Concepts of Chemical Bonding and Molecular Geometry
Part 2

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Theories of Chemical Bonding
Valence Bond Theory and Orbital Overlap

• We think of covalent bonds forming through the sharing of electrons by adjacent atoms.
• In this approach, we assume bonding occurs when orbitals on two atoms overlap.

Overlaps can occur between s, p, and d orbitals on two different atoms.
Orbital Overlap and Bonding

- Increased overlap brings the electrons and nuclei closer together while simultaneously decreasing electron-electron repulsion.
- However, if atoms get too close, the internuclear repulsion greatly raises the energy.

This diagram is known as a Heitler-London diagram.
Orbital Overlap

Some examples of orbital overlaps are shown in the diagram below.
The areas of bond overlap are shaded. This represents increased electron density of the chemical bond.
Valence Bond Theory

• There are two ways orbitals can overlap to form bonds between atoms.
  - Sigma bonds, symbolized by $\sigma$
  - Pi bonds, symbolized by $\pi$

• Hybridization, which we will look at later, is a major player in this approach to bonding.
Sigma (σ) Bonds

- Sigma bonds are characterized by
  - Head-to-head overlap.
  - Cylindrical symmetry of electron density about the internuclear axis.
Pi (\(\pi\)) Bonds

- Pi bonds are characterized by:
  - Side-to-side overlap.
  - Electron density above and below the internuclear axis.
- Although there are two areas of overlap, this is a single 2-electron bond.
Pi ($\pi$) Bonds

A more accurate picture of a pi bond is not a side-to-side overlap, it is better represented by electron density above and below the internuclear axis.
Single Bonds

Single bonds are always $\sigma$ bonds, because $\sigma$ overlap is greater, resulting in a stronger bond and more energy lowering.
Multiple Bonds

In a multiple bond one of the bonds is a $\sigma$ bond and the rest are $\pi$ bonds.
Multiple Bonds

In a multiple bond one of the bonds is a $\sigma$ bond and the rest are $\pi$ bonds.
In a molecule like formaldehyde, an $sp^2$ orbital on carbon overlaps in $\sigma$ fashion with the corresponding orbital on the oxygen. The unhybridized $\rho$ orbitals overlap in $\pi$ fashion.
Multiple Bonds

In triple bonds, as in acetylene, two $sp$ orbitals form a $\sigma$ bond between the carbons, and two pairs of $p$ orbitals overlap in $\pi$ fashion to form the two $\pi$ bonds.
Orbital Overlap and Molecular Shapes

There is a problem with the simple orbital overlap model.

- The shapes of the molecules are linear due to s-s, s-p, or p-p (end to end overlaps)
- The shapes are predicted to be at 90° angles if there are two or three different p orbitals involved on one atom

In order to account for tetrahedral, trigonal bipyramidal, and other geometries arising from the atomic orbitals we recognize, we use a model known as HYBRID ORBITALS.
Orbital Overlap and Molecular Shapes

In the water molecule, the shape due to orbital overlaps are predicted to be at 90° angles.

This model shows the two hydrogen atoms in a water molecule at 90°

In order to account for the actual 104° angle, we assume that the 2 H atoms repel each other.
Orbital Overlap and Molecular Shapes

In the ammonia molecule, the shape due to orbital overlaps is also predicted to be at 90° angles.

This model shows the three hydrogen atoms in an ammonia molecule at 90°.
In order to account for the actual 107° angle, we assume that the 3 H atoms repel each other.

In order to account for tetrahedral, trigonal pyramidal, and other geometries arising from the atomic orbitals we recognize, we use a model known as HYBRID ORBITALS.
Hybrid Orbitals

- Consider beryllium:
  ➢ In its ground electronic state, it would not be able to form bonds because it has no singly-occupied orbitals.
Hybrid Orbitals

But if it absorbs the small amount of energy needed to promote an electron from the 2s to the 2p orbital, it can form two bonds.
Hybrid Orbitals

- Mixing the $s$ and $p$ orbitals yields two degenerate orbitals that are hybrids of the two orbitals.
  - These are called **$sp$ hybrid orbitals** and have two lobes like a $p$ orbital.
  - One of the lobes is larger and more rounded as is the $s$ orbital.
Hybrid Orbitals

• These two degenerate orbitals would align themselves 180° from each other.
• This is consistent with the observed geometry of beryllium compounds: linear.

The hybrid orbitals obey VSEPR Geometry. In this case, linear geometry.
Hybrid Orbitals

- With hybrid orbitals the orbital diagram for beryllium would look like this.
- The $sp$ orbitals are higher in energy than the $1s$ orbital but lower than the $2p$. 
Hybrid Orbitals

Using a similar model for boron leads to...

The mixing of one s and two p orbitals gives rise to three sp² hybrid orbitals
Hybrid Orbitals

...three degenerate $sp^2$ orbitals.

The three $sp^2$ orbitals arrange to form trigonal planar bonds.
Hybrid Orbitals

With carbon we get...

The mixing of one $s$ and three $p$ orbitals gives rise to four $sp^3$ hybrid orbitals
Hybrid Orbitals

...four degenerate $sp^3$ orbitals.

The four $sp^3$ orbitals arrange to form tetrahedral bonds.
**sp³ Hybrid Orbitals**

The bonding in CH₄:

- C has 4 valence e⁻ (each sp³ hybrid contains 1 e⁻).
- Hybrids obey VSEPR (tetrahedral geometry).
- Each H needs 1 valence e⁻ to complete a bond.
- Four equivalent C–H covalent bonds form.

Most compounds with 2nd or 3rd period central atoms are sp³ hybridized.
$sp^3$ Hybrid Orbitals

N and O compounds are often $sp^3$ hybridized:

Four $sp^3$ hybrid orbitals on N

- Lone pair
- Sigma bond pairs

(Only the five valence electrons of N are shown.)

N is $sp^3$ hybridized with one lone pair and three bonding pairs occupying the four $sp^3$ hybrid orbitals...

...which are oriented at angles of 109.5°, close to the experimental bond angle of 107.5°.

Four $sp^3$ hybrid orbitals on O

- Lone pairs
- Sigma bond pairs

(Only the six valence electrons of O are shown.)

O is $sp^3$ hybridized with two lone pairs and two bonding pairs occupying the four $sp^3$ hybrid orbitals...

...which are oriented at angles of 109.5°. The experimental bond angle is 104.5°.
Hybrid Orbitals with $d$ Orbitals

For geometries involving expanded octets on the central atom, we must use $d$ orbitals in our hybrids.

The mixing of one $s$ and three $p$ and one $d$ orbitals gives rise to five $sp^{3}d$ hybrid orbitals.
Hybrid Orbitals with $d$ Orbitals

This leads to five degenerate $sp^3d$ orbitals. This forms a trigonal bipyramidal molecule.

In a similar manner, the mixing of one $s$, three $p$ and two $d$ orbitals gives rise to six degenerate $sp^3d^2$ hybrid orbitals. This forms an octahedral molecule.
Hybrid Orbitals

Once you know the electron-domain geometry, you know the hybridization state of the atom.
Multiple Bonds

- Utilize both hybridized and unhybridized orbitals
- In the formaldehyde molecule (shown at left) an $sp^2$ orbital on carbon overlaps in a $\sigma$ bond with the corresponding orbital on the oxygen.
- Hydrogen atoms overlap with the remaining $sp^2$ orbitals on the carbon.
- The unhybridized $p$ orbitals overlap in $\pi$ bond.
In acetylene, two \( sp \) orbitals form a \( \sigma \) bond between the carbons, and two pairs of \( p \) orbitals overlap in \( \pi \) fashion to form the two \( \pi \) bonds.

\[ H-C≡C-H \]
Delocalized Electrons: Resonance

Previously, we observed, when writing Lewis structures for species like the nitrate ion, we drew resonance structures to more accurately reflect the structure of the molecule or ion.

\[
\text{[O=O]}^- \quad \leftrightarrow \quad \text{[O=N=O]}^- \quad \leftrightarrow \quad \text{[O=O]}^-
\]
Delocalized Electrons: Resonance

- In reality, each of the four atoms in the nitrate ion has a \( p \) orbital.
- The \( p \) orbitals on all three oxygens overlap with the \( p \) orbital on the central nitrogen.
Delocalized Electrons: Resonance

This means the $\pi$ electrons are not localized between the nitrogen and one of the oxygens, but rather are delocalized throughout the ion.
The organic molecule benzene has six $\sigma$ bonds and a $\rho$ orbital on each carbon atom.
Resonance

• In reality the $\pi$ electrons in benzene are not localized, but delocalized.
• The even distribution of the $\pi$ electrons in benzene makes the molecule unusually stable.
Molecular Orbital (MO) Theory

Though valence bond theory effectively conveys most observed properties of ions and molecules, there are some concepts better represented by molecular orbitals. Molecular orbitals are considered to be formed from a linear combination of atomic orbitals (LCAO). If two atomic orbitals combine, then two molecular orbitals will be formed.
Molecular Orbital (MO) Theory

- In MO theory, we invoke the wave nature of electrons.
- If waves interact constructively, the resulting orbital is lower in energy: a bonding molecular orbital.
Molecular Orbital (MO) Theory

If waves interact destructively, the resulting orbital is higher in energy: an antibonding molecular orbital.
Molecular Orbital (MO) Theory

- The resulting molecular orbital (MO) energy diagram will look like this.
- The two atomic 1s orbitals form 2 molecular orbitals.
- The atomic orbitals are shown on the left and right side.
- The molecular orbitals are shown in the middle.
- The lower MO is the $\sigma_{1s}$ bonding orbital.
- The upper MO is the $\sigma^*_{1s}$ antibonding orbital.
MO Theory

- This is the MO energy diagram for an H₂ molecule.
- In H₂ the two electrons go into the bonding molecular orbital.
- The bond order is one half the difference between the number of bonding and antibonding electrons.

Bond Order = \( \frac{1}{2} \left[ (\text{no. of } e^- \text{ in bonding MOs}) - (\text{no. of } e^- \text{ in antibonding MOs}) \right] \)
MO Theory

For hydrogen, with two electrons in the bonding MO and none in the antibonding MO, the bond order is calculated:

\[
\text{Bond Order} = \frac{1}{2} \left[ (\text{no. of } e^- \text{ in bonding MOs}) - (\text{no. of } e^- \text{ in antibonding MOs}) \right]
\]

\[
= \frac{1}{2} \left[ (2) - (0) \right] = 1
\]

This is interpreted to mean that there is 1 bond between the two hydrogen atoms.
In the case of $\text{He}_2$, the bond order would be

$$\frac{1}{2}(2 - 2) = 0$$

The zero is interpreted to mean there are no bonds.

Therefore, $\text{He}_2$ does not exist.
MO Theory

• For atoms with both s and p orbitals, there are two types of interactions:
  ➢ The s and the p orbitals that face each other overlap in σ fashion.
  ➢ The other two sets of p orbitals overlap in π fashion.
MO Theory

• The resulting MO energy diagram looks like this.
• There are both $\sigma$ and $\pi$ bonding molecular orbitals and $\sigma^*$ and $\pi^*$ antibonding molecular orbitals.
• The smaller $p$-block elements in the second period have a sizeable interaction between the $s$ and $p$ orbitals.
• This flips the order of the $s$ and $p$ molecular orbitals in these elements.
### Second-Row MO Diagrams

#### Large 2s-2p interaction

<table>
<thead>
<tr>
<th></th>
<th>B$_2$</th>
<th>C$_2$</th>
<th>N$_2$</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\sigma^*_{2p}$</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\pi^*_{2p}$</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\sigma_{2p}$</td>
<td></td>
<td></td>
<td>$\downarrow$</td>
</tr>
<tr>
<td>$\pi_{2p}$</td>
<td>$\uparrow$ $\uparrow$</td>
<td>$\uparrow$ $\uparrow$</td>
<td>$\uparrow$ $\uparrow$</td>
</tr>
<tr>
<td>$\sigma^*_{2s}$</td>
<td>$\downarrow$</td>
<td>$\downarrow$</td>
<td>$\downarrow$</td>
</tr>
<tr>
<td>$\sigma_{2s}$</td>
<td>$\downarrow$</td>
<td>$\downarrow$</td>
<td>$\downarrow$</td>
</tr>
</tbody>
</table>

