

## Chapter 4 Chemical Reactions (Sections [4.2](#), [3.3](#), and [4.5](#))

### Types of Chemical Reactions

- A precipitation reaction forms an insoluble ionic solid.
- An acid-base reaction involves the transfer of an  $\text{H}^{+1}$  ion (which is just one proton only).
- A red-ox reaction (reduction and oxidation) involves the transfer of electrons ( $\text{e}^{-1}$ 's).

### Ionic Theory of Solutions

- Ionic substances produce freely moving ions when dissolved in water, and the ions carry electric current. (S. Arrhenius, 1884)
- An [electrolyte](#) is a substance that conducts electricity when dissolved in water.
- A strong electrolyte exists entirely as dissociated ions when dissolved in a solution. This includes all soluble ionic compounds:  $\text{NaCl}_{(s)} \rightarrow \text{Na}^{+1}_{(aq)} + \text{Cl}^{-1}_{(aq)}$
- HCl is a strong acid and is molecular. But, it is a strong electrolyte because it dissociates entirely when it is dissolved in water.  $\text{HCl}_{(g)} + \text{H}_2\text{O}_{(L)} \rightarrow \text{H}_3\text{O}^{+1}_{(aq)} + \text{Cl}^{-1}_{(aq)}$
- With a weak electrolyte, only a small % of the dissolved substance exists as a pair of ions. Weak electrolytes are generally molecular substances for which most, but not all, of the dissolved molecules are not ionized. This includes all weak acids and bases.
- Only 1 % of the weak base  $\text{NH}_3$  is converted into an ion pair (forward reaction). The remainder (99 %) exists as dissolved molecules.  
 $\text{NH}_{3(aq)} + \text{H}_2\text{O}_{(L)} \rightarrow \text{NH}_4^{+1}_{(aq)} + \text{OH}^{-1}_{(aq)}$
- A nonelectrolyte is a dissolved molecule that does not normally form ions. It does not conduct electricity when dissolved.
- Methanol is a soluble molecule that does not easily ionize in water.  
 $\text{CH}_3\text{OH}_{(L)} \rightarrow \text{CH}_3\text{OH}_{(aq)}$

[Solubility Rules](#) – See also [Table 4.1](#) in Open Stax and Appendix G in Lab Manual

### Ex 4.01 Determine if Ionic Compounds are Soluble

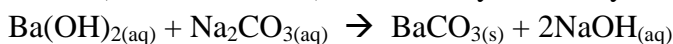
- $\text{PbCl}_2$  Most chlorides are soluble, but see the exceptions to rule 3.
- $\text{LiBr}$  See rules 1 (alkali metals in column 1A) and 3 (halogens in column 7A).  
Are any  $\text{Li}^{+1}$  salts insoluble? Is  $\text{Li}^{+1}$  an exception to the rule for  $\text{Br}^{-1}$ ?
- $\text{Hg}(\text{NO}_3)_2$  See rule 2. Are any  $\text{NO}_3^{-1}$  salts insoluble?

### Precipitation Reactions

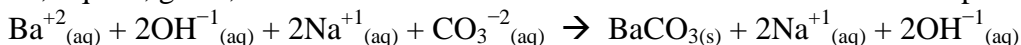
- A precipitate is an insoluble ionic compound formed from reaction of soluble compounds.
- A metathesis reaction is a “Double Exchange” of parts between two reactants:  
 $\text{AB} + \text{CD} \rightarrow \text{AD} + \text{CB}$  (Either AD or CB will be a solid precipitate.)
- Precipitation occurs when the anions and cations exchange to form a solid.

## Chemical Equations

- The molecular equation always has reactants and products written as molecules (or formula units), even if they exist only as dissolved ions.

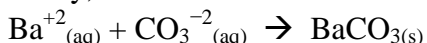


- The complete ionic equation has the dissolved ionic compounds written as ions. Solids, liquids, gases, and dissolved molecules are written as non-ionized compounds.



- The net ionic equation does not include spectator ions that do not change form.

Generally, the net reaction is the formation of a solid or gas.



### Ex 4.02 Use [Solubility Rules](#) to Determine if a Precipitation Reaction Occurs

- a. Potassium chloride and cobalt(II) nitrate are mixed, and the partners are switched.

See rules 1, 2, and 3. Are there any insoluble compounds involved? (No.)

Are there any ions that do not cancel from the net equation? (No.)

- b. Iron(III) sulfate and sodium hydroxide are mixed, and the partners are switched.

See rules 1, 4, and 8. Which sulfates are insoluble? Which hydroxides are insoluble?

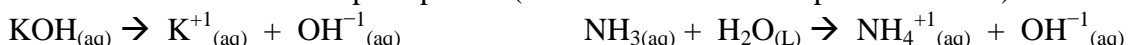
$\text{Na}^{+1}$  and  $\text{SO}_4^{-2}$  are spectators, so the net ionic reaction is  $2\text{Fe}^{+3}_{(\text{aq})} + 6\text{OH}^{-1}_{(\text{aq})} \rightarrow 2\text{Fe}(\text{OH})_{3(\text{s})}$

## Acid-Base Reactions

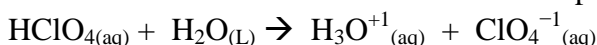
- An acid and a base neutralize each other to form less reactive substances (water and a salt).

- An acid is a substance that donates a proton (produces  $\text{H}^{+1}$ )  $\text{HClO}_4 \rightarrow \text{H}^{+1} + \text{ClO}_4^{-1}$

- A base is a substance that accepts a proton (reacts with  $\text{H}^{+1}$  and/or produces  $\text{OH}^{-1}$ )



- The net acid-base reaction is the transfer of one proton.



- A [strong acid or base](#) dissociates completely in water, existing entirely as ions.

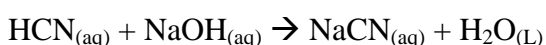
They are always strong electrolytes. Examples include HCl and NaOH.

- A weak acid or base only partially ionizes, and exists primarily as a dissolved molecule.

They are always weak electrolytes. Examples include  $\text{CH}_3\text{COOH}$  and  $\text{NH}_3$ .

- A neutralization salt is an ionic compound that results from an acid-base reaction.

NaCN is the neutralization salt in this reaction:



### Ex 4.03 Net Ionic Equations for Strong and Weak Acids reacting with a Strong Base

- a.  $2\text{HNO}_{3(\text{aq})} + \text{Ba}(\text{OH})_{2(\text{aq})} \rightarrow \text{Ba}(\text{NO}_3)_{2(\text{aq})} + 2\text{H}_2\text{O}_{(\text{L})}$

$\text{Ba}^{+2}_{(\text{aq})}$  and  $\text{NO}_3^{-1}_{(\text{aq})}$  are spectator ions, which do not change and can be cancelled out.

- b.  $\text{CH}_3\text{COOH}_{(\text{aq})} + \text{KOH}_{(\text{aq})} \rightarrow \text{CH}_3\text{COOK}_{(\text{aq})} + \text{H}_2\text{O}_{(\text{L})}$

$\text{K}^{+1}_{(\text{aq})}$  is a spectator that cancels out.

$\text{CH}_3\text{COO}^{-1}_{(\text{aq})}$  is not, because  $\text{CH}_3\text{COOH}_{(\text{aq})}$  is a weak electrolyte that is not ionized.

#### Ex 4.04 [Strong Acids and Bases](#)

- Assume acids and bases not on the table are weak.
- KOH and HCl are on the table (strong), CH<sub>3</sub>COOH and HNO<sub>2</sub> are not on the table (weak).

#### Ex 4.05 Net Ionic Equation for a Neutralization Reaction (Strong Base with a Weak Acid)

- Molecular Equation:  $\text{KOH}_{(\text{aq})} + \text{HNO}_{2(\text{aq})} \rightarrow \text{H}_2\text{O}_{(\text{L})} + \text{KNO}_{2(\text{aq})}$
- The Net Ionic Equation still has the weak acid:  $\text{OH}^{-1}_{(\text{aq})} + \text{HNO}_{2(\text{aq})} \rightarrow \text{H}_2\text{O}_{(\text{L})} + \text{NO}_2^{-1}_{(\text{aq})}$

#### [Gas Formation from Acid-Base Reactions](#)

- Carbon dioxide (CO<sub>2</sub>) forms when HCO<sub>3</sub><sup>-1</sup> or CO<sub>3</sub><sup>-2</sup> react with an acid:  
 $\text{HCO}_3^{-1}_{(\text{aq})} + \text{H}^{+1}_{(\text{aq})} \rightarrow \text{H}_2\text{O}_{(\text{L})} + \text{CO}_{2(\text{g})}$
- Sulfur dioxide (SO<sub>2</sub>) forms when sulfite (SO<sub>3</sub><sup>-2</sup>) reacts with an acid:  
 $\text{SO}_3^{-2}_{(\text{aq})} + 2\text{H}^{+1}_{(\text{aq})} \rightarrow \text{H}_2\text{O}_{(\text{L})} + \text{SO}_{2(\text{g})}$
- Hydrogen sulfide (H<sub>2</sub>S) forms when HS<sup>-1</sup> or S<sup>-2</sup> react with an acid:  
 $\text{HS}^{-1}_{(\text{aq})} + \text{H}^{+1}_{(\text{aq})} \rightarrow \text{H}_2\text{S}_{(\text{g})}$

#### Ex 4.06 Net Ionic Equation for an Acid-Base Reaction with Gas Formation

- Molecular equation:  $\text{NiS}_{(\text{s})} + 2\text{HCl}_{(\text{aq})} \rightarrow \text{H}_2\text{S}_{(\text{g})} + \text{NiCl}_{2(\text{aq})}$
- Net Ionic Equation (still has solid and gas):  $\text{NiS}_{(\text{s})} + 2\text{H}^{+1}_{(\text{aq})} \rightarrow \text{H}_2\text{S}_{(\text{g})} + \text{Ni}^{+2}_{(\text{aq})}$

#### Reduction – Oxidation (RedOx Reactions) – Involves transfer of electrons (e<sup>-1</sup>)

- Molecular Equation:  $\text{Fe}_{(\text{s})} + \text{CuSO}_{4(\text{aq})} \rightarrow \text{FeSO}_{4(\text{aq})} + \text{Cu}_{(\text{s})}$
- Net Ionic Equation:  $\text{Fe}_{(\text{s})} + \text{Cu}^{+2}_{(\text{aq})} \rightarrow \text{Fe}^{2+}_{(\text{aq})} + \text{Cu}_{(\text{s})}$
- Fe<sup>+2</sup> is Iron (II), its oxidation number is +2 (equivalent to charge, per rules for ox numbers)
- Ox numbers change in a redox reaction: Fe goes from 0 to +2, while Cu goes from +2 to 0.

#### Rules for Oxidation Numbers

- An uncharged atom of any element has ox number of zero (0).
- A monatomic ion has an ox number that is equal to its charge.
- Oxygen in a compound always has an ox number of -2, as in H<sub>2</sub>O, except for peroxides, such as H<sub>2</sub>O<sub>2</sub>, where it is -1.
- Hydrogen in a compound always has an ox number of +1, as in H<sub>2</sub>O.  
Hydrides, such as NaH, are the exceptions, where the ox number is -1.
- Halogens (F, Cl, Br, & I) in a binary compound always have ox # of -1, as in HCl and NaCl.  
However for oxoanions, there are several possible ox numbers for halogens.  
For example, in hypochlorous acid, HClO, the Cl has an ox number of +1.
- For neutral compounds, the sum of the ox numbers is always zero.  
With Fe<sub>2</sub>O<sub>3</sub>, we have 2(+3) + 3(-2) = 0
- For polyatomic ions, the sum of ox numbers is the same as the charge.  
With the sulfate ion, SO<sub>4</sub><sup>-2</sup>, we have S + 4(-2) = -2, so S = +6

#### Ex 4.07 Assign Oxidation Numbers to Cl atoms by using Algebra Equations

- a.  $\text{KClO}_4$      $(+1) + \text{Cl} + 4(-2) = 0$              $\text{Cl} = +7$   
b.  $\text{ClO}_2^{-1}$      $\text{Cl} + 2(-2) = -1$              $\text{Cl} = +3$

RedOx Reactions can be split into two parts (half-cells)

- Oxidation is loss of  $e^{-1}$  (leo or oil)             $\text{Fe}_{(s)} \rightarrow \text{Fe}^{+2}_{(aq)} + 2e^{-1}$
- Reduction is gain of  $e^{-1}$  (ger or rig)             $\text{Cu}^{+2}_{(aq)} + 2e^{-1} \rightarrow \text{Cu}_{(s)}$
- A reducing agent (like  $\text{Fe}_{(s)}$ ) reduces another substance, and is itself oxidized.
- An oxidizing agent (like  $\text{Cu}^{+2}_{(aq)}$ ) oxidizes another substance, and is itself reduced.

Types of RedOx Reactions – Combination, Decomposition, Displacement, and Combustion

- Combination occurs when two substances combine to form a third substance.  
 $2\text{K}_{(s)} + \text{Cl}_{2(g)} \rightarrow 2\text{KCl}_{(s)}$
- Decomposition occurs when a single substance breaks down into two or more substances.  
 $2\text{NaH}_{(s)} \rightarrow 2\text{Na}_{(s)} + \text{H}_{2(g)}$
- Displacement occurs when an element reacts with a compound to displace another element.  
 $\text{Mg}_{(s)} + 2\text{AgNO}_{3(aq)} \rightarrow 2\text{Ag}_{(s)} + \text{Mg}(\text{NO}_3)_{2(aq)}$
- Combustion is a reaction with  $\text{O}_2$ , and usually creates heat or a flame.  
 $2\text{C}_3\text{H}_{8(g)} + 10\text{O}_{2(g)} \rightarrow 6\text{CO}_{2(g)} + 8\text{H}_2\text{O}_{(g)}$

Reactions that appear like RedOx, but are NOT (because ox numbers do not change)

- $\text{MgO}_{(s)} + \text{SO}_{2(g)} \rightarrow \text{MgSO}_{3(s)}$
- $\text{CaCO}_{3(s)} \rightarrow \text{CaO}_{(s)} + \text{CO}_{2(g)}$

Activity Series

- The [Activity Series of Metals](#) shows the order in which they will displace each other.
- An atom of an element will react with a compound to displace an atom below it on the chart.  
For instance,  $\text{Ca}_{(s)}$  will displace  $\text{Mg}^{+2}_{(aq)}$  in  $\text{MgCl}_{2(aq)}$  to form  $\text{CaCl}_{2(aq)}$  and  $\text{Mg}_{(s)}$ .
- The most reactive (top of chart) will react with water to create  $\text{H}_{2(g)}$ .  
 $2\text{K}_{(s)} + 2\text{H}_2\text{O}_{(L)} \rightarrow 2\text{KOH}_{(aq)} + \text{H}_{2(g)}$
- The middle section (above  $\text{H}_2$ ) will react with acids ( $\text{H}^{+1}$ ) to create  $\text{H}_{2(g)}$ .
- The bottom section is not reactive towards water or acids.

Balancing RedOx Reactions

- First, determine the species being oxidized and reduced, and separate them into half-cells.
- Balance each element so that it has the same number of atoms on each side of the reactions.
- Then, balance charges with  $e^{-1}$ 's so that each side has same total charge.  
Put  $e^{-1}$ 's on the right side of the oxidation half-cell, and on the left side of reduction half-cell.
- Each of the two half cells needs to have the same number of  $e^{-1}$ 's.  
So, multiply the balanced half-cells by factors where necessary.

#### Ex 4.08 Balance the RedOx Reaction which forms Magnesium Nitride

- Unbalanced reaction:  $\text{Mg}_{(s)} + \text{N}_{2(g)} \rightarrow \text{Mg}_3\text{N}_{2(s)}$
- Note that  $\text{Mg}_3\text{N}_2$  is composed of  $\text{Mg}^{+2}$  and  $\text{N}^{-3}$
- So, reaction can be split into half-cells (unbalanced):  
 $\text{Mg}_{(s)} \rightarrow \text{Mg}^{+2}$  and  $\text{N}_{2(g)} \rightarrow 2\text{N}^{-3}$  (Note same number of each atom on each side!)
- Add  $e^{-1}$ 's to balance charges:  
 $\text{Mg}_{(s)} \rightarrow \text{Mg}^{+2} + 2e^{-1}$  and  $\text{N}_{2(g)} + 6e^{-1} \rightarrow 2\text{N}^{-3}$  (Note same total charge on each side!)
- Use factors to give each half-cell same number of  $e^{-1}$ 's:  
 $3(\text{Mg}_{(s)} \rightarrow \text{Mg}^{+2} + 2e^{-1})$  or  $3\text{Mg}_{(s)} \rightarrow 3\text{Mg}^{+2} + 6e^{-1}$
- The sum is:  $3\text{Mg}_{(s)} + \text{N}_{2(g)} \rightarrow 3\text{Mg}^{+2} + 2\text{N}^{-3}$
- Or:  $3\text{Mg}_{(s)} + \text{N}_{2(g)} \rightarrow \text{Mg}_3\text{N}_{2(s)}$

#### Working with Solutions

- Solutions are necessary in the lab for species which react only in dissolved form.
- The solute is the dissolved species.
- The solvent is the liquid substance which contains the dissolved substance.
- Concentration is the quantity of solute per quantity of solution or solvent.
- Molarity (M) is the moles of solute per liter of solution.

#### Ex 4.09 Determine Molarity for 0.811 g $\text{KNO}_3$ in 75.0 ml of Aqueous Solution

- Moles = Mass  $\div$  Molar Mass =  $(0.811 \text{ g } \text{KNO}_3) \left( \frac{1 \text{ mole}}{101.1 \text{ g}} \right) = 0.00802 \text{ mol } \text{KNO}_3$
- $(75.0 \text{ ml}) \left( \frac{1 \text{ L}}{1000 \text{ ml}} \right) = 0.0750 \text{ L}$
- Molarity =  $\frac{0.00802 \text{ mole}}{0.0750 \text{ L}} = 0.107 \text{ mol/L} = 0.107 \text{ M}$

#### Ex 4.10 Use Molarity (0.250 M) as a Conversion Factor between Moles and Volume

- Moles = Mass  $\div$  Molar Mass =  $(0.254 \text{ g } \text{KOH}) \left( \frac{1 \text{ mole}}{56.1 \text{ g}} \right) = 0.00453 \text{ mol } \text{KOH}$
- Volume = Moles  $\div$  Molarity =  $(0.00453 \text{ mol}) \left( \frac{1 \text{ L}}{0.250 \text{ mole}} \right) = 0.0181 \text{ L}$
- $(0.0181 \text{ L}) \left( \frac{1000 \text{ ml}}{1 \text{ L}} \right) = 18.1 \text{ ml}$

#### Diluting Solutions

- If only the pure solvent is added, then the total moles of solute do not change.
- Moles = Molarity  $\times$  Volume =  $M \times V = \left( \frac{\text{moles}}{\text{L}} \right) \times (\text{L})$
- Moles =  $M_i V_i = M_f V_f$  where M is molarity (mol/L) and V is total volume (L).
- Rearrange to get  $V_f = V_i \left( \frac{M_i}{M_f} \right)$  and  $M_f = M_i \left( \frac{V_i}{V_f} \right)$
- $V_f$  is the final total volume, so  $V_f - V_i =$  volume of solvent added.

Ex 4.11 Dilute an 8.50 M solution to obtain 125.0 ml of a 1.00 M solution

- Moles of Solute =  $M_f V_f = (1.00 \text{ mol/L})(0.1250 \text{ L}) = 0.125 \text{ mol}$
- $V_i = \frac{(M_f V_f)}{M_i} = \frac{\text{moles}}{M_i} = \frac{0.125 \text{ mole}}{8.50 \frac{\text{mole}}{\text{L}}} = 0.0147 \text{ L} \quad V_i = (0.0147 \text{ L})\left(\frac{1000 \text{ ml}}{1 \text{ L}}\right) = 14.7 \text{ ml}$
- Volume of H<sub>2</sub>O Added =  $V_f - V_i = 125.0 \text{ ml} - 14.7 \text{ ml} = 110.3 \text{ ml}$

#### Quantitative Analysis

- Determination of the amount of a substance present.
- Result of analysis is generally a mass, volume, or concentration.

#### Gravimetric Analysis (a type of quantitative analysis)

- Weigh the product to determine the mass of a substance that is present.
- For example, a precipitate can be filtered from a solution, and then dried and weighed.

Ex 4.12 Find mass of Pb<sup>+2</sup> in 1.000 L of polluted water if 294.6 mg PbSO<sub>4</sub> are precipitated.

- $(294.6 \text{ mg PbSO}_4)(1 \text{ g} / 1000 \text{ mg}) = 0.2946 \text{ g PbSO}_4$
- $(0.2946 \text{ g PbSO}_4)\left(\frac{1 \text{ mol}}{303.3 \text{ g}}\right) = 9.713 \times 10^{-4} \text{ mol PbSO}_4$
- $(9.713 \times 10^{-4} \text{ mol PbSO}_4)(1 \text{ mol Pb}^{+2} / 1 \text{ mol PbSO}_4) = 9.713 \times 10^{-4} \text{ mol Pb}^{+2}$
- $(9.713 \times 10^{-4} \text{ mol Pb}^{+2})\left(\frac{207.2 \text{ g}}{1 \text{ mole}}\right) = 0.2013 \text{ g Pb}^{+2}$
- $(0.2013 \text{ g Pb}^{+2})\left(\frac{1000 \text{ mg}}{1 \text{ g}}\right) = 201.3 \text{ mg of Pb}^{+2} \text{ (per liter)}$

#### Volumetric Analysis (another type of quantitative analysis)

- Measure the solution volume which is equivalent to the mass or moles of solute present.
- Generally uses a titration, where a carefully measured volume of solution is added to reach a stoichiometric endpoint.

Ex 4.13 Calculate the volume of 0.225 M KOH needed to neutralize (or to reach endpoint).

- For an acid-base titration, the stoichiometric endpoint is where the moles of H<sup>+1</sup> (in 45.5 ml of 0.125 M H<sub>2</sub>SO<sub>4</sub>) equal the moles of OH<sup>-1</sup> (from 0.225 M KOH).
- $(45.5 \text{ ml H}_2\text{SO}_4)(1 \text{ L} / 1000 \text{ ml})(0.125 \text{ mol/L}) = 0.00569 \text{ mol H}_2\text{SO}_4$
- $(0.00569 \text{ mol H}_2\text{SO}_4)(2 \text{ H}^{+1} / 1 \text{ H}_2\text{SO}_4) = 0.01138 \text{ mol H}^{+1} = 0.01138 \text{ mol OH}^{-1}$
- $(0.01138 \text{ mol OH}^{-1})(1 \text{ KOH} / 1 \text{ OH}^{-1}) = 0.01138 \text{ mol KOH}$
- $(0.01138 \text{ mol KOH}) / (0.225 \text{ mol/L}) = 0.0506 \text{ L}$
- $(0.0506 \text{ L})(1000 \text{ ml} / 1 \text{ L}) = 50.6 \text{ ml of } 0.225 \text{ M KOH}$

Ex 4.14 Calculate the mass of HCl titrated by 3.85 ml of 0.225 mol/L KOH

- $(3.85 \text{ ml})(1 \text{ L} / 1000 \text{ ml})(0.225 \text{ mol/L KOH}) = 0.000866 \text{ mol KOH}$
- $(0.000866 \text{ mol KOH})(1 \text{ mol HCl} / 1 \text{ mol KOH}) = 0.000866 \text{ mol HCl}$
- $(0.000866 \text{ mol HCl})(36.5 \text{ g/mol}) = 0.0316 \text{ g HCl}$