized: atom, molecule, or ion becomes more positively charged.
- **Reduction** is the gain of electrons  
Reduced: atom, molecule, or ion becomes less positively charged.

# Memory Devices

– *Oxidation or Reduction???*

– **LEO the lion goes GER**

– Lose Electrons Oxidation

– Gain Electrons Reduction

– **OIL RIG**

– Oxidation Is Loss

– Reduction Is Gain

# Balancing a Molecular Redox Equation



- Assign oxidation numbers to determine which substance(s) are oxidized and which are reduced.
- Adjust coefficients to balance the atoms of the elements that are being oxidized and reduced.
- Adjust the coefficients of the oxidized and reduced substances so that the total # electrons lost equals the total # electrons gained.
- Adjust coefficients to balance the atoms of all other elements by inspection.



# KMnO<sub>4</sub> Redox Molecular Equation

• *Now you try it!*



# Steps for Balancing Ionic Redox Reactions

- Write the **net ionic equation** with the oxidation numbers.
- Break equation apart into **half-reactions**.
- **Change coefficients** to balance elements *other than* oxygen or hydrogen
- Balance oxygen by adding **H<sub>2</sub>O**
- Balance hydrogen by adding **H<sup>+</sup>**

# Steps for Balancing Ionic Redox Reactions

- Balance **electric charge** on both sides by adding electrons
- Multiply each half-reaction by some number to make number of **electrons transferred equal**
- **Combine** half-reactions and **cancel**
- If in **basic** solution, add one  $\text{OH}^-$  *to each side* for each  $\text{H}^+$  and simplify water.

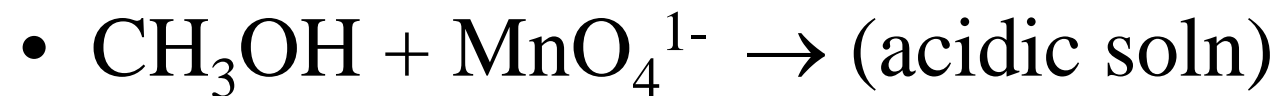
## Balance an Ionic Redox Equation

- $\text{Cr}_2\text{O}_7^{2-} + \text{Cl}^{1-} \rightarrow$  (acidic solution)

## Balance an Ionic Redox Equation

- $\text{Cr}_2\text{O}_7^{2-} + \text{Cl}^{1-} \rightarrow$  (basic solution)

# Balance an Ionic Redox Equation



# Synthesis Reactions

- Find an example of each type in your packet -
  - a. metal + nonmetal  $\rightarrow$  binary salt
  - b. metal + oxygen  $\rightarrow$  metal oxide
  - c. nonmetal + oxygen  $\rightarrow$  nonmetal oxide
  - d. metal oxide + nonmetal oxide  $\rightarrow$  ternary salt
  - e. metal oxide + water  $\rightarrow$  base
  - f. nonmetal oxide + water  $\rightarrow$  oxyacid
  - g. ammonia + acid  $\rightarrow$  ammonium salt

# Decomposition

- a. binary salt or binary oxide  $\rightarrow$  elements
- b. metal carbonates  $\rightarrow$  metal oxide + carbon dioxide
- c. metal hydrogen carbonates  $\rightarrow$  metal oxide +  
carbon dioxide + water
- d. metal hydroxide  $\rightarrow$  metal oxide + water
- e. oxyacids  $\rightarrow$  nonmetal oxide + water
- f. metal chlorates  $\rightarrow$  salt + oxygen



# Single Replacement and Redox

- a. metal replaces metal
- b. metal replaces H from water or acid
- c. halogen replaces halide

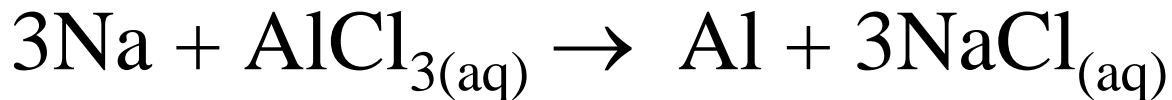
Use activity series

- d. An oxidizing agent in presence acid (or base) forms water, oxidizes whatever else is in solution and the reduced products (know products from classic oxidizing agents)

Li  
K  
Ba  
Sr  
Ca  
Na  
Mg  
Al  
Mn  
Zn  
Fe  
Cd  
Co  
Ni  
Sn  
Pb  
H  
Cu  
Ag  
Hg  
Au

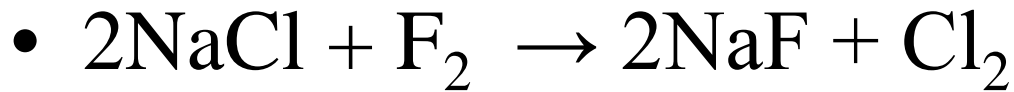
# Single Replacement Reactions

- Remember the Activity Series!

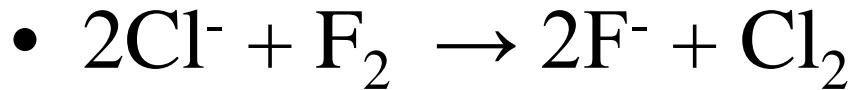


Which one occurs for real? Why?

# Single Replacement



Or



F

Cl

Br

I

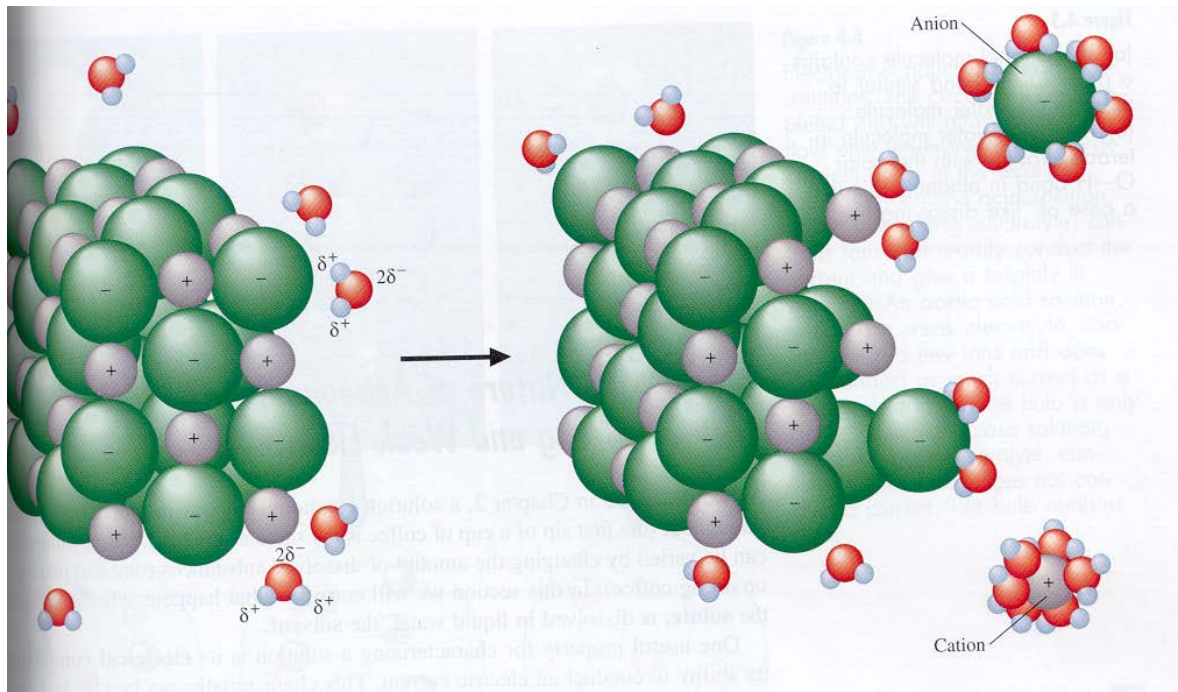
- The activity of the halides is in their order on the periodic table

# Double Replacement

- a. formation of a precipitate
- b. acid/base reaction
  - a. forms a molecule like water or weak acid

# Precipitation Reactions

- What is a precipitate?
  - Rain is called precipitation, Why?
- When an ionic compound dissolves in water, what actually happens?



# Precipitation Reactions

## Double Replacement

- $\text{K}_2\text{CrO}_{4(\text{aq})} + \text{Ba}(\text{NO}_3)_{2(\text{aq})} \rightarrow \text{Products}$ 
  - A yellow precipitate forms. What is it?
- Consider what is in solution

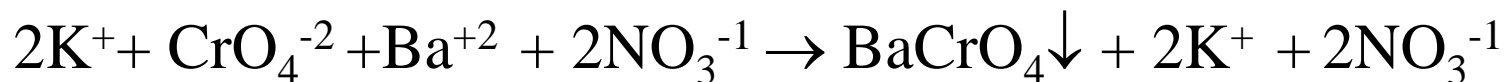
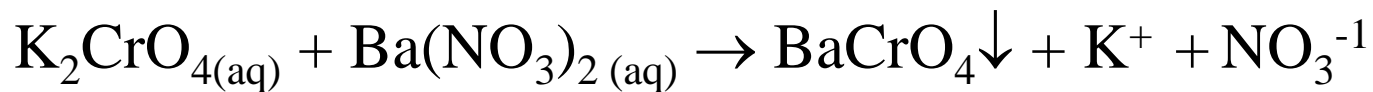


- Only two choices to precipitate



# Double Replacement Reactions

- Why does one replace the other?
- Need to know the solubility of compounds to predict these.
- Yes, you must memorize them!



# Combustion of a Hydrocarbon

- Produces carbon dioxide and water



## ● *The Structure of Complex Ions*

- Ions of transition metals lose electrons to form small, highly charged cations.
- These positive ions attract negatively charged anions and/or the negative end of polar molecules, called *ligands*, to form coordinate covalent bonds, where both electrons from the shared pair come from the anion.
- The resulting complex ions are charged: They have a metal ion (can be zero oxidation state) bonded to a number of ligands. Example,  $[\text{Ag}(\text{NH}_3)_2]^+$ .

# • *The Structure of Complex Ions*

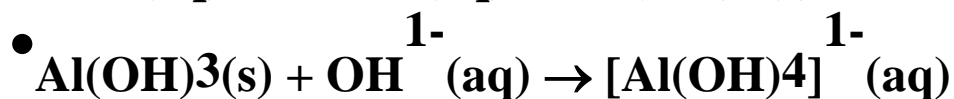
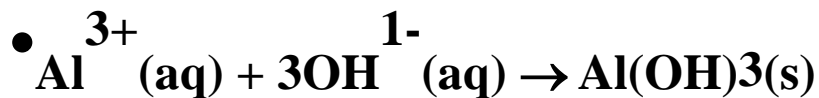
- The formation of complex ions can occur when insoluble compounds react with ligands to dissolve in solution.

- Examples:

• **Excess concentrated ammonia solution is added to a suspension of silver chloride.**

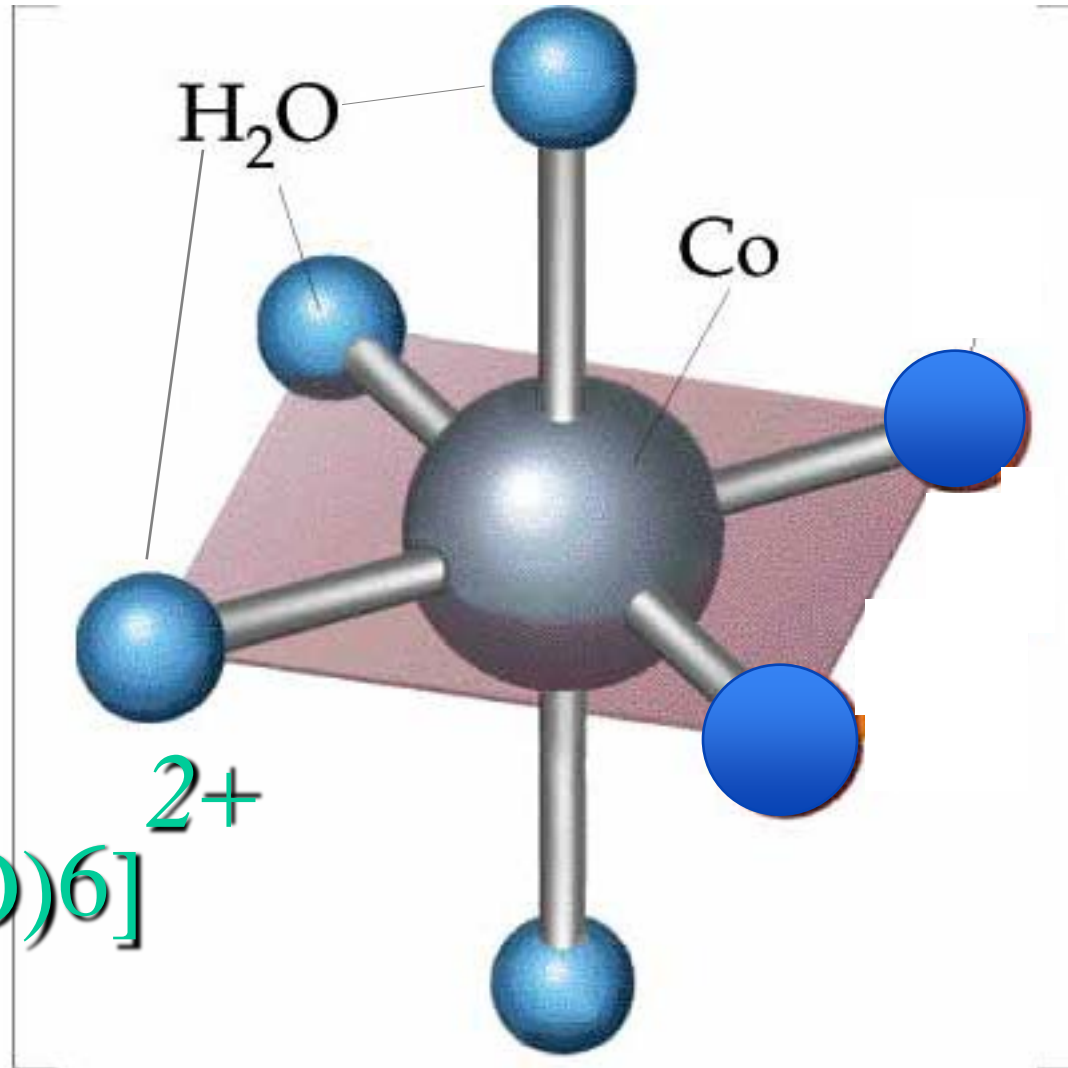


- **Excess potassium hydroxide solution is added to a solution of aluminum nitrate.**

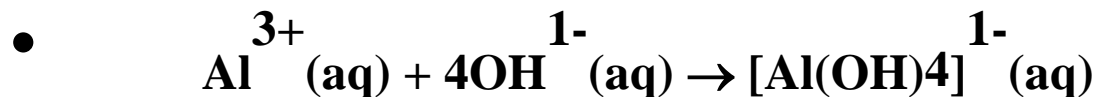


# • *The Structure of Complex Ions*

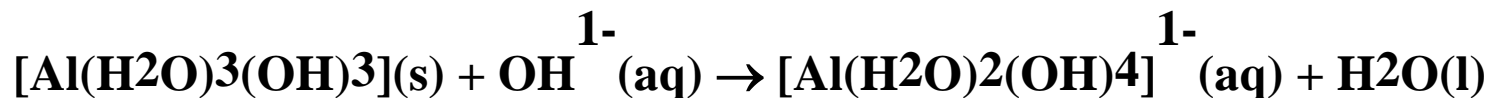
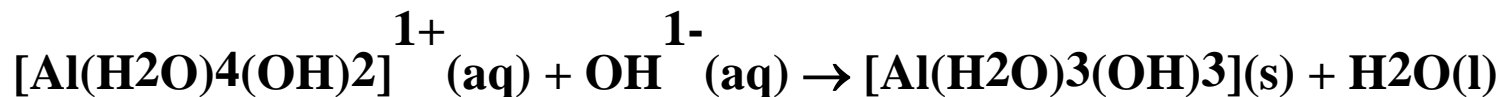
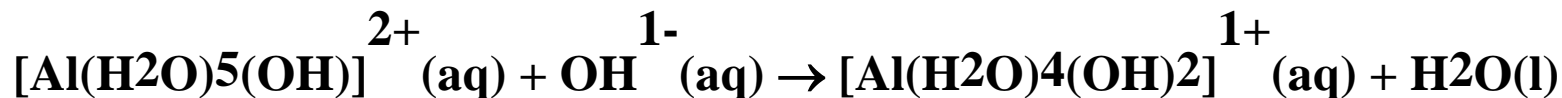
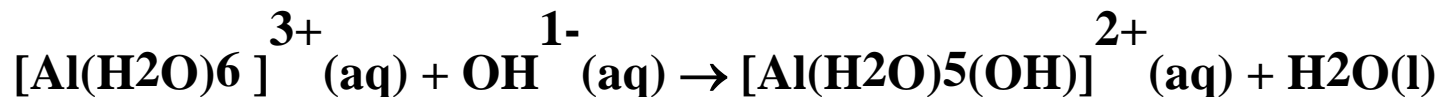
- Most metal ions in water exist as  $[M(H_2O)_6]^{n+}$ .



## • *The Structure of Complex Ions*

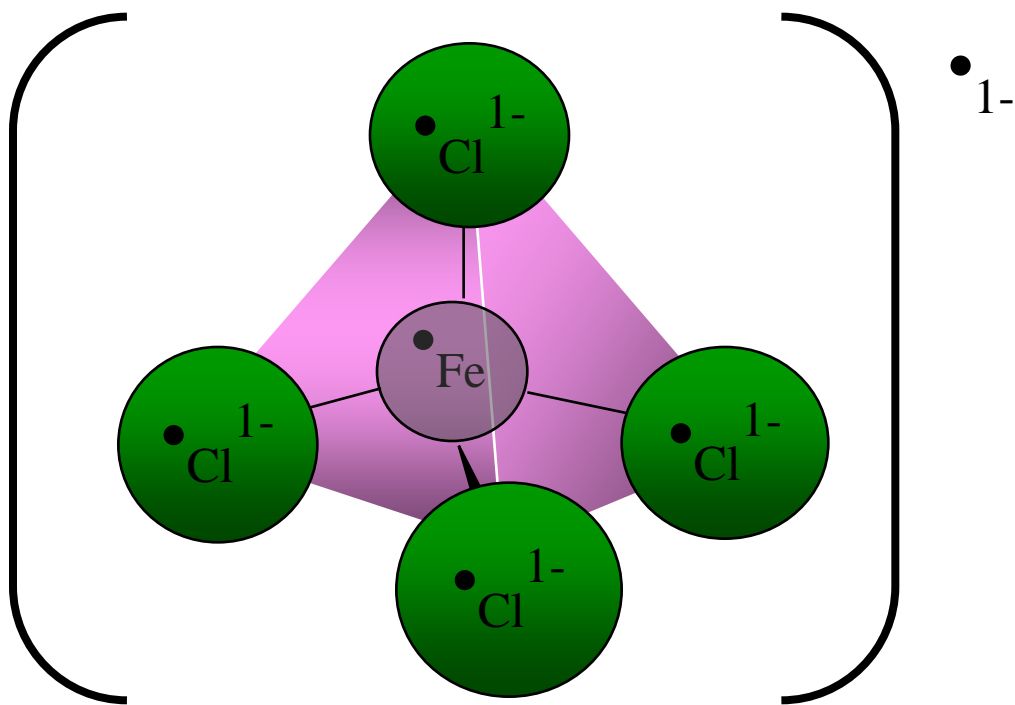


- The reaction of aluminum and hydroxide ions actually occurs in several steps:



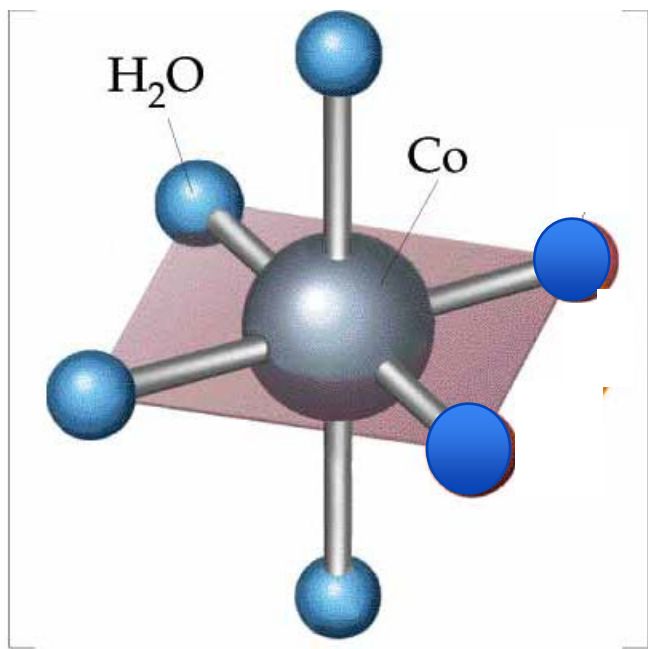
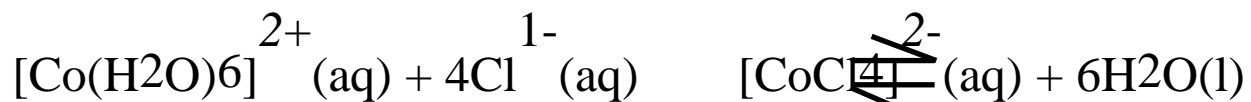
## • *The Structure of Complex Ions*

- The number of ligands attached to the metal is called the *coordination number*.
- The size of the ligand affects the coordination number (e.g.  $[\text{FeF}_6]^{3-}$  forms but only  $[\text{FeCl}_4]^{-}$  is stable).



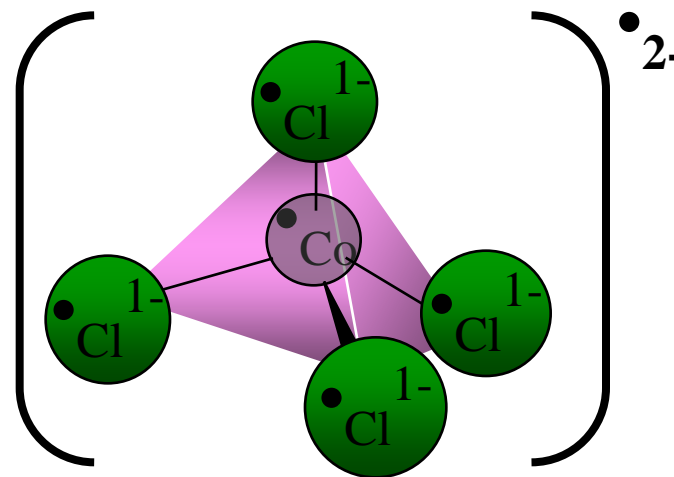
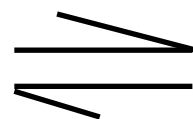
# The Structure of Complex Ions

Concentrated hydrochloric acid is added to a solution of cobalt (II) chloride.



pink color

• 2+

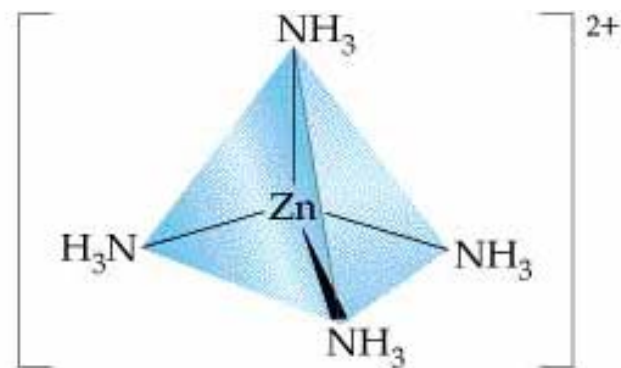


blue color

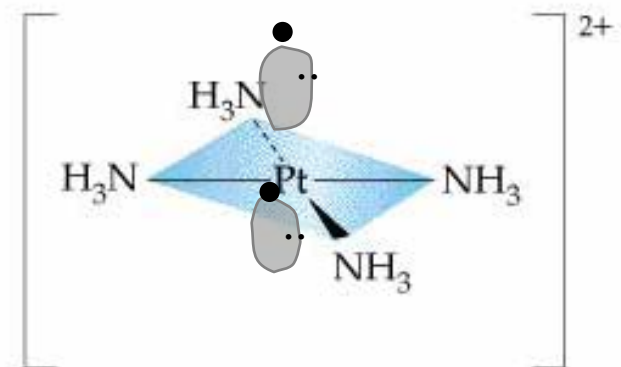
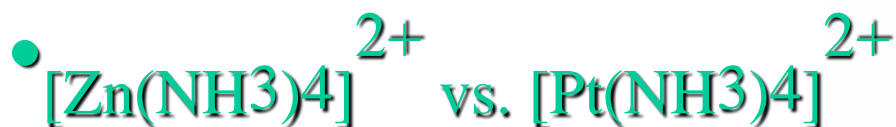
# The Structure of Complex Ions

- As a general rule, the most stable complexes have a coordination number that is twice the charge of the metallic ion.

- Complexes with a coordination number of 4 are either tetrahedral or square planar (commonly seen for  $d$  metal ions).



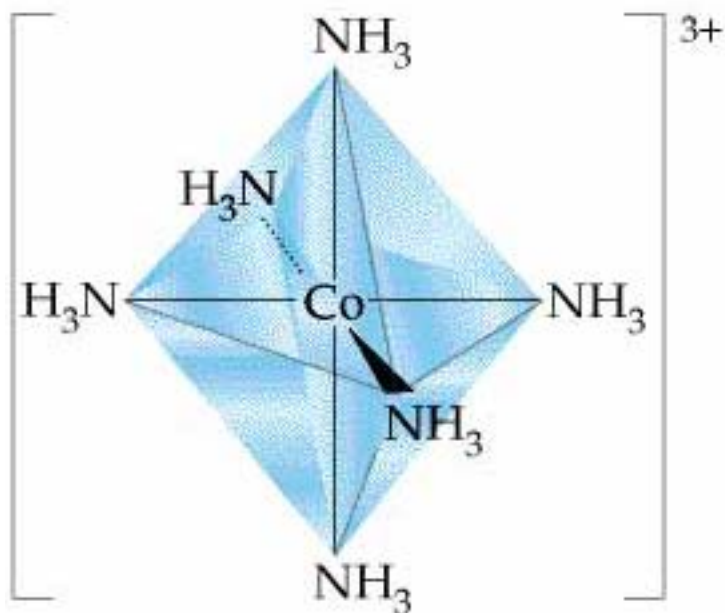
(a)



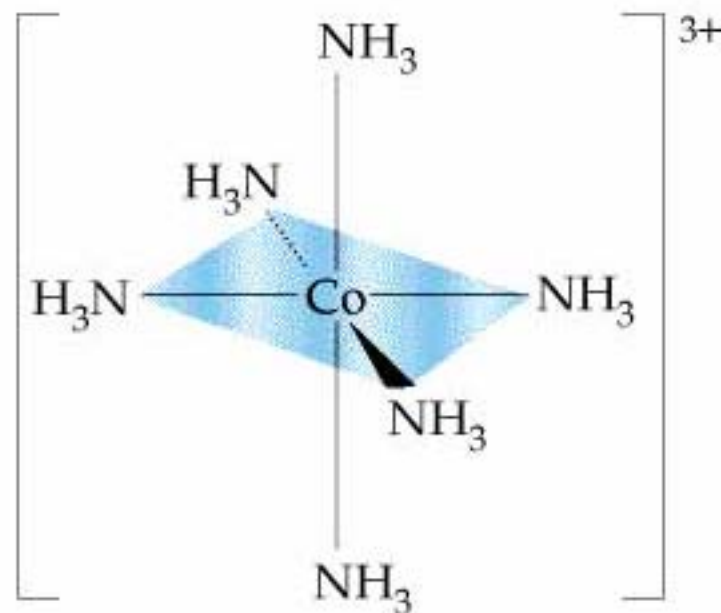
(b)

## • *The Structure of Complex Ions*

- Complexes with a coordination number of 6 form octahedral structures.



(a)



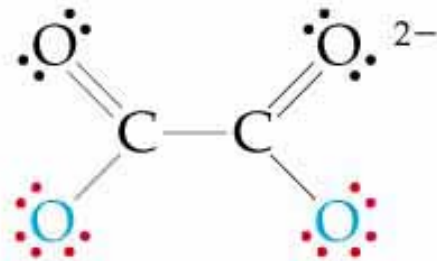
(b)

- The relative strength of the bonds between the metallic ions and common ligands is:

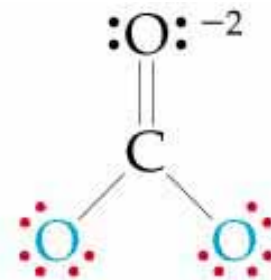




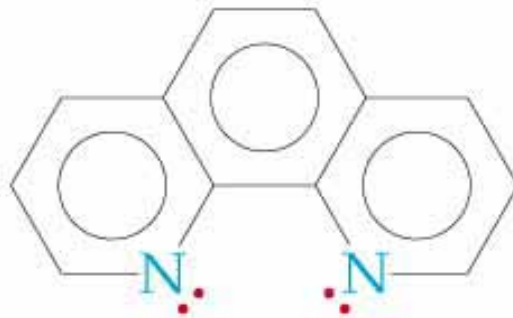
- Polydentate ligands (or chelating agents) bind to the metallic ion through more than one donor atom per ligand.



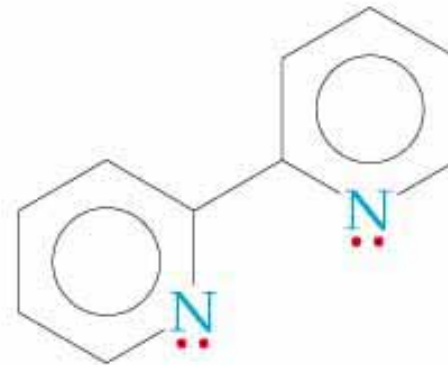
Oxalate ion



Carbonate ion



*Ortho*-phenanthroline  
(*o*-phen)

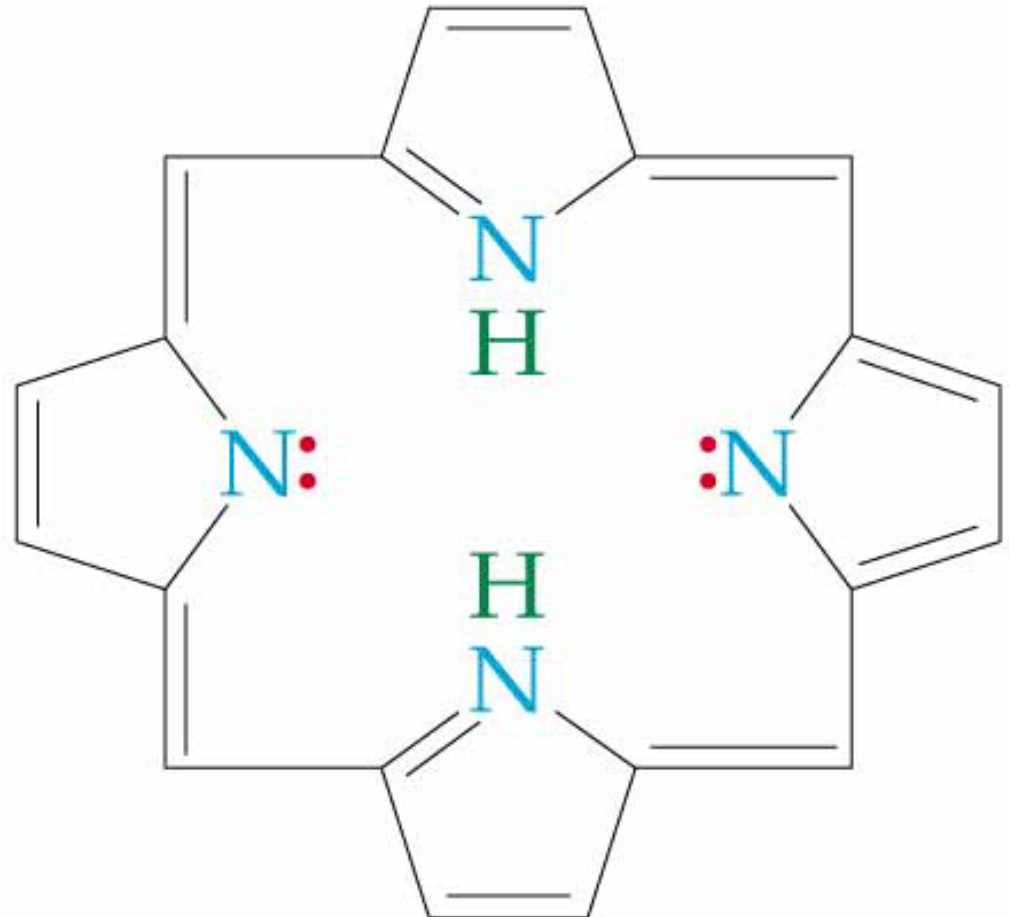


Bipyridine  
(bipy)

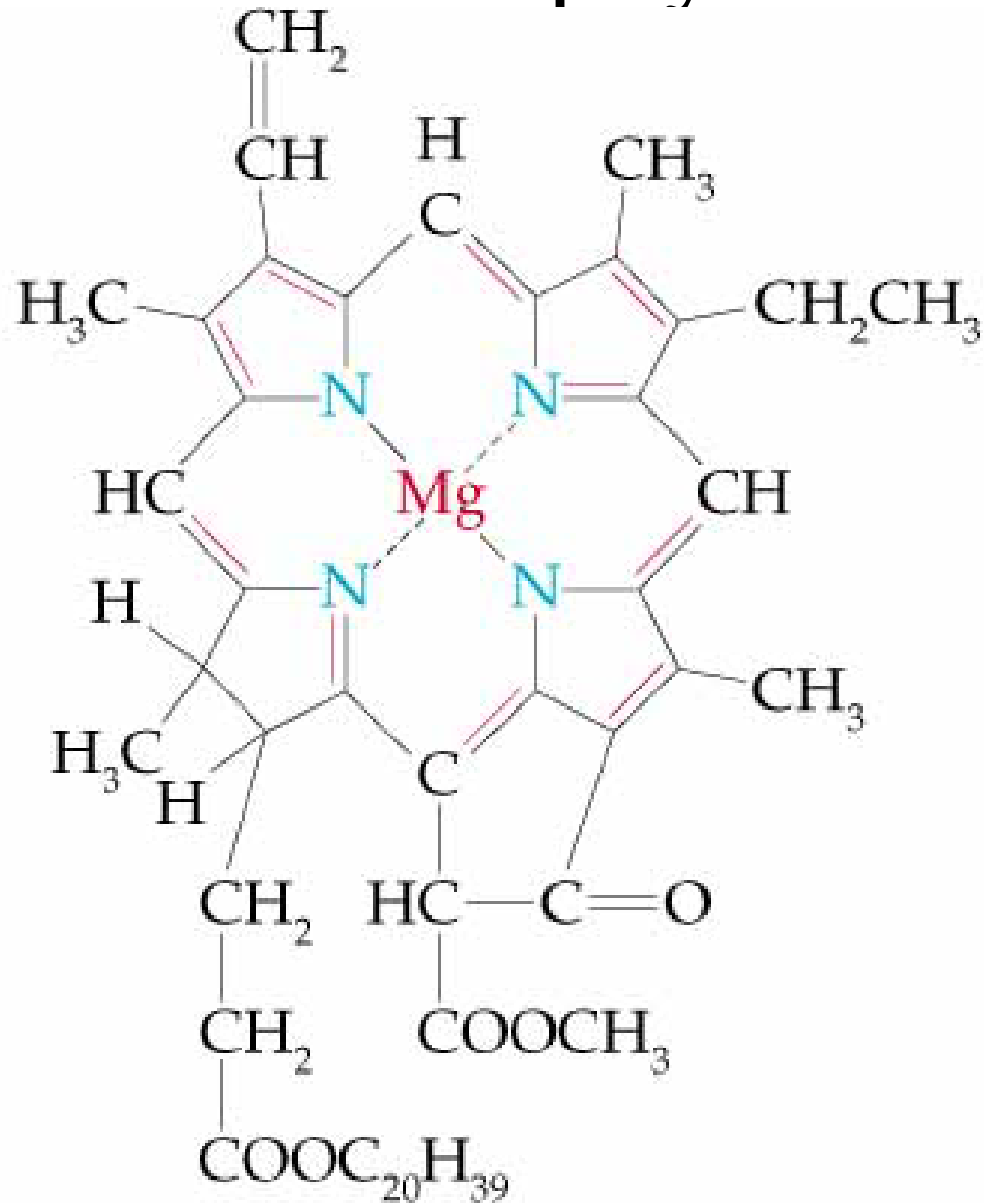
# ● *Metals and Chelates in Living Systems*

- Many natural chelates are designed around the *porphyrin* molecule.

- After the two H atoms bound to N are lost, porphyrin is a tetradentate ligand.



# Chlorophyll a





## Nomenclature

- Rules: (in case you wanted to know)

- For salts, name the cation before the anion. Example in  $[\text{Co}(\text{NH}_3)_5\text{Cl}]\text{Cl}_2$  we name  $[\text{Co}(\text{NH}_3)_5\text{Cl}]$  before the chloride,  $\text{Cl}$ .

- Within a complex ion, the ligands are named (in alphabetical order) before the metal. Example:  $[\text{Co}(\text{NH}_3)_5\text{Cl}]$  is pentaamminechlorocobalt(III). Note the penta portion is an indication of the number of  $\text{NH}_3$  groups and is therefore not considered in the alphabetizing of the ligands.

- Anionic ligands end in o and neutral ligands are simply the name of the molecule. Exceptions:  $\text{H}_2\text{O}$  (aqua) and  $\text{NH}_3$  (ammine).