#### Small 2s-2p interaction

<table>
<thead>
<tr>
<th></th>
<th>O$_2$</th>
<th>F$_2$</th>
<th>Ne$_2$</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\sigma^*_{2p}$</td>
<td></td>
<td></td>
<td>$\downarrow$</td>
</tr>
<tr>
<td>$\pi^*_{2p}$</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\sigma_{2p}$</td>
<td></td>
<td>$\downarrow$</td>
<td>$\downarrow$</td>
</tr>
<tr>
<td>$\pi_{2p}$</td>
<td>$\downarrow$ $\downarrow$</td>
<td>$\downarrow$ $\downarrow$</td>
<td></td>
</tr>
<tr>
<td>$\sigma^*_{2s}$</td>
<td>$\downarrow$</td>
<td>$\downarrow$</td>
<td>$\downarrow$</td>
</tr>
<tr>
<td>$\sigma_{2s}$</td>
<td>$\downarrow$</td>
<td>$\downarrow$</td>
<td>$\downarrow$</td>
</tr>
</tbody>
</table>

#### Bond Order, Enthalpy, Length, and Magnetic Behavior

<table>
<thead>
<tr>
<th></th>
<th>B$_2$</th>
<th>C$_2$</th>
<th>N$_2$</th>
<th>O$_2$</th>
<th>F$_2$</th>
<th>Ne$_2$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Bond order</td>
<td>1</td>
<td>2</td>
<td>3</td>
<td>2</td>
<td>1</td>
<td>0</td>
</tr>
<tr>
<td>Bond enthalpy (kJ/mol)</td>
<td>290</td>
<td>620</td>
<td>941</td>
<td>495</td>
<td>155</td>
<td>—</td>
</tr>
<tr>
<td>Bond length (Å)</td>
<td>1.59</td>
<td>1.31</td>
<td>1.10</td>
<td>1.21</td>
<td>1.43</td>
<td>—</td>
</tr>
<tr>
<td>Magnetic behavior</td>
<td>Paramagnetic</td>
<td>Diamagnetic</td>
<td>Diamagnetic</td>
<td>Paramagnetic</td>
<td>Diamagnetic</td>
<td>—</td>
</tr>
</tbody>
</table>
## MO Properties for Diatomic Elements

<table>
<thead>
<tr>
<th>Predicted Properties</th>
<th>Observed Properties</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Unpaired e⁻</td>
</tr>
<tr>
<td>Li₂</td>
<td>0</td>
</tr>
<tr>
<td>Be₂</td>
<td>0</td>
</tr>
<tr>
<td>B₂</td>
<td>2</td>
</tr>
<tr>
<td>C₂</td>
<td>0</td>
</tr>
<tr>
<td>N₂</td>
<td>0</td>
</tr>
<tr>
<td>O₂</td>
<td>2</td>
</tr>
<tr>
<td>F₂</td>
<td>0</td>
</tr>
<tr>
<td>Ne₂</td>
<td>0</td>
</tr>
</tbody>
</table>
Metallic Bonds
Physical Properties of Metals

• Conduct heat and electricity.
• Malleable (can be pressed or hammered into sheets).
• Ductile (can be drawn into wire).
• Atoms can slip past each other.
  ➢ So metals aren’t as brittle as other solids.
Electron-Sea Model

- Metals can be thought of as cations suspended in “sea” of valence electrons.
- Attractions hold electrons near cations, but not so tightly as to impede their flow.
Electron-Sea Model

- This explains properties of metals—
  - Conductivity of heat and electricity
  - Deformation

[Diagram of metal ions and electrons]
Molecular Orbital Model

- Electron-sea model does not explain observed trends in melting point, boiling point, heat of fusion, etc.
  - Suggests these properties should increase with increasing number of valence electrons.

<table>
<thead>
<tr>
<th>Metal</th>
<th>Group 3B</th>
<th>Group 6B</th>
<th>Group 8B</th>
</tr>
</thead>
<tbody>
<tr>
<td>Metal</td>
<td>Sc</td>
<td>Cr</td>
<td>Ni</td>
</tr>
<tr>
<td>Melting point (°C)</td>
<td>1541</td>
<td>1857</td>
<td>1455</td>
</tr>
<tr>
<td>Metal</td>
<td>Y</td>
<td>Mo</td>
<td>Pd</td>
</tr>
<tr>
<td>Melting point (°C)</td>
<td>1522</td>
<td>2617</td>
<td>1554</td>
</tr>
<tr>
<td>Metal</td>
<td>La</td>
<td>W</td>
<td>Pt</td>
</tr>
<tr>
<td>Melting point (°C)</td>
<td>918</td>
<td>3410</td>
<td>1772</td>
</tr>
</tbody>
</table>
These trends can be explained by energy bands created by large number of molecular orbitals formed as metal atoms bond with each other.
Molecular Orbital Model

- As with nonmetals, bond order apexes in center of row, then decreases.
- Thus, attractions (and melting point, etc.) apex in center of transition metals. (Group 6B)

<table>
<thead>
<tr>
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<th>Group 3B</th>
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<th>Group 8B</th>
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<tr>
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<td>Metal</td>
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<td>Pt</td>
</tr>
<tr>
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<td>3410</td>
<td>1772</td>
</tr>
</tbody>
</table>
## Alloys

- Mixtures of elements that have properties characteristic of metals.
- Many ordinary uses of metals involve alloys.

<table>
<thead>
<tr>
<th>Primary Element</th>
<th>Name of Alloy</th>
<th>Composition by Mass</th>
<th>Properties</th>
<th>Uses</th>
</tr>
</thead>
<tbody>
<tr>
<td>Bismuth</td>
<td>Wood’s metal</td>
<td>50% Bi, 25% Pb, 12.5% Sn, 12.5% Cd</td>
<td>Low melting point (70°C)</td>
<td>Fuse plugs, automatic sprinklers</td>
</tr>
<tr>
<td>Copper</td>
<td>Yellow brass</td>
<td>67% Cu, 33% Zn</td>
<td>Ductile, takes polish</td>
<td>Hardware items</td>
</tr>
<tr>
<td>Iron</td>
<td>Stainless steel</td>
<td>80.6% Fe, 0.4% C, 18% Cr, 1% Ni</td>
<td>Resists corrosion</td>
<td>Tableware</td>
</tr>
<tr>
<td>Lead</td>
<td>Plumber’s solder</td>
<td>67% Pb, 33% Sn</td>
<td>Low melting point (275°C)</td>
<td>Soldering joints</td>
</tr>
<tr>
<td>Silver</td>
<td>Sterling silver</td>
<td>92.5% Ag, 7.5% Cu</td>
<td>Bright surface</td>
<td>Tableware</td>
</tr>
<tr>
<td></td>
<td>Dental amalgam</td>
<td>70% Ag, 18% Sn, 10% Cu, 2% Hg</td>
<td>Easily worked</td>
<td>Dental fillings</td>
</tr>
</tbody>
</table>
Solution Alloys

Components of alloys are dispersed uniformly

Substitutional alloys:
- Particles close in size.
- Solute particles take place of solvent metal atoms.

Interstitial alloys:
- Solute particles smaller than solvent.
- Solute particles find their way into holes between solvent metal atoms.

In a Substitutional Alloy, atoms of the solute take positions normally occupied by a solvent atom.

In an Interstitial Alloy, solute atoms occupy interstitial positions in the "holes" between the solvent atoms.
Intermetallic Compounds

- Homogeneous alloys with definite properties and compositions.
- $\text{Co}_5\text{Sm}$
  - Used for permanent magnets in headsets and speakers.
Transition Metals

- Many important metals are included in this group.
- Comprised of elements in $d$ block of periodic table.
### Physical Properties of Transition Metals

<table>
<thead>
<tr>
<th>Group: Element:</th>
<th>3B</th>
<th>4B</th>
<th>5B</th>
<th>6B</th>
<th>7B</th>
<th>8B</th>
<th>1B</th>
<th>2B</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>S c</td>
<td>Ti</td>
<td>V</td>
<td>Cr</td>
<td>Mn</td>
<td>Fe</td>
<td>Co</td>
<td>Ni</td>
</tr>
<tr>
<td>Electron config</td>
<td>3d^14s^2</td>
<td>3d^24s^2</td>
<td>3d^34s^2</td>
<td>3d^54s^1</td>
<td>3d^54s^2</td>
<td>3d^64s^2</td>
<td>3d^74s^2</td>
<td>3d^84s^2</td>
</tr>
<tr>
<td>First ionization energy (kJ/mol)</td>
<td>631</td>
<td>658</td>
<td>650</td>
<td>653</td>
<td>717</td>
<td>759</td>
<td>758</td>
<td>737</td>
</tr>
<tr>
<td>Bonding atomic radius (Å)</td>
<td>1.44</td>
<td>1.36</td>
<td>1.25</td>
<td>1.27</td>
<td>1.39</td>
<td>1.25</td>
<td>1.26</td>
<td>1.21</td>
</tr>
<tr>
<td>Density (g/cm³)</td>
<td>3.0</td>
<td>4.5</td>
<td>6.1</td>
<td>7.9</td>
<td>7.2</td>
<td>7.9</td>
<td>8.7</td>
<td>8.9</td>
</tr>
<tr>
<td>Melting point (°C)</td>
<td>1541</td>
<td>1660</td>
<td>1917</td>
<td>1857</td>
<td>1244</td>
<td>1537</td>
<td>1494</td>
<td>1455</td>
</tr>
</tbody>
</table>

- Some of their properties (such as ionization energy, atomic radius, etc.) are suggestive of isolated atoms.
- Others (such as density, melting point, etc.) suggest bulk solid metal.
Atomic Radii

- Trends are similar across all three rows of transition metals.
- While $Z_{\text{eff}}$ increases across row, so does number of nonbonding electrons.
  - These repel each other and increase radius.
Electron Configurations and Oxidation States

- Transition metals often have more than one common oxidation state.
  - Most have +2 state due to loss of $s$ electrons.
  - Oxidation numbers greater than 2 are due to loss of $d$ electrons as well as $s$. 
Electron Configurations and Oxidation States

Many form compounds that have colors.
Colors in Stained Glass
Transition Metals, Nanoparticles and Murano Glass

- The Venetian island of Murano is located 1.5 km from the city of Venice.
- Murano originated its glassmaking in 8th-century Rome with significant Asian and Muslim influences.
- Murano glass is made up of 70% silica sand, with 30% of other substances called “fluxes” and "stabilizers" (soda and lime).
- The colors are obtained by adding small amounts of minerals, oxides, and chemical derivatives to the base composition of the glass powder.

<table>
<thead>
<tr>
<th>Substance</th>
</tr>
</thead>
<tbody>
<tr>
<td>cobalt</td>
</tr>
<tr>
<td>copper</td>
</tr>
<tr>
<td>chromium</td>
</tr>
<tr>
<td>iron</td>
</tr>
<tr>
<td>manganese</td>
</tr>
<tr>
<td>gold</td>
</tr>
</tbody>
</table>

A Ruby-Red Colloidal Gold demonstration kit is available from Flinn Scientific
Electron Configurations and Oxidation States

• Many have significant magnetic properties.
  ➢ In diamagnetic elements, all electron spins are paired.
  ➢ Therefore, there is no net magnetic moment.
Electron Configurations and Oxidation States

- In paramagnetic atoms and ions, there are unpaired spins.
- The magnetic fields are randomly arranged, though, unless placed in an external magnetic field.
Electron Configurations and Oxidation States

In ferromagnetic substances the orientations of magnetic fields from unpaired electrons are affected by spins from electrons around them.
Electron Configurations and Oxidation States

When an external field is applied and then removed, the substance maintains the magnetic moment and becomes a permanent magnet.
Chromium

- Oxidized by HCl or H$_2$SO$_4$ to form blue Cr$^{2+}$ ion.
- Cr$^{2+}$ oxidized by O$_2$ in air to form green Cr$^{3+}$.
- Cr also found in +6 state as in CrO$_4^{2-}$ and the strong oxidizer Cr$_2$O$_7^{2-}$.
Iron

- Exists in solution in $+2$ or $+3$ state.
- Elemental iron reacts with non-oxidizing acids to form $\text{Fe}^{2+}$, which oxidizes in air to $\text{Fe}^{3+}$. 
Iron

- Brown water running from a faucet is caused by insoluble Fe$_2$O$_3$.
- Fe$^{3+}$ soluble in acidic solution, but forms a hydrated oxide as red-brown gel in basic solution.
Copper

- In solution exists in $+1$ or $+2$ state.
- $+1$ salts generally white, insoluble.
- $+2$ salts commonly blue, water-soluble.