## Precipitation Reactions

- A precipitate forms when a chemical reaction of solutions produces a solid product.
- A salt is an ionic compound.  Ionic compounds result when metals bind with non-metals (elements from left side of Periodic Table with elements from right side of Periodic Table) or with the common anions ($NO_3^-$, $SO_4^{2-}$, etc.).

## Precipitation Reactions

- Salts may be either soluble or insoluble in water.
- When soluble salts dissolve in solution, they produce cations and anions in the aqueous phase.

## Precipitation Reactions

Solubility Rules

- Salts containing $NH_4^+$ or alkali metal cations (Li, Na, K, Rb, Cs, Fr) are soluble.
- Salts containing the following anions are soluble with the exceptions noted:

$NO_3^-$: no exceptions

$ClO_4^-$: no exceptions

$HSO_4^-$: no exceptions

$CH_3CO_2^-$: no exceptions

## Precipitation Reactions

Solubility Rules (con't.)

- Salts containing the following anions are soluble with the exceptions noted:

$Cl^-$: AgCl, $Hg_2Cl_2$, $PbCl_2$

$Br^-$: AgBr, $Hg_2Br_2$, $PbBr_2$

$I^-$: AgI, $Hg_2I_2$, $PbI_2$

$SO_4^{2-}$: $Hg_2SO_4$, $PbSO_4$, $CaSO_4$, $SrSO_4$, $BaSO_4$

## Precipitation Reactions

Solubility Rules (con't.)

- $Ca(OH)_2$, $Sr(OH)_2$, and $Ba(OH)_2$ are soluble—other hydroxide are insoluble (except group 1)

## Precipitation Reactions

Does a precipitate form when solutions of $NaCl(aq)$ and $AgNO_3(aq)$ are mixed?

1. What does the mixture look like initially?

   $Na^+(aq) + Cl^-(aq) + Ag^+(aq) + NO_3^-(aq)$

## Precipitation Reactions

2. Form possible cation/anion combinations—do any of the resulting species form a precipitate?

Ions in solution: $Na^+$(aq), $Ag^+$(aq), $Cl^-$(aq), $NO_3^-$(aq)

→ NaCl(s)? No—all salts with $Na^+$ are soluble

→ $AgNO_3$(s)? No—all salts with $NO_3^-$ are soluble

→ $NaNO_3$(s)? No—$Na^+$ and $NO_3^-$ solubility rules

→ AgCl(s)? Yes—because this is one of the exceptions to the $Cl^-$ family of soluble salts

## Precipitation Reactions

- The ***net ionic equation*** involves only those ions participating in the chemical reaction:

$$Na^+(aq) + Cl^-(aq) + Ag^+(aq) + NO_3^-(aq) \rightarrow AgCl(s) + Na^+(aq) + NO_3^-(aq)$$

net ionic eqn: $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

- $Na^+$(aq) and $NO_3^-$(aq) are called ***spectator ions*** because they do not participate in the formation of products—their role is to balance the total electrical charge of the solution. Spectator ions are not included in the net ionic equation.

## Aqueous Solutions

- A solution is composed of two parts: the solute and the solvent.
  - The solute is the minor component of the solution.
  - The solvent is the major component of the solution and is the liquid into which the solute is added.
- Aqueous solutions are those in which water acts as the solvent.

## Concentration Units

- *Molarity* (M) is a very commonly used unit of concentration.

  molarity ≡ mol solute / L solution
- Units are mol $L^{-1}$
- Note that the denominator is L total solution (total volume of solution), not L solvent (volume of solvent added).

## Molarity

Determine concentration of a solution in which 6.081 g $NaNO_3$ is dissolved to a total volume of 843 mL.

1. Calculate moles of solute
   a. molar mass of $NaNO_3$

   Na: 1 (22.990 g $mol^{-1}$) = 22.990 g $mol^{-1}$

   N: 1 (14.007 g $mol^{-1}$) = 14.007 g $mol^{-1}$

   O: 3 (15.999 g $mol^{-1}$) = 47.997 g $mol^{-1}$

   molar mass ($NaNO_3$) = 84.994 g $mol^{-1}$

## Molarity

1. Calculate moles of solute (con't.)
   b. moles of $NaNO_3$

   6.081 ~~g~~ / 84.994 ~~g~~ $mol^{-1}$ = 7.155 x $10^{-2}$ mol
2. Calculate molarity of solution
   a. convert volume to L

   843 mL = 0.843 L

   b. calc M

   7.155 x $10^{-2}$ mol / 0.843 L = **8.49 x $10^{-2}$ M**

## Dilutions

- When diluting a solution by adding more solvent, the amount of solute does not change (you do not add or take away the solute), only the concentration changes.

$C_1V_1 = C_2V_2$ (C = concentration
V = volume)

## Dilutions

50.0 mL of .650 M NaCl solution is diluted with 1000.0 mL of water. Determine concentration of final solution.

1. $C_1V_1$ = (0.650 M)(0.0500 L) = 0.0325 mol NaCl
2. final volume = $V_2$ = 50.0 mL + 1000.0 mL
   = 1050.0 mL
3. $C_2V_2 = C_1V_1$ = 0.0325 mol NaCl
   $C_2 = C_1V_1 / V_2$ = 0.0325 mol / 1.0500 L
   = $3.10 \times 10^{-2}$ M

## Acid-Base Reactions

- The simplest definitions of acid and base are:

  An acid donates $H^+$ ion

  A base accepts $H^+$ ion

- An acid reacts with a base to form water:

  $H_3O^+(aq) + OH^-(aq) \rightarrow 2\ H_2O$

  acid base

## Acid-Base Reactions

- The hydronium ion, $H_3O^+$, is actually a hydrated hydrogen ion, $H^+ \cdot H_2O$. ($H^+$ is the chemical formula of a proton, and protons do not hang out in solution alone.)
- Many bases form hydroxide ion, $OH^-$, when dissolved in water.

## Strong and Weak Acids

- A “strong” acid is a compound that dissociates nearly completely in solution to form hydronium ion.

Strong acids

Hydrochloric acid:

$$HCl + H_2O \rightarrow H_3O^+(aq) + Cl^-(aq)$$

Nitric acid:

$$HNO_3 + H_2O \rightarrow H_3O^+(aq) + NO_3^-(aq)$$

## Strong and Weak Acids

Strong acids (con’t.)

Sulfuric acid:

$$H_2SO_4 + H_2O \rightarrow H_3O^+(aq) + HSO_4^-(aq)$$

Perchloric acid:

$$HClO_4 + H_2O \rightarrow H_3O^+(aq) + ClO_4^-(aq)$$

Hydrobromic acid:

$$HBr + H_2O \rightarrow H_3O^+(aq) + Br^-(aq)$$

## Strong and Weak Acids

- A "weak" acid dissociate only partially in water to form hydronium ion.

  Acetic acid:

  $CH_3CO_2H + H_2O \rightarrow H_3O^+(aq) + CH_3CO_2^-(aq)$

  A solution of acetic acid contains all three species: $H_3O^+(aq)$, $CH_3CO_2^-(aq)$, *and* $CH_3CO_2H(aq)$.

## Strong and Weak Bases

- A "strong" base produces stoichiometric quantities of $OH^-(aq)$.

  Group 1 hydroxides: NaOH, KOH, etc.

  Some Group 2 hydroxides: $Ca(OH)_2$, $Sr(OH)_2$, and $Ba(OH)_2$

- A "weak" base accepts a proton from $H_3O^+$, but not water.

  $NH_3 + H_2O \rightarrow NH_3(aq)$

  $NH_3(aq) + H_3O^+(aq) \rightarrow NH_4^+(aq) + H_2O$

## Acid-Base Titrations

- Titrations use a solution of known concentration to determine the concentration of an unknown solution.
- Titrations may be performed with either an acid or a base as the *titrant*—the known solution employed in the titration.
- A acid of known concentration is used to determine a basic solution: a base of known concentration is used to determine an acidic solution.

## Acid-Base Titrations

- The *equivalence point* (the textbook calls this the stoichiometric point) of the titration is reached when:

  moles $H_3O^+(aq)$ = moles $OH^-(aq)$

- If we know (a) the concentration of the titrant, (b) the volume of titrant used, and (c) the volume of the unknown solution, we can determine the concentration of the unknown solution.

## Acid-Base Titrations

Determine the concentration of acetic acid in a sample of vinegar:

$CH_3CO_2H(aq) + OH^-(aq) \rightarrow CH_3CO_2^-(aq) + H_2O$

Experimental Conditions

15.21 g NaOH dissolved in 500.0 mL $H_2O$

25.0 mL of vinegar initially in flask

Volume of NaOH soln. used = 29.97 mL

## Acid-Base Titrations

Determine the concentration of acetic acid in a sample of vinegar:

NaOH Titrant

(15.21 g NaOH) / (39.997 g $mol^{-1}$) = .3803 mol

(.3803 mol) / (.500 L) = .7606 M NaOH soln.

(.7606 M NaOH) (.02997 L) = .02280 mol NaOH at equivalence point

## Acid-Base Titrations

Determine the concentration of acetic acid in a sample of vinegar:

Acetic Acid

At equivalence point:

.02280 mol $OH^-(aq)$ = .02280 mol $H_3O^+(aq)$

25.0 mL of vinegar contains .02280 mol $CH_3CO_2H$:

$[CH_3CO_2H]$ = (.02280 mol) / (.025 L) = .912 M

## Oxidation-Reduction Reactions

- Oxidation-reduction reactions involve the transfer of electrons between reactants and products.
- The *oxidation number* of an atom is the electrical charge on that atom.

$O_2$: O has an oxidation number of 0

$O^{2-}$: O has an oxidation number of -2

$O^+$: O has an oxidation number of +1

## Oxidation-Reduction Reactions

- Oxidation corresponds to a loss of electrons—the *oxidation number* of the atom increases.

  $Na(s) \rightarrow Na^{+}(aq) + e^{-}$

- Reduction corresponds to a gain of electrons—the *oxidation number* of the atom decreases.

  $Cl_2(g) + 2\ e^{-} \rightarrow 2\ Cl^{-}(aq)$

## Oxidation-Reduction Reactions

Rust Formation

$Fe(s) \rightarrow Fe^{2+} + 2\ e^{-}$ $Fe^{0}$ oxidized to $Fe^{2+}$

$Fe^{2+} \rightarrow Fe^{3+} + e^{-}$ $Fe^{2+}$ oxidized to $Fe^{3+}$

$O_2(g) + 2\ H_2O + 4\ e^{-} \rightarrow 4\ OH^{-}$

$O^{0}$ reduced to $O^{2-}$

The electrons produced in the oxidation of Fe are used in the reduction of $O_2$.

$4\ Fe^{2+} + O_2(g) + 4\ H_2O \rightarrow 2\ Fe_2O_3 + 8H^{+}(aq)$

## Oxidation-Reduction Reactions

- *Redox* reactions (oxidation-reduction reactions) can be separated into *half-reactions*—one reaction for the oxidation process and one reaction for the reduction process:

$Mg(s) \rightarrow Mg^{2+} + 2\,e^-$ oxidation

$2\,H_3O^+ + 2\,e^- \rightarrow H_2(g) + 2\,H_2O$ reduction

net: $Mg(s) + 2\,H_3O^+ \rightarrow Mg^{2+} + H_2(g) + 2\,H_2O$

## Balancing Redox Reactions

- *Redox* reactions can be most easily balanced by separating the reaction into its half reactions (one for oxidation and one for reduction), and then balancing the electrons transferred in the process.

## Balancing Redox Reactions

Example

$Cr^{3+}(aq) + H_2(g) + H_2O \rightarrow Cr(s) + H_3O^+(aq)$

Step 1: Which species is oxidized and which species is reduced?

$Cr^{3+}(aq) + H_2(g) + H_2O \rightarrow Cr(s) + H_3O^+(aq)$

Oxid. State: +3 0 +1 -2 0 +1 -2

Cr is reduced (+3 $\rightarrow$ 0)

H from $H_2(g)$ is oxidized (0 $\rightarrow$ +1)

## Balancing Redox Reactions

Example

$Cr^{3+}(aq) + H_2(g) + H_2O \rightarrow Cr(s) + H_3O^+(aq)$

Step 2: Write half reactions for oxidation and reduction processes

oxidation: $H_2(g) + 2\ H_2O \rightarrow 2\ H_3O^+(aq) + 2\ e^-$

reduction: $Cr^{3+}(aq) + 3\ e^- \rightarrow Cr(s)$

## Balancing Redox Reactions

Example

$Cr^{3+}(aq) + H_2(g) + H_2O \rightarrow Cr(s) + H_3O^+(aq)$

Step 3: Balance number of electrons transferred

oxidation: $H_2(g) + 2\ H_2O \rightarrow 2\ H_3O^+(aq) + 2\ e^-$

reduction: $Cr^{3+}(aq) + 3\ e- \rightarrow Cr(s)$

If we multiply the oxidation reaction by 3 and the reduction reaction by 2, then each reaction involves a total of 6 electrons.

## Balancing Redox Reactions

Example

$Cr^{3+}(aq) + H_2(g) + H_2O \rightarrow Cr(s) + H_3O^+(aq)$

Step 3: Balance number of electrons transferred

$3\ H_2(g) + 6\ H_2O \rightarrow 6\ H_3O^+(aq) + 6\ e^-$

$2\ Cr^{3+}(aq) + 6\ e- \rightarrow 2\ Cr(s)$

Summing the two half reactions give the net overall reaction:

$2\ Cr^{3+}(aq) + 3\ H_2(g) + 6\ H_2O \rightarrow 2\ Cr(s) + 6\ H_3O^+(aq)$

## Balancing Redox Reactions

Example

$Cr^{3+}(aq) + H_2(g) + H_2O \rightarrow Cr(s) + H_3O^+(aq)$

Step 4: Check that atoms and charges are balanced

$2\ Cr^{3+}(aq) + 3\ H_2(g) + 6\ H_2O \rightarrow 2\ Cr(s) + 6\ H_3O^+(aq)$

2 Cr on each side

18 H on each side

6 O on each side

+6 charge on reactants; +6 charge on products

## Balancing Redox Reactions

Example

$Hg(\ell) + H_2S(g) + SnF_6^{2-}(aq) + H_2O \rightarrow$

$HgS(s) + Sn(s) + F^-(aq) + H_3O^+(aq)$

Step 1: Which species is oxidized and which species is reduced?

$Hg(\ell) + H_2S(g) + SnF_6^{2-}(aq) + H_2O \rightarrow$

0 +1 -2 +4 -1 +1 -2

$HgS(s) + Sn(s) + 6F^-(aq) + H_3O^+(aq)$

+2 -2 0 -1 +1 -2

Oxidation: Hg is oxidized from 0 to +2 oxidation state

Reduction: Sn is reduced from +4 to 0 oxidation state

# Balancing Redox Reactions

Example

$Hg(\ell) + H_2S(g) + SnF_6^{2-}(aq) + H_2O \rightarrow$

$HgS(s) + Sn(s) + F^-(aq) + H_3O^+(aq)$

Step 2: Write half reactions for oxidation and reduction processes

Oxidation: $Hg(\ell) + H_2S(g) + H_2O \rightarrow HgS(s) + H_3O^+(aq)$

mass balance: Hg, S, and O are balanced as written, but not H

$Hg(\ell) + H_2S(g) + 2H_2O \rightarrow HgS(s) + 2H_3O^+(aq)$

add electrons to achieve charge balance:

$Hg(\ell) + H_2S(g) + 2H_2O \rightarrow HgS(s) + 2H_3O^+(aq) + 2e^-$

# Balancing Redox Reactions

Example

$Hg(\ell) + H_2S(g) + SnF_6^{2-}(aq) + H_2O \rightarrow$

$HgS(s) + Sn(s) + F^-(aq) + H_3O^+(aq)$

Step 2: Write half reactions for oxidation and reduction processes

Reduction: $SnF_6^{2-}(aq) \rightarrow Sn(s) + F^-(aq)$

mass balance: Sn is balanced as written, but not F

$SnF_6^{2-}(aq) \rightarrow Sn(s) + 6F^-(aq)$

add electrons to achieve charge balance:

$SnF_6^{2-}(aq) + 4e^- \rightarrow Sn(s) + 6F^-(aq)$

# Balancing Redox Reactions

Example

$Hg(\ell) + H_2S(g) + SnF_6^{2-}(aq) + H_2O \rightarrow$

$HgS(s) + Sn(s) + F^-(aq) + H_3O^+(aq)$

Step 3: Balance number of electrons transferred

$(Hg(\ell) + H_2S(g) + 2H_2O \rightarrow HgS(s) + 2H_3O^+(aq) + 2e^-) \times 2$

$SnF_6^{2-}(aq) + 4e^- \rightarrow Sn(s) + 6F^-(aq)$

Result:

$2Hg(\ell) + 2H_2S(g) + 4H_2O \rightarrow 2HgS(s) + 4H_3O^+(aq) + 4e^-$

$SnF_6^{2-}(aq) + 4e^- \rightarrow Sn(s) + 6F^-(aq)$

# Balancing Redox Reactions

Example

$Hg(\ell) + H_2S(g) + SnF_6^{2-}(aq) + H_2O \rightarrow$

$HgS(s) + Sn(s) + F^-(aq) + H_3O^+(aq)$

Step 3: Balance number of electrons transferred

$2Hg(\ell) + 2H_2S(g) + 4H_2O \rightarrow 2HgS(s) + 4H_3O^+(aq) + 4e^-$

$SnF_6^{2-}(aq) + 4e^- \rightarrow Sn(s) + 6F^-(aq)$

Sum of chemical equations:

$2Hg(\ell) + 2H_2S(g) + SnF_6^{2-}(aq) + 4H_2O \rightarrow$

$2HgS(s) + Sn(s) + 6F^-(aq) + 4H_3O^+(aq)$

# Balancing Redox Reactions

Example

$2Hg(\ell) + 2H_2S(g) + SnF_6^{2-}(aq) + 4H_2O \rightarrow$

$2HgS(s) + Sn(s) + 6F^-(aq) + 4H_3O^+(aq)$

Step 4: Check mass and charge balance

| | Reactants | Products |
|---|---|---|
| Hg | 2 | 2 |
| S | 2 | 2 |
| Sn | 1 | 1 |
| F | 6 | 6 |
| H | 12 | 12 |
| O | 4 | 4 |
| charge | -2 | -6 + (+4) = -2 |

# Oxidation-Reduction Reactions