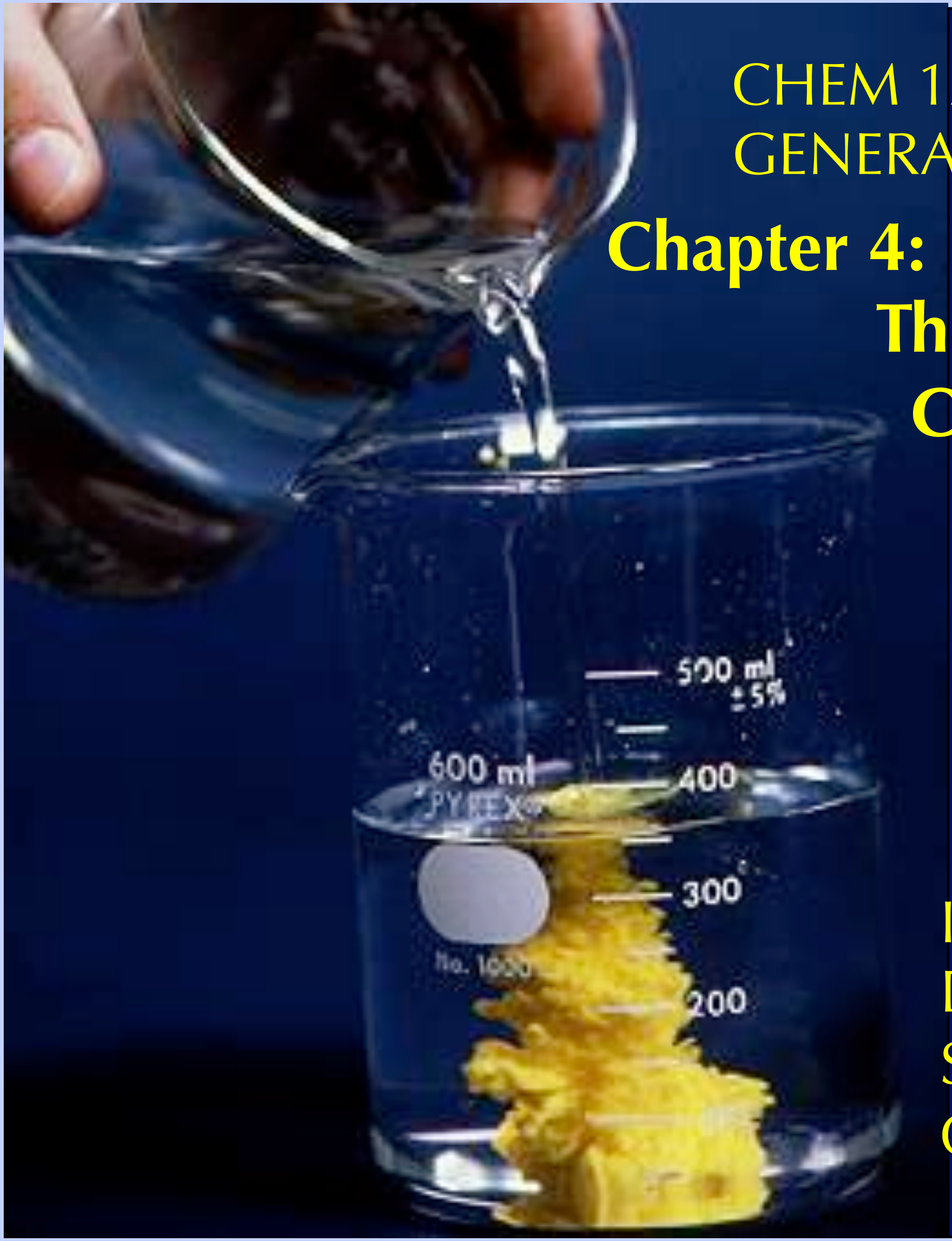
CHEM 1A:
GENERAL CHEMISTRY

Chapter 4:
Three Major Types of Chemical Reactions



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Chapter 4

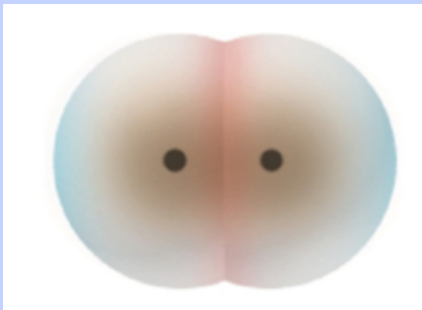
Three Major Classes of Chemical Reactions

The Major Classes of Chemical Reactions

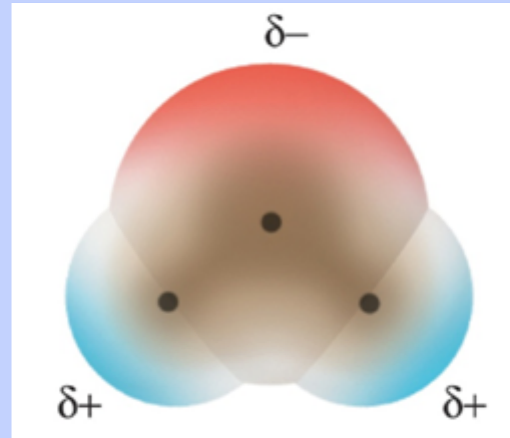
Water as a Solvent

- Water is a polar molecule
  - since it has uneven electron distribution
  - and a bent molecular shape.
- Water readily dissolves a variety of substances.
- Water interacts strongly with its solutes and often plays an active role in aqueous reactions.

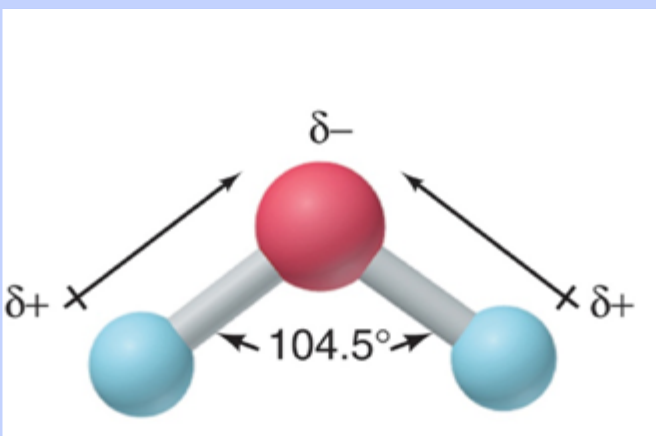
Figure 4.1 Electron distribution in molecules of H_2 and H_2O.



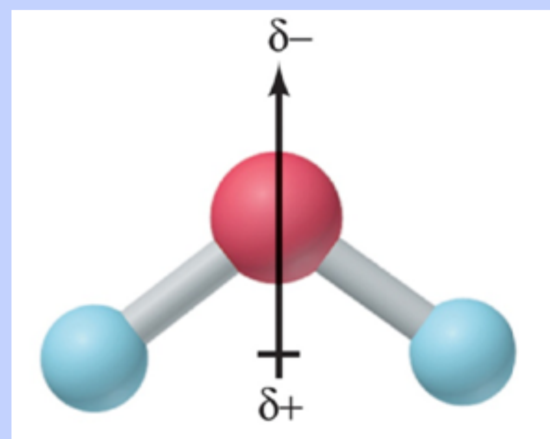
A. Electron charge distribution in H_2 is symmetrical.



B. Electron charge distribution in H_2O is asymmetrical.



C. Each bond in H_2O is polar.



D. The whole H_2O molecule is polar.

Figure 4.2 An ionic compound dissolving in water.

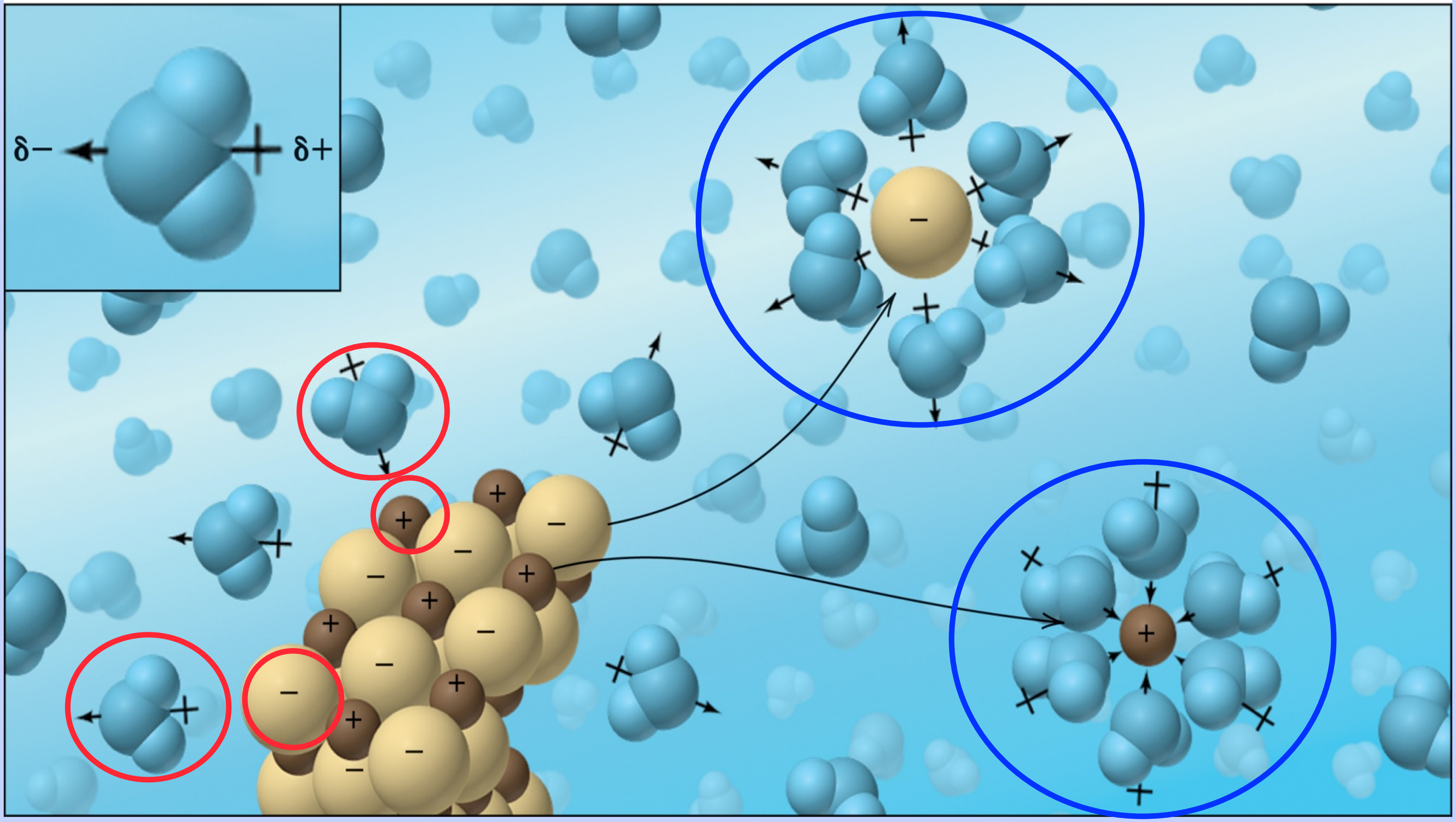
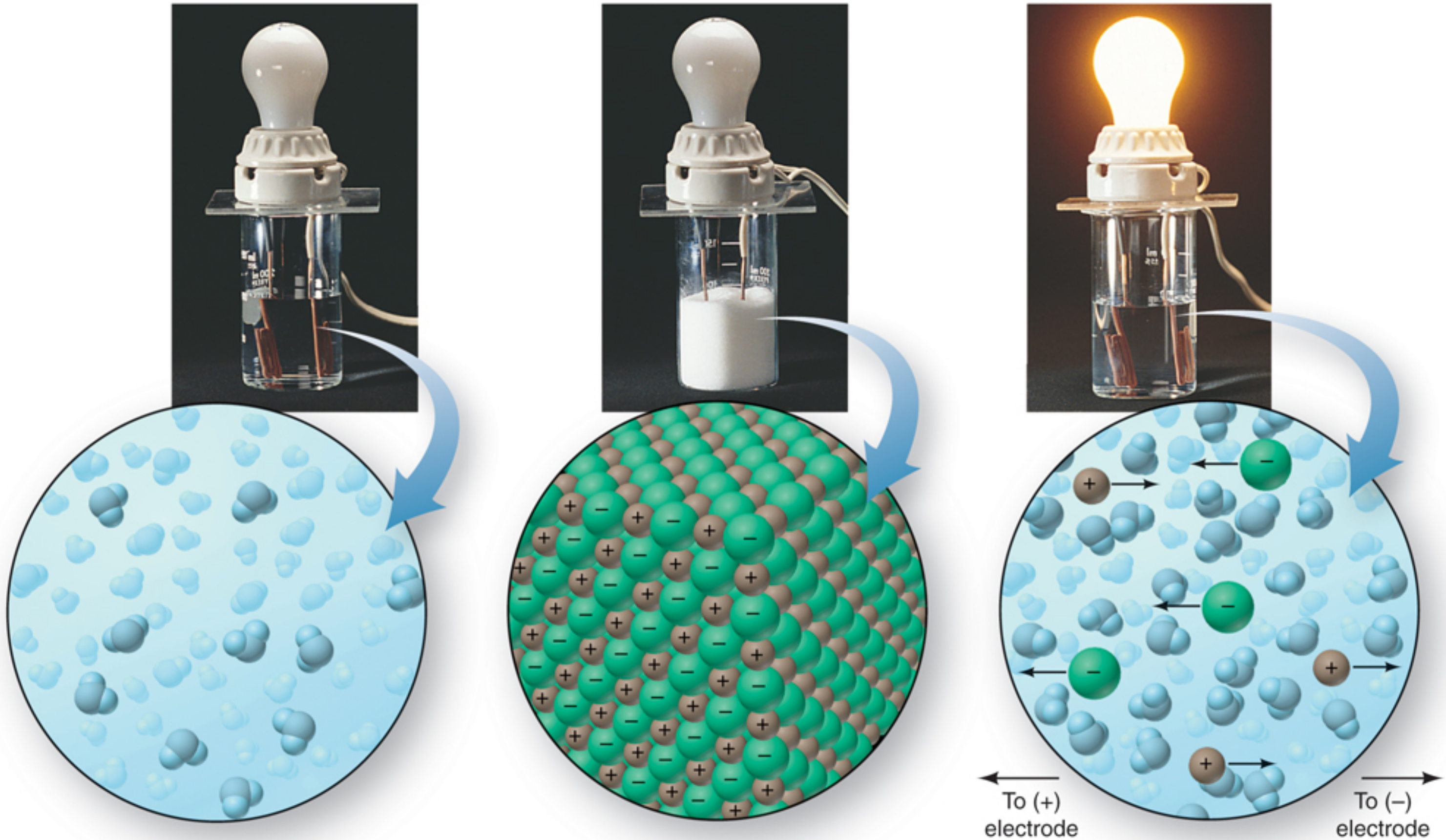


Figure 4.3 The electrical conductivity of ionic solutions.



A *Distilled water does not conduct a current.*

B *Positive and negative ions fixed in a solid do not conduct a current.*

C *In solution, positive and negative ions move and conduct a current.*

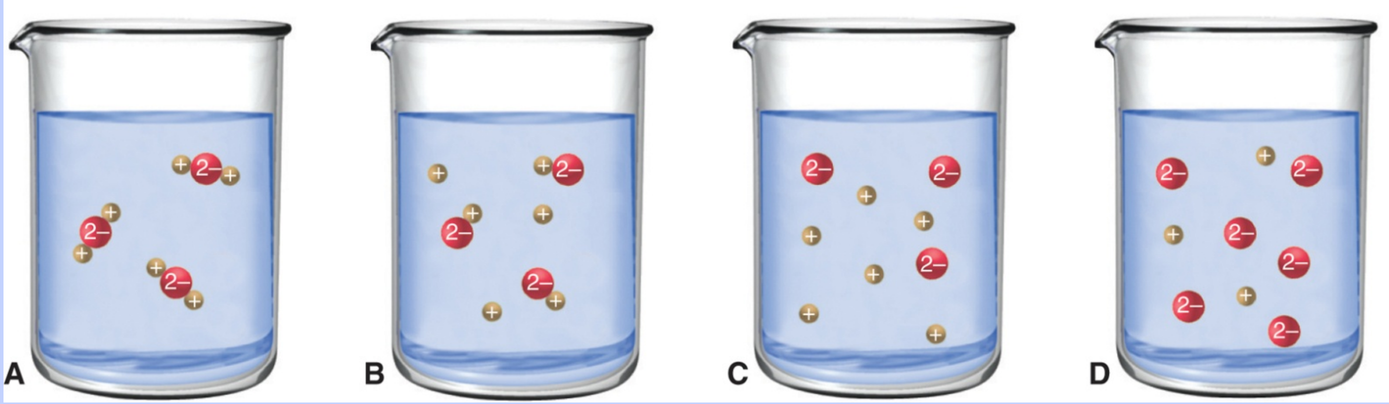
Sample Problem 4.1

Using Molecular Scenes to Depict an Ionic Compound in Aqueous Solution

PROBLEM: The beakers shown below contain aqueous solutions of the strong electrolyte potassium sulfate.

(a) Which beaker best represents the compound in solution? (H_2O molecules are not shown).

(b) If each particle represents 0.10 mol, what is the total number of particles in solution?



Sample Problem 4.1

PLAN: (a) Determine the formula and write and equation for the dissociation of 1 mol of compound. Potassium sulfate is a strong electrolyte; it therefore dissociates completely in solution. *Remember that polyatomic ions remain intact in solution.*

(b) Count the number of separate particles in the relevant beaker, then multiply by 0.1 mol and by Avogadro's number.

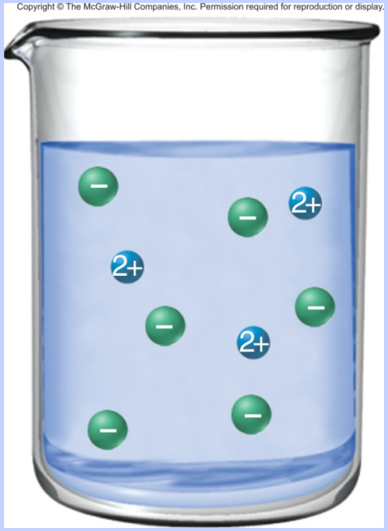
SOLUTION:

(a) The formula is K_2SO_4, so the equation for dissociation is:

$$\mathbf{K_2SO_4\ (\mathit{s}) \rightarrow 2K^+\ (\mathit{aq}) + SO_4^{2-}\ (\mathit{aq})}$$

Sample Problem 4.1

There should be 2 cations for every 1 anion; beaker C represents this correctly.



(b) Beaker C contains 9 particles, 6 K^+ ions and 3 SO_4^{2-} ions.

$$9 \times 0.1\ \cancel{\text{mol}} \times \frac{6.022\times10^{23}\ \text{particles}}{1\ \cancel{\text{mol}}} \boxed{= \mathbf{5.420\times10^{23}\ particles}}$$

Sample Problem 4.2 **Determining Amount (mol) of Ions in Solution**

PROBLEM: What amount (mol) of each ion is in each solution?

(a) 5.0 mol of ammonium sulfate dissolved in water

(b) 78.5 g of cesium bromide dissolved in water

(c) 7.42×10^{22} formula units of copper(II) nitrate dissolved in water

(d) 35 mL of 0.84 *mol/L* zinc chloride

PLAN: Write an equation for the dissociation of 1 mol of each compound. Use this information to calculate the actual number of moles represented by the given quantity of substance in each case.

Sample Problem 4.2

SOLUTION:

(a) The formula is $(NH_4)_2SO_4$ so the equation for dissociation is:

$$\mathbf{(NH_4)_2SO_4\ (\mathit{s}) \rightarrow 2NH_4^+\ (\mathit{aq}) + SO_4^{2-}\ (\mathit{aq})}$$

$$5.0\ \cancel{\text{mol } (NH_4)_2SO_4} \times \frac{2\ \text{mol } NH_4^+}{1\ \cancel{\text{mol } (NH_4)_2SO_4}} \quad \boxed{= 10.\ \text{mol } NH_4^+}$$

$$5.0\ \cancel{\text{mol } (NH_4)_2SO_4} \times \frac{1\ \text{mol } SO_4^{2-}}{1\ \cancel{\text{mol } (NH_4)_2SO_4}} \quad \boxed{= 5.0\ \text{mol } NH_4^+}$$

Sample Problem 4.2

SOLUTION:

(b) The formula is CsBr so the equation for dissociation is:

$$\mathbf{CsBr}\ (\boldsymbol{s}) \rightarrow \mathbf{Cs^+}\ (\boldsymbol{aq}) + \mathbf{Br^-}\ (\boldsymbol{aq})$$

$$78.5\ \cancel{\text{g CsBr}} \times \frac{1\ \cancel{\text{mol CsBr}}}{212.8\ \cancel{\text{g CsBr}}} \times \frac{1\ \text{mol Cs}^+}{1\ \cancel{\text{mol CsBr}}} \ \boxed{= \mathbf{0.369\ mol\ Cs^+}}$$

There is one Cs^+ ion for every Br^- ion, so the number of moles of Br^- is also equation to **0.369 mol.**

Sample Problem 4.2

SOLUTION:

(c) The formula is $Cu(NO_3)_2$ so the formula for dissociation is:

$$\mathbf{Cu(NO_3)_2\ (\mathit{s}) \rightarrow Cu^{2+}\ (\mathit{aq}) + 2NO_3^-\ (\mathit{aq})}$$

$$7.42\times10^{22}\ \cancel{\text{formula units}}\ Cu(NO_3)_2 \times \frac{1\ \text{mol}}{6.022\times10^{23}\ \cancel{\text{formula units}}}$$

$$= 0.123\ \text{mol}\ Cu(NO_3)_2$$

$$0.123\ \cancel{\text{mol}\ Cu(NO_3)_2} \times \frac{1\ \text{mol}\ Cu^{2+}}{1\ \cancel{\text{mol}\ Cu(NO_3)_2}}\ \boxed{\mathbf{= 0.123\ mol\ Cu^{2+}\ ions}}$$

There are 2 NO_3^- ions for every 1 Cu^{2+} ion, so there are **0.246 mol NO_3^- ions.**

Sample Problem 4.2

SOLUTION:

(d) The formula is $ZnCl_2$ so the formula for dissociation is:

$$\mathbf{ZnCl_2\ (\mathit{s}) \rightarrow Zn^{2+}\ (\mathit{aq}) + 2Cl^-\ (\mathit{aq})}$$

$$35\ \cancel{\text{mL soln}} \times \frac{1\ \cancel{\text{L}}}{10^3\ \cancel{\text{mL}}} \times \frac{0.84\ \text{mol}\ ZnCl_2}{1\ \cancel{\text{L soln}}} \quad \boxed{= \mathbf{2.9 \times 10^{-2}\ mol\ ZnCl_2}}$$

$$2.9 \times 10^{-2}\ \cancel{\text{mol}\ ZnCl_2} \times \frac{2\ \text{mol}\ Cl^-}{1\ \cancel{\text{mol}\ ZnCl_2}} \quad \boxed{= \mathbf{5.8 \times 10^{-2}\ mol\ Cl^-}}$$

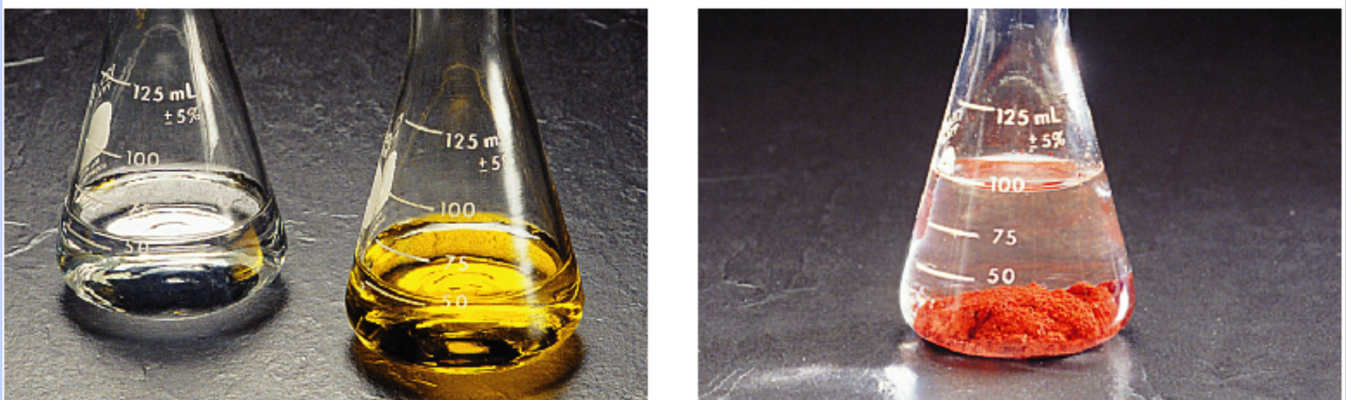
There is 1 mol of Zn^{2+} ions for every 1 mol of $ZnCl_2$, so there are **2.9×10^{-2} mol Zn^{2+} ions**.

Writing Equations for Aqueous Ionic Reactions

The **molecular equation** shows all reactants and products as if they were *intact, undissociated compounds*.

This gives the least information about the species in solution.

$$\mathbf{2AgNO_3\,(\mathit{aq}) + Na_2CrO_4\,(\mathit{aq}) \rightarrow Ag_2CrO_4\,(\mathit{s}) + 2NaNO_3\,(\mathit{aq})}$$



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When solutions of silver nitrate and sodium chromate mix, a brick-red precipitate of silver chromate forms.

The **total ionic equation** shows all soluble ionic substances *dissociated into ions*.

This gives the most accurate information about species in solution.

$$2Ag^+ (aq) + 2NO_3^- (aq) + 2Na^+ (aq) + CrO_4^{2-} (aq) \longrightarrow Ag_2CrO_4 (s) + 2Na^+ (aq) + NO_3^- (aq)$$

Spectator ions are ions that are not involved in the actual chemical change. Spectator ions appear unchanged on both sides of the total ionic equation.

$$2Ag^+ (aq) + 2NO_3^- (aq) + 2Na^+ (aq) + CrO_4^{2-} (aq) \longrightarrow Ag_2CrO_4 (s) + 2Na^+ (aq) + 2NO_3^- (aq)$$

The **net ionic equation** eliminates the ***spectator ions*** and shows only the *actual chemical change*.

$$2Ag^{+}\,(aq) + CrO_4^{2-}\,(aq) \longrightarrow Ag_2CrO_4\,(s)$$

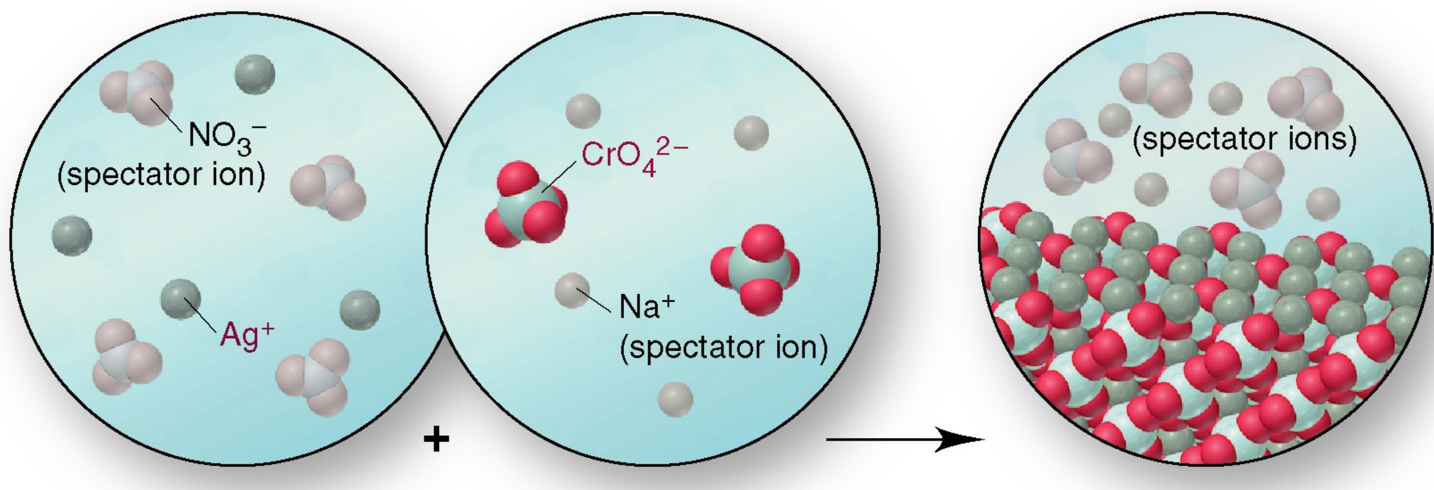
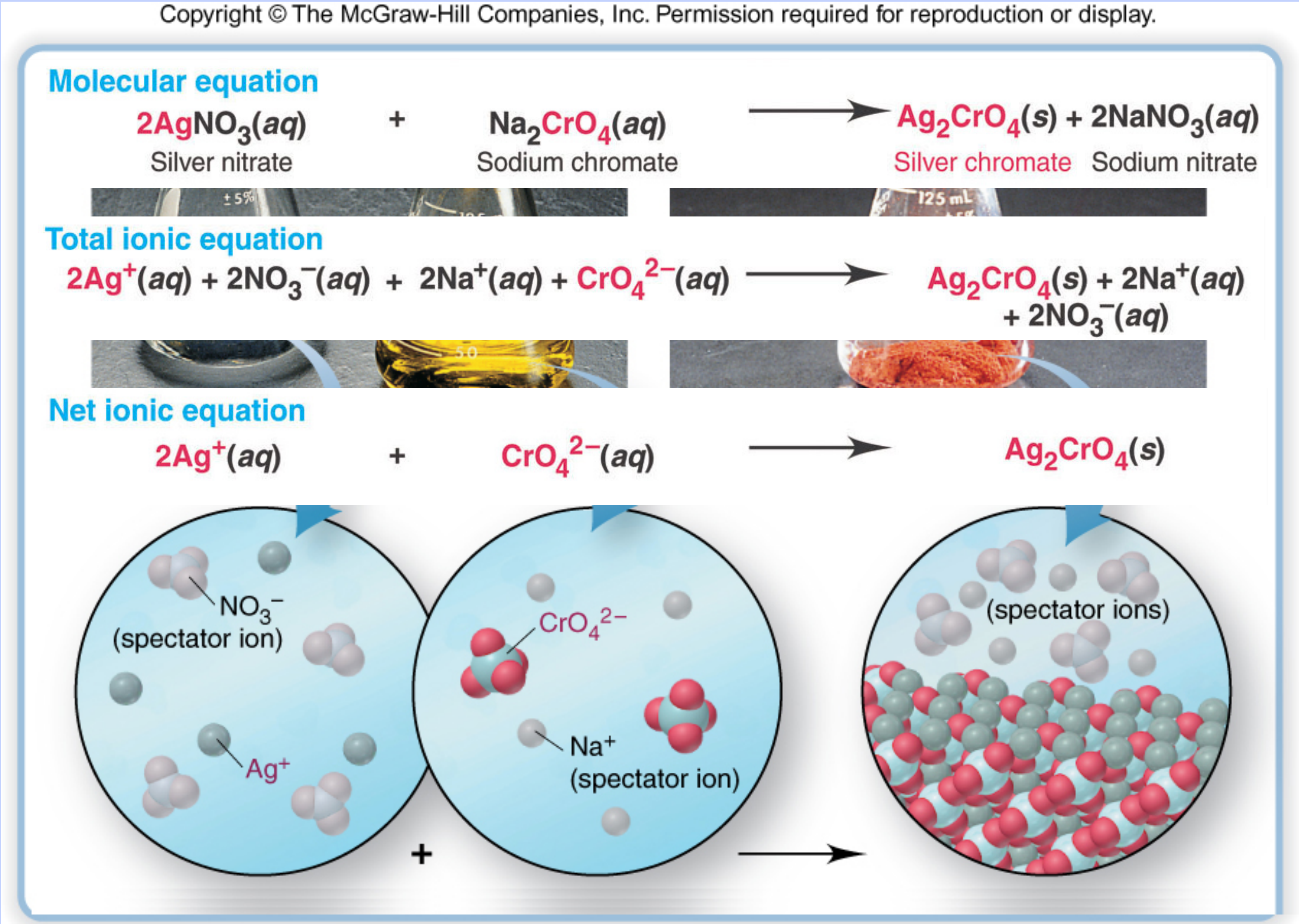


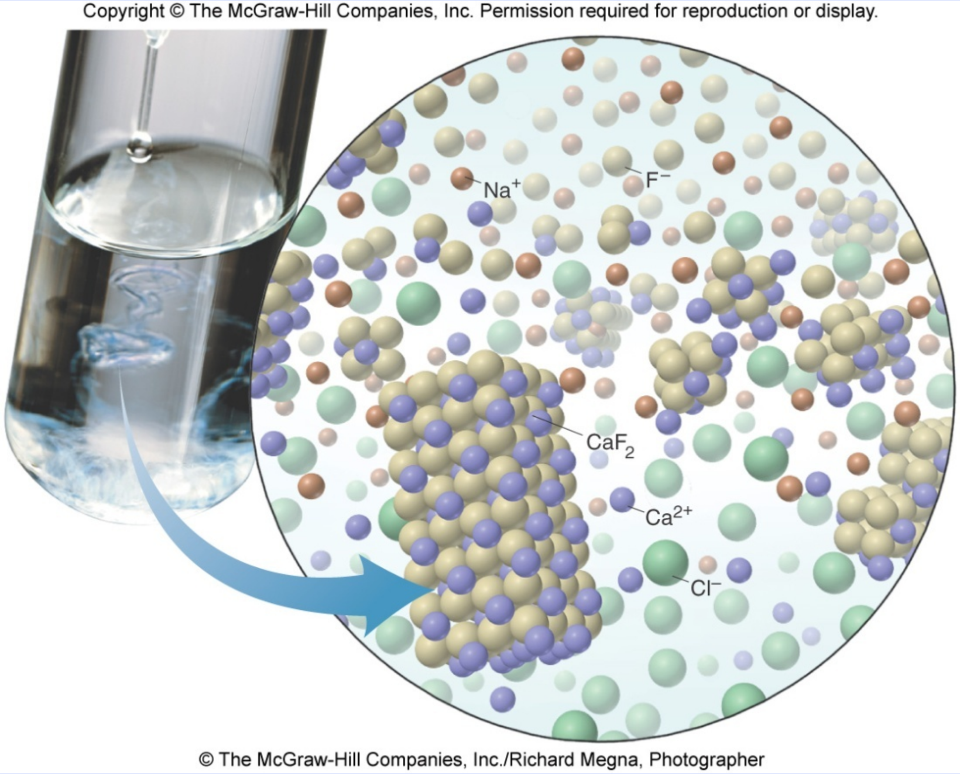
Figure 4.4 An aqueous ionic reaction and the three types of equations.



Precipitation Reactions

- In a **precipitation reaction** two soluble ionic compounds react to give an insoluble products, called a ***precipitate***.
- The precipitate forms through the net removal of ions from solution.
- It is possible for more than one precipitate to form in such a reaction.

Figure 4.5 The precipitation of calcium fluoride.



$$2NaF\ (aq) + CaCl_2\ (aq) \longrightarrow CaF_2\ (s) + 2NaCl\ (aq)$$

$$2\ Na^+\ (aq) + 2\ F^-\ (aq) + Ca^{2+}\ (aq) + 2\ Cl^-\ (aq) \longrightarrow CaF_2(s) + 2\ Na^+\ (aq) + 2\ Cl^-\ (aq)$$

$$2\ NaF(aq) + CaCl_2\ (aq) \longrightarrow CaF_2(s) + 2\ NaCl\ (aq)$$

Figure 4.6 The precipitation of PbI_2, a metathesis reaction.

$$2NaI\ (aq) + Pb(NO_3)_2\ (aq) \longrightarrow PbI_2\ (s) + NaNO_3\ (aq)$$

$$2Na^+(aq) + 2I^-\ (aq) + Pb^{2+}\ (aq) + 2NO_3^-\ (aq)$$
$$\longrightarrow PbI_2\ (s) + 2Na^+\ (aq) + 2NO_3^-(aq)$$

$$Pb^{2+}\ (aq) + 2I^-\ (aq) \longrightarrow PbI_2\ (s)$$

Precipitation reactions are also called **double displacement** reactions or **metathesis**.

$$2NaI\ (aq) + Pb\ (NO_3)_2\ (aq) \longrightarrow PbI_2\ (s) + 2NaNO_3\ (aq)$$

Ions exchange partners and a precipitate forms, so there is an exchange of bonds between reacting species.

Predicting Whether a Precipitate Will Form

- Note the ions present in the reactants.
- Consider all possible cation-anion combinations.
- Use the ***solubility rules*** to decide whether any of the ion combinations is insoluble.
  - An insoluble combination identifies the precipitate that will form.

Table 4.1 Solubility Rules for Ionic Compounds in Water

Soluble Ionic Compounds

1. All common compounds of Group 1A(1) ions (Li^+, Na^+, K^+, etc.) and ammonium ion (NH_4^+) are soluble.
2. All common nitrates (NO_3^-), acetates (CH_3COO^- or $C_2H_3O_2^-$) and most perchlorates (ClO_4^-) are soluble.
3. All common chlorides (Cl^-), bromides (Br^-) and iodides (I^-) are soluble, *except* those of Ag^+, Pb^{2+}, Cu^+, and Hg_2^{2+}. All common fluorides (F^-) are soluble *except* those of Pb^{2+} and Group 2A(2).
4. All common sulfates (SO_2^{2-}) are soluble, *except* those of Ca^{2+}, Sr^{2+}, Ba^{2+}, Ag^+, and Pb^{2+}.

Insoluble Ionic Compounds

1. All common metal hydroxides are insoluble, *except* those of Group 1A(1) and the larger members of Group 2A(2)(beginning with Ca^{2+}).
2. All common carbonates (CO_3^{2-}) and phosphates (PO_4^{3-}) are insoluble, *except* those of Group 1A(1) and NH_4^+.
3. All common sulfides are insoluble *except* those of Group 1A(1), Group 2A(2) and NH_4^+.

Sample Problem 4.3

Predicting Whether a Precipitation Reaction Occurs; Writing Ionic Equations

PROBLEM: Predict whether or not a reaction occurs when each of the following pairs of solutions are mixed. If a reaction does occur, write balanced molecular, total ionic, and net ionic equations, and identify the spectator ions.

(a) potassium fluoride (*aq*) + strontium nitrate (*aq*) →

(b) ammonium perchlorate (*aq*) + sodium bromide (*aq*) →

PLAN: Note reactant ions, write the possible cation-anion combinations, and use Table 4.1 to decide if the combinations are insoluble.
Write the appropriate equations for the process.

Sample Problem 4.3

SOLUTION: **(a)** The reactants are KF and $Sr(NO_3)_2$. The possible products are KNO_3 and SrF_2. KNO_3 is soluble, but SrF_2 is an insoluble combination.

Molecular equation:

$$2KF\ (aq) + Sr(NO_3)_2\ (aq) \rightarrow 2\ KNO_3\ (aq) + SrF_2\ (s)$$

Total ionic equation:

$$2K^+\ (aq) + 2F^-\ (aq) + Sr^{2+}\ (aq) + 2NO_3^-\ (aq) \rightarrow 2K^+\ (aq) + 2NO_3^-\ (aq) + SrF_2\ (s)$$

K^+ and NO_3^- are spectator ions

Net ionic equation:

$$Sr^{2+}\ (aq) + 2F^-\ (aq) \rightarrow SrF_2\ (s)$$

Sample Problem 4.3

SOLUTION: **(b)** The reactants are NH_4ClO_4 and NaBr. The possible products are NH_4Br and $NaClO_4$. Both are soluble, so no precipitate forms.

Molecular equation:

$$NH_4ClO_4\ (aq) + NaBr\ (aq) \rightarrow NH_4Br\ (aq) + NaClO_4\ (aq)$$

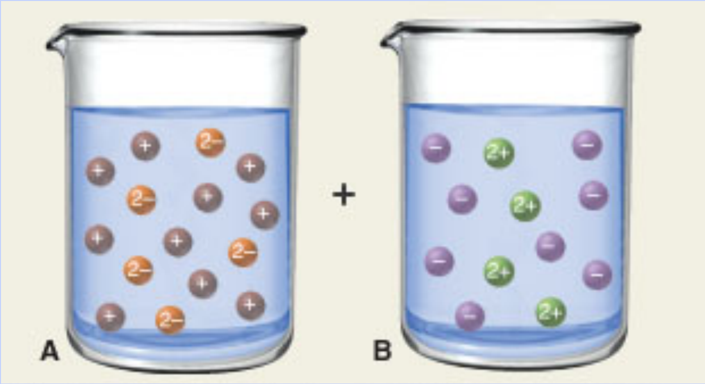
Total ionic equation:

$$NH_4^+\ (aq) + ClO_4^-\ (aq) + Na^+\ (aq) + Br^-\ (aq) \rightarrow NH_4^+\ (aq) + Br^-\ (aq) + Na^+\ (aq) + ClO_4^-\ (aq)$$

All ions are spectator ions and there is no net ionic equation.

Sample Problem 4.4 **Using Molecular Depictions in Precipitation Reactions**

PROBLEM: The following molecular views show reactant solutions for a precipitation reaction (with H_2O molecules omitted for clarity).



(a) Which compound is dissolved in beaker A: KCl, Na_2SO_4, $MgBr_2$, or Ag_2SO_4?
(b) Which compound is dissolved in beaker B: NH_4NO_3, $MgSO_4$, $Ba(NO_3)_2$, or CaF_2?

Sample Problem 4.4

PLAN: Note the number and charge of each kind of ion and use Table 4.1 to determine the ion combinations that are soluble.

SOLUTION:

(a) Beaker A contains two 1+ ion for each 2- ion. Of the choices given, only Na_2SO_4 and Ag_2SO_4 are possible. Na_2SO_4 is soluble while Ag_2SO_4 is not.

Beaker A therefore contains Na_2SO_4.

(b) Beaker B contains two 1- ions for each 2+ ion. Of the choices given, only CaF_2 and $Ba(NO_3)_2$ match this description. CaF_2 is not soluble while $Ba(NO_3)_2$ is soluble.

Beaker B therefore contains $Ba(NO_3)_2$.

Sample Problem 4.4

PROBLEM: (c) Name the precipitate and spectator ions when solutions A and B are mixed, and write balanced molecular, total ionic, and net ionic equations for this process.

(d) If each particle represents 0.010 mol of ions, what is the maximum mass (g) of precipitate that can form (assuming complete reaction)?

PLAN: (c) Consider the cation-anion combinations from the two solutions and use Table 4.1 to decide if either of these is insoluble.

SOLUTION: The reactants are $Ba(NO_3)_2$ and Na_2SO_4. The possible products are $BaSO_4$ and $NaNO_3$. $BaSO_4$ is insoluble while $NaNO_3$ is soluble.

Sample Problem 4.4

Molecular equation:

$Ba(NO_3)_2$ (*aq*) + Na_2SO_4 (*aq*) → $2NaNO_3$ (*aq*) + $BaSO_4$ (*s*)

Total ionic equation:

Ba^{2+} (*aq*) + $2NO_3^-$ (*aq*) + $2Na^+$ (*aq*) + SO_4^{2-} (*aq*) → $2Na^+$ (*aq*) + $2NO_3^-$ (*aq*) + $BaSO_4$ (*s*)

Na^+ and NO_3^- are spectator ions

Net ionic equation:

Ba^{2+} (*aq*) + SO_4^{2-} (*aq*) → $BaSO_4$ (*s*)

Sample Problem 4.4

PLAN: (d) Count the number of each kind of ion that combines to form the solid. Multiply the number of each reactant ion by 0.010 mol and calculate the mol of product formed from each. Decide which ion is the limiting reactant and use this information to calculate the mass of product formed.

SOLUTION: There are 4 Ba^{2+} particles and 5 SO_4^{2-} particles depicted.

$$4\ \cancel{Ba^{2+}\ \text{particles}} \times \frac{0.010\ \cancel{\text{mol}\ Ba^{2+}}}{1\ \cancel{\text{particle}}} \times \frac{1\ \text{mol}\ BaSO_4}{1\ \cancel{\text{mol}\ Ba^{2+}}} = 0.040\ \text{mol}\ BaSO_4$$

$$4\ \cancel{SO_4^{2-}\ \text{particles}} \times \frac{0.010\ \cancel{\text{mol}\ SO_4^{2-}}}{\cancel{1\ \text{particle}}} \times \frac{1\ \text{mol}\ BaSO_4}{1\ \cancel{\text{mol}\ SO_4^{2-}}} = 0.050\ \text{mol}\ BaSO_4$$

Sample Problem 4.4

Ba^{2+} ion is the limiting reactant, since it yields less $BaSO_4$.

$$0.040\ \cancel{\text{mol BaSO}_4} \times \frac{233.4\ \text{g BaSO}_4}{1\ \cancel{\text{mol BaSO}_4}}\ \boxed{= \mathbf{9.3\ g\ BaSO_4}}$$

Acid-Base Reactions

An **acid** is a substance that produces H^+ ions when dissolved in H_2O.

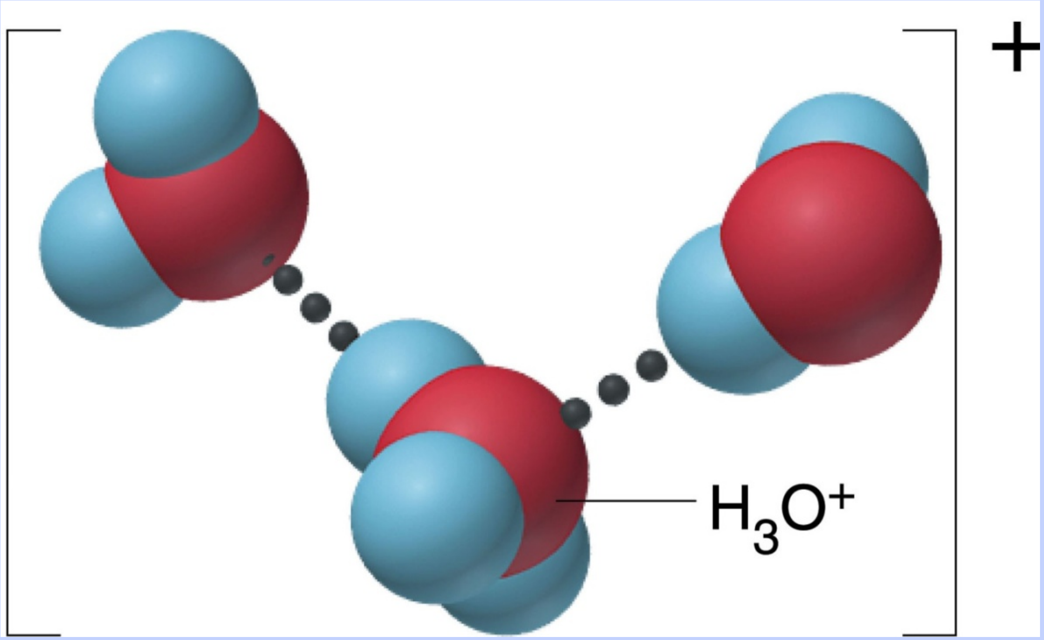
$$HX \xrightarrow{H_2O} H^+ (aq) + X^- (aq)$$

A **base** is a substance that produces OH^- ions when dissolved in H_2O.

$$MOH \xrightarrow{H_2O} mol/L^+ (aq) + OH^- (aq)$$

An **acid-base reaction** is also called a **neutralization** reaction.

Figure 4.7 The H^+ ion as a solvated hydronium ion.



H^+ interacts strongly with H_2O, forming H_3O^+ in aqueous solution.

Table 4.2 Selected Acids and Bases

Acids	Bases
Strong	**Strong**
hydrochloric acid, HCl	sodium hydroxide, $NaOH$
hydrobromic acid, HBr	potassium hydroxide, KOH
hydriodic acid, HI	calcium hydroxide, $Ca(OH)_2$
nitric acid, HNO_3	strontium hydroxide, $Sr(OH)_2$
sulfuric acid, H_2SO_4	barium hydroxide, $Ba(OH)_2$
perchloric acid, $HClO_4$	
Weak	**Weak**
hydrofluoric acid, HF	ammonia, NH_3
phosphoric acid, H_3PO_4	
acetic acid, CH_3COOH (or $HC_2H_3O_2$)	

Figure 4.8 Acids and bases as electrolytes.

Strong acids and strong bases dissociate completely into ions in aqueous solution.

They are ***strong electrolytes*** and conduct well in solution.

A Strong acid (or base) = strong electrolyte

Figure 4.8 Acids and bases as electrolytes.

Weak acids and weak bases dissociate very little into ions in aqueous solution.

They are ***weak electrolytes*** and conduct poorly in solution.

B Weak acid (or base) = weak electrolyte

Sample Problem 4.5 **Determining the Number of H^+ (or OH^-) Ions in Solution**

PROBLEM: How many $H^+(aq)$ ions are in 25.3 mL of 1.4 *mol/L* nitric acid?

PLAN: Use the volume and concentration to determine the mol of acid present. Since HNO_3 is a strong acid, moles acid = moles H^+.

volume of HNO_3

↓ **convert mL to L and multiply by *mol/L***

mol of HNO_3

↓ **mole of H^+ = mol of HNO_3**

mol of H^+

↓ **multiply by Avogadro's number**

number of H^+ ions

Sample Problem 4.5

SOLUTION:

$$35.3\ \cancel{\text{mL soln}} \times \frac{1\ \cancel{\text{L}}}{10^3\ \cancel{\text{mL}}} \times \frac{1.4\ \text{mol HNO}_3}{1\ \cancel{\text{L soln}}} = 0.035\ \text{mol HNO}_3$$

One mole of $H^+(aq)$ is released per mole of nitric acid (HNO_3).

$$HNO_3\,(aq) \xrightarrow{H_2O} H^+\,(aq) + NO_3^-\,(aq)$$

$$= 0.035\ \cancel{\text{mol HNO}_3} \times \frac{1\ \cancel{\text{mol}}\ \text{H}^+}{1\ \cancel{\text{mol HNO}_3}} \times \frac{6.022\text{x}10^{23}\ \text{ions}}{1\ \cancel{\text{mol}}}\quad \boxed{\mathbf{= 2.1x10^{22}\ H^+\ ions}}$$

Sample Problem 4.6 — **Writing Ionic Equations for Acid-Base Reactions**

PROBLEM: Write balanced molecular, total ionic, and net ionic equations for the following acid-base reactions and identify the spectator ions.

(a) hydrochloric acid (*aq*) + potassium hydroxide (*aq*) →

(b) strontium hydroxide (*aq*) + perchloric acid (*aq*) →

(c) barium hydroxide (*aq*) + sulfuric acid (*aq*) →

PLAN: All reactants are strong acids and bases (see Table 4.2). The product in each case is H_2O and an **ionic salt**.
Write the molecular reaction in each case and use the solubility rules to determine if the product is soluble or not.

Sample Problem 4.6

SOLUTION:

(a) hydrochloric acid (*aq*) + potassium hydroxide (*aq*) →

Molecular equation:

$HCl\ (aq) + KOH\ (aq) \rightarrow KCl\ (aq) + H_2O\ (l)$

Total ionic equation:

$H^+\ (aq) + Cl^-\ (aq) + K^+\ (aq) + OH^-\ (aq) \rightarrow K^+\ (aq) + Cl^-\ (aq) + H_2O\ (l)$

Net ionic equation:

$H^+\ (aq) + OH^-\ (aq) \rightarrow H_2O\ (l)$

Spectator ions are K^+ and Cl^-

Sample Problem 4.6

SOLUTION:

(b) strontium hydroxide (*aq*) + perchloric acid (*aq*) →

Molecular equation:

$Sr(OH)_2\ (aq) + 2HClO_4\ (aq) \rightarrow Sr(ClO_4)_2\ (aq) + 2H_2O\ (l)$

Total ionic equation:

$Sr^{2+}\ (aq) + 2OH^-\ (aq) + 2H^+\ (aq) + 2ClO_4^-\ (aq) \rightarrow Sr^{2+}\ (aq) + 2ClO_4^-\ (aq) + 2H_2O\ (l)$

Net ionic equation:

$2H^+\ (aq) + 2OH^-\ (aq) \rightarrow 2H_2O\ (l)$ or $H^+\ (aq) + OH^-\ (aq) \rightarrow H_2O\ (l)$

Spectator ions are Sr^{2+} and ClO_4^-

Sample Problem 4.6

SOLUTION:

(c) barium hydroxide (*aq*) + sulfuric acid (*aq*) →

Molecular equation:

$Ba(OH)_2\ (aq) + H_2SO_4\ (aq) \rightarrow BaSO_4\ (s) + 2H_2O\ (l)$

Total ionic equation:

$Ba^{2+}\ (aq) + 2OH^-\ (aq) + 2H^+\ (aq) + SO_4^{2-}\ (aq) \rightarrow BaSO_4\ (s) + H_2O\ (l)$

The net ionic equation is the **same** as the total ionic equation since there are **no spectator ions**.

This reaction is both a neutralization reaction and a precipitation reaction.

Figure 4.9 An aqueous strong acid-strong base reaction as a proton-transfer process.

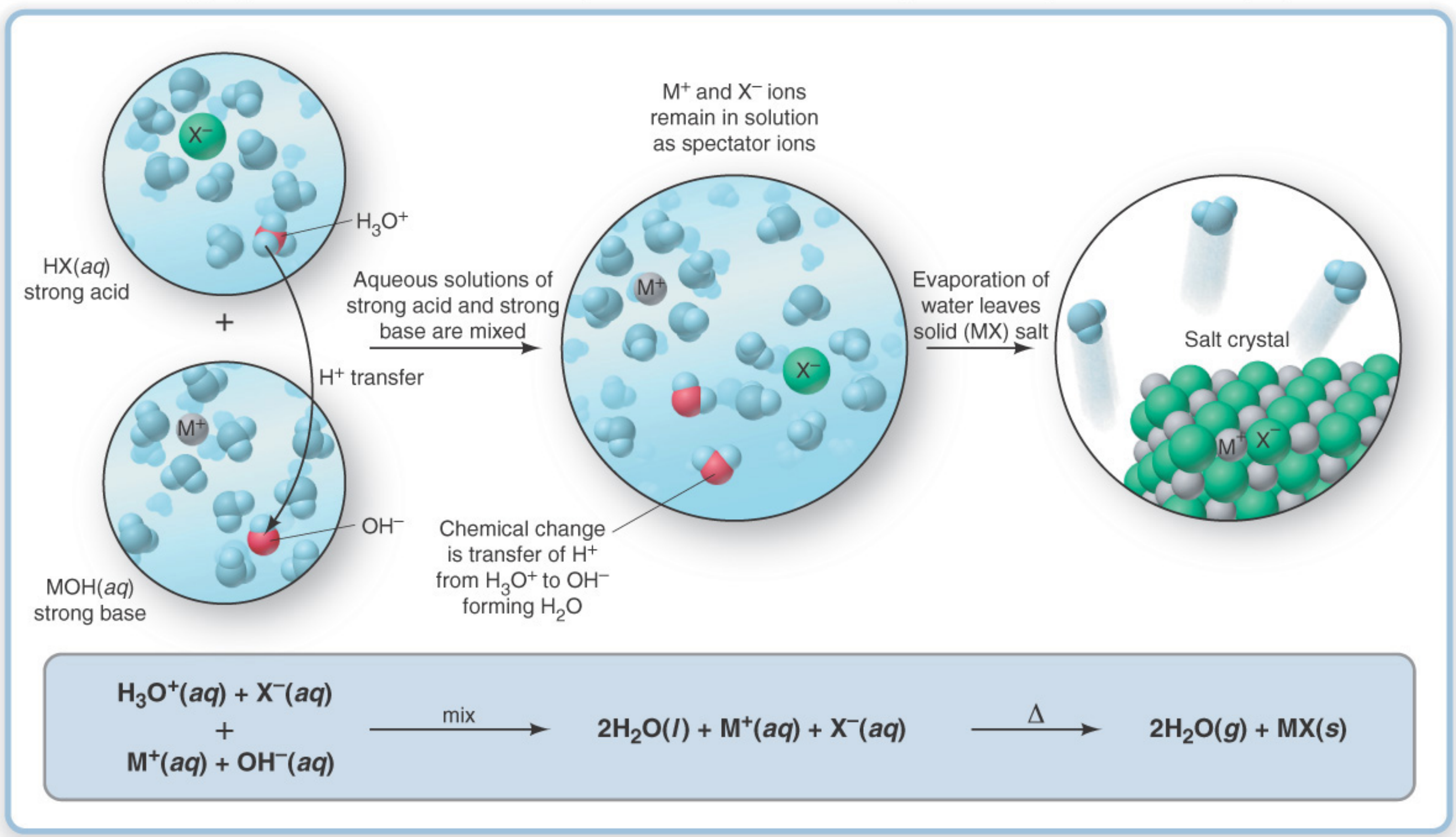


Figure 4.11 A gas-forming reaction with a weak acid.

Molecular equation

$$NaHCO_3\,(aq) + CH_3COOH(aq) \longrightarrow CH_3COONa\,(aq) + CO_2\,(g) + H_2O\,(l)$$

Total ionic equation

$$Na^+\,(aq) + HCO_3^-\,(aq) + CH_3COOH\,(aq) \longrightarrow CH_3COO^-\,(aq) + Na^+\,(aq) + CO_2\,(g) + H_2O\,(l)$$

Net ionic equation

$$HCO_3^-(aq) + CH_3COOH\,(aq) \longrightarrow CH_3COO^-\,(aq) + CO_2\,(g) + H_2O\,(l)$$

Sample Problem 4.7

Writing Proton-Transfer Equations for Acid-Base Reactions

PROBLEM: Write balanced total and net ionic equations for the following reactions and use curved arrows to show how the proton transfer occurs.

(a) hydriodic acid (*aq*) + calcium hydroxide (*aq*) →

Give the name and formula of the salt present when the water evaporates.

(b) potassium hydroxide (*aq*) + propionic acid (*aq*) →

Note that propionic acid is a weak acid. Be sure to identify the spectator ions in this reaction.

Sample Problem 4.7

PLAN: In (a) the reactants are a strong acid and a strong base. The acidic species is therefore H_3O^+, which transfers a proton to the OH^- from the base.

SOLUTION:

Total Ionic Equation:

$$2H_3O^+\ (aq) + 2I^-\ (aq) + Ca^{2+}\ (aq) + 2OH^-\ (aq) \rightarrow 2I^-\ (aq) + Ca^{2+}\ (aq) + 4H_2O\ (l)$$

H^+ transferred to OH^-

Net Ionic Equation:

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

When the water evaporates, the salt remaining is CaI_2, calcium iodide.

Sample Problem 4.7

PLAN: In (b) the acid is weak; therefore it does not dissociate much and largely exists as intact molecules in solution.

SOLUTION:

Total Ionic Equation:

H^+ transferred to OH^-

$K^+ (aq) + OH^- (aq) + CH_3CH_2COOH (aq) \rightarrow K^+ (aq) + H_2O (l) + CH_3CH_2COO^- (aq)$

Net Ionic Equation:

