Formation of a molecular species

- It is the same as precipitates or gases except a liquid is formed.
- Acid base neutralization reactions will produce water.
- $HNO_3 + NaOH \rightarrow H_2O_{(l)} + NaNO_{3\,(aq)}$
- Net ionic
- $H^+ + OH^- \rightarrow H_2O_{(l)}$

Strong acids

Acid	formula	Acid	Formula
Hydrochloric acid	**HCl**	**Sulfuric Acid**	H_2SO_4
Hydrobromic acid	**HBr**	**Nitric Acid**	HNO_3
Hydriodic acid	**HI**	**Perchloric Acid**	$HClO_4$
		Chloric Acid	$HClO_3$

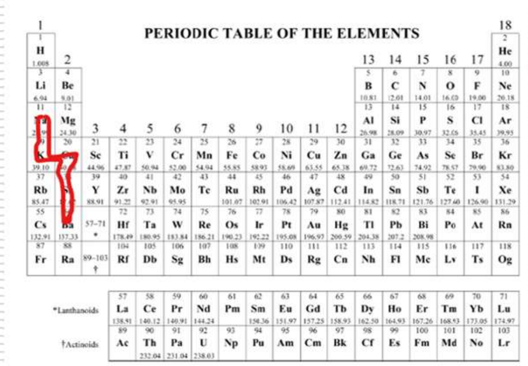
Strong bases

- All of group 1 and group 2 metals (not H) make strong bases.
- However, most of them are not very soluble.
- For example, $Mg(OH)_2$ is a saturated solution at 1.6×10^{-4} M

Common Strong Bases

these make a lightning bolt on the periodic table!

Name	Formula	Name	Formula
Sodium Hydroxide	**NaOH**	**Calcium Hydroxide**	$Ca(OH)_2$
Potassium Hydroxide	**KOH**	**Strontium Hydroxide**	$Sr(OH)_2$
		Barium Hydroxide	$Ba(OH)_2$



Strong acids and bases

- Strong acids and bases are not at equilibrium, there is no reverse reaction.
- Strong acids and bases will never be formed in a net ionic equation.
- All other acids/bases can be formed, and will be formed by reacting the appropriate ion with a strong acid/base.
- *Most other bases are insoluble

Examples

- Calcium hydroxide reacts with chloric acid
- Hydrochloric acid reacts with calcium nitrite
- Nitric acid reacts with sodium chlorite
- Sodium chloride is mixed with sulfuric acid

Chapter 10 Aqueous Solutions and Ionic Equations

- This chapter has already been covered
- *Only dissociate soluble ionic compounds
- Molecular equation
- $Na_2S + CrCl_2 \rightarrow CrS + NaCl$
- Full Ionic Equation
- $2Na^+ + S^{2-} + Cr^{2+} + 2Cl^- \rightarrow CrS + Na^+ + Cl^-$
- Net Ionic Equation
- $Cr^{2+} + S^{2-} \rightarrow CrS$

Chapter 11 Redox Reactions

- Redox or oxidation-reduction reactions are reactions that involve a transfer of electrons.
- Oxidation is the **loss** of electrons.
- Reduction is the **gain** of electrons.
- (think of the charge, OIL RIG)
- $4\ K + O_2 \rightarrow 4\ K^+ + 2\ O^{2-}$
- Potassium get oxidized, oxygen get reduced

Using oxidation states

- In the reaction…
- $2\ Na + 2\ H_2O \rightarrow 2\ NaOH + H_2$
- 0 +1 -2 +1 -2 +1 0
- Note the changes
- Sodium went from 0 to 1
- 2 of the hydrogen atoms went from +1 to 0 (the other two were unchanged)

Breaking into two half reactions

- Sodium must have lost 2 electrons
- $2\ Na \rightarrow 2Na^+ + 2\ e^-$
- And Hydrogen gained two electrons
- $2\ H_2O + 2\ e^- \rightarrow 2\ OH^- + H_2$
- Sodium is oxidized, hydrogen is reduced in this reaction
- Oxidation is an increase in oxidation state
- Reduction is a decrease in oxidation state

Balancing Redox Equations

- by Half Reactions Method or oxidation state method
- The book does not separate these into half reactions, although it adds another step I think it makes it easier

Half reactions

- $Ce^{4+} + Sn^{2+} \rightarrow Ce^{3+} + Sn^{4+}$
- Half reactions
- $Ce^{4+} + e^- \rightarrow Ce^{3+}$
- $Sn^{2+} \rightarrow 2e^- + Sn^{4+}$
- **Electrons lost must equal electrons gained!**
- $\mathbf{2}\ Ce^{4+} + \mathbf{2}\ e^- \rightarrow \mathbf{2}\ Ce^{3+}$
- Merge the two half reactions
- $2\ Ce^{4+} + Sn^{2+} \rightarrow 2\ Ce^{3+} + Sn^{4+}$

Redox reactions in acidic solutions

- It will be noted in the problem
- Balance all elements except hydrogen and oxygen.
- Balance oxygen by adding H_2O (which is always prevalent in an acidic solution)
- Balance hydrogen by adding H^+
- Then balance the charge adding electrons and proceed as normal.

Example

- In an acidic solution
- $Cr_2O_7{}^{2-} + Cl^- \rightarrow Cr^{3+} + Cl_2$
- Half reactions
- $Cr_2O_7{}^{2-} \rightarrow Cr^{3+}$
- $Cl^- \rightarrow Cl_2$

Reduction side

- $Cr_2O_7{}^{2-} \rightarrow Cr^{3+}$
- $Cr_2O_7{}^{2-} \rightarrow 2\ Cr^{3+}$
- $Cr_2O_7{}^{2-} \rightarrow 2\ Cr^{3+} + 7\ H_2O$
- $Cr_2O_7{}^{2-} + 14\ H^+ \rightarrow 2\ Cr^{3+} + 7\ H_2O$
- $Cr_2O_7{}^{2-} + 14\ H^+ + 6\ e^- \rightarrow 2Cr^{3+} + 7\ H_2O$

Oxidation side

- $Cl^- \rightarrow Cl_2$
- $2\ Cl^- \rightarrow Cl_2$
- $2\ Cl^- \rightarrow Cl_2 + 2\ e^-$
- I have to equal 6 e^- so multiply by 3
- $6\ Cl^- \rightarrow 3\ Cl_2 + 6\ e^-$

Combine my half reactions

- $Cr_2O_7{}^{2-} + 14\ H^+ + 6\ e^- \rightarrow 2\ Cr^{3+} + 7\ H_2O$
- $6\ Cl^- \rightarrow 3\ Cl_2 + 6\ e^-$
- And you get
- $Cr_2O_7{}^{2-} + 14\ H^+ + 6Cl^- \rightarrow 2Cr^{3+} + 3\ Cl_2 + 7H_2O$
- The electrons cancel out .

Example

- In an acidic solution
- $MnO_4^- + H_2O_2 \rightarrow Mn^{2+} + O_2$
- Half reactions
- $MnO_4^- \rightarrow Mn^{2+}$
- $H_2O_2 \rightarrow O_2$

Top Equation

- $MnO_4^- \rightarrow Mn^{2+}$
- $MnO_4^- \rightarrow Mn^{2+} + 4\ H_2O$
- $MnO_4^- + 8\ H^+ \rightarrow Mn^{2+} + 4\ H_2O$
- $MnO_4^- + 8\ H^+ + 5\ e^- \rightarrow Mn^{2+} + 4\ H_2O$

Bottom Equation

- $H_2O_2 \rightarrow O_2$
- $H_2O_2 \rightarrow O_2 + 2\ H^+$
- $H_2O_2 \rightarrow O_2 + 2\ H^+ + 2\ e^-$
- I need to equal 5 e^- so…
- That won't work…
- $2MnO_4^- + 16\ H^+ + 10\ e^- \rightarrow 2\ Mn^{2+} + 8\ H_2O$
- $5\ H_2O_2 \rightarrow 5\ O_2 + 10\ H^+ + 10\ e^-$

Add them together

- $2MnO_4^- + 16\ H^+ + 10\ e^- \rightarrow 2\ Mn^{2+} + 8\ H_2O$
- $5\ H_2O_2 \rightarrow 5\ O_2 + 10\ H^+ + 10\ e^-$
- And you get
- $2\ MnO_4^- + 6\ H^+ + 5\ H_2O_2$
  $\rightarrow 2\ Mn^{2+} + 5\ O_2 + 8\ H_2O$
- Notice the H^+ canceled out as well.

Balancing Redox Equations in a basic solution

- Look for the words basic or alkaline
- Follow all rules for an acidic solution.
- After you have completed the acidic reaction add OH^- to each side to neutralize any H^+.
- Combine OH^- and H^+ to make H_2O.
- Cancel out any extra waters from both sides of the equation.

Example

- We will use the same equation as before
- In a basic solution
- $MnO_4^- + H_2O_2 \rightarrow Mn^{2+} + O_2$

- $2\ MnO_4^- + 6\ H^+ + 5\ H_2O_2$
  $\rightarrow 2\ Mn^{2+} + 5\ O_2 + 8\ H_2O$

Basic solution

- Since this is a basic solution we can't have excess H^+.
- We will add OH^- to each side to neutralize all H^+
- $2\ MnO_4^- + 6\ H^+ + 5\ H_2O_2 + 6OH^-$
  $\rightarrow 2\ Mn^{2+} + 5\ O_2 + 8\ H_2O + 6OH^-$
- We added 6 OH^- because there were $6H^+$

Cont.

- $H^+ + OH^- \rightarrow H_2O$
- Combine the hydroxide and hydrogen on the reactant side to make water
- $2\ MnO_4^- + 6\ H_2O + 5\ H_2O_2$
  $\rightarrow 2\ Mn^{2+} + 5\ O_2 + 8\ H_2O + 6OH^-$
- Cancel out waters on both sides

$2\ MnO_4^- + 5\ H_2O_2$
$\rightarrow 2\ Mn^{2+} + 5\ O_2 + 2\ H_2O + 6OH^-$

Another example

- In a basic solution
- $MnO_4^- + SO_3^{2-} \rightarrow MnO_4^{2-} + SO_4^{2-}$
- Half reactions
- $MnO_4^- \rightarrow MnO_4^{2-}$
- $SO_3^{2-} \rightarrow SO_4^{2-}$

Half reactions

- $MnO_4^- \rightarrow MnO_4^{2-}$
- $MnO_4^- + e^- \rightarrow MnO_4^{2-}$

- $SO_3^{2-} \rightarrow SO_4^{2-}$
- $H_2O + SO_3^{2-} \rightarrow SO_4^{2-}$
- $H_2O + SO_3^{2-} \rightarrow SO_4^{2-} + 2\ H^+$
- $H_2O + SO_3^{2-} \rightarrow SO_4^{2-} + 2\ H^+ + 2e^-$
- Double the top reaction

- $2\ MnO_4^- + 2\ e^- \rightarrow 2\ MnO_4^{2-}$
- $H_2O + SO_3^{2-} \rightarrow SO_4^{2-} + 2\ H^+ + 2e^-$
- Combine them
- $2\ MnO_4^- + H_2O + SO_3^{2-}$
  $\rightarrow 2\ MnO_4^{2-} + SO_4^{2-} + 2\ H^+$
- Add OH^-

$2\ MnO_4^- + H_2O + SO_3^{2-} + 2\ OH^-$
$\rightarrow 2\ MnO_4^{2-} + SO_4^{2-} + 2\ H^+ + 2\ OH^-$

- $2\ MnO_4^- + H_2O + SO_3^{2-} + 2\ OH^-$
  $\rightarrow 2\ MnO_4^{2-} + SO_4^{2-} + 2\ H_2O$
- finishing
- $2\ MnO_4^- + SO_3^{2-} + 2\ OH^-$
  $\rightarrow 2\ MnO_4^{2-} + SO_4^{2-} + H_2O$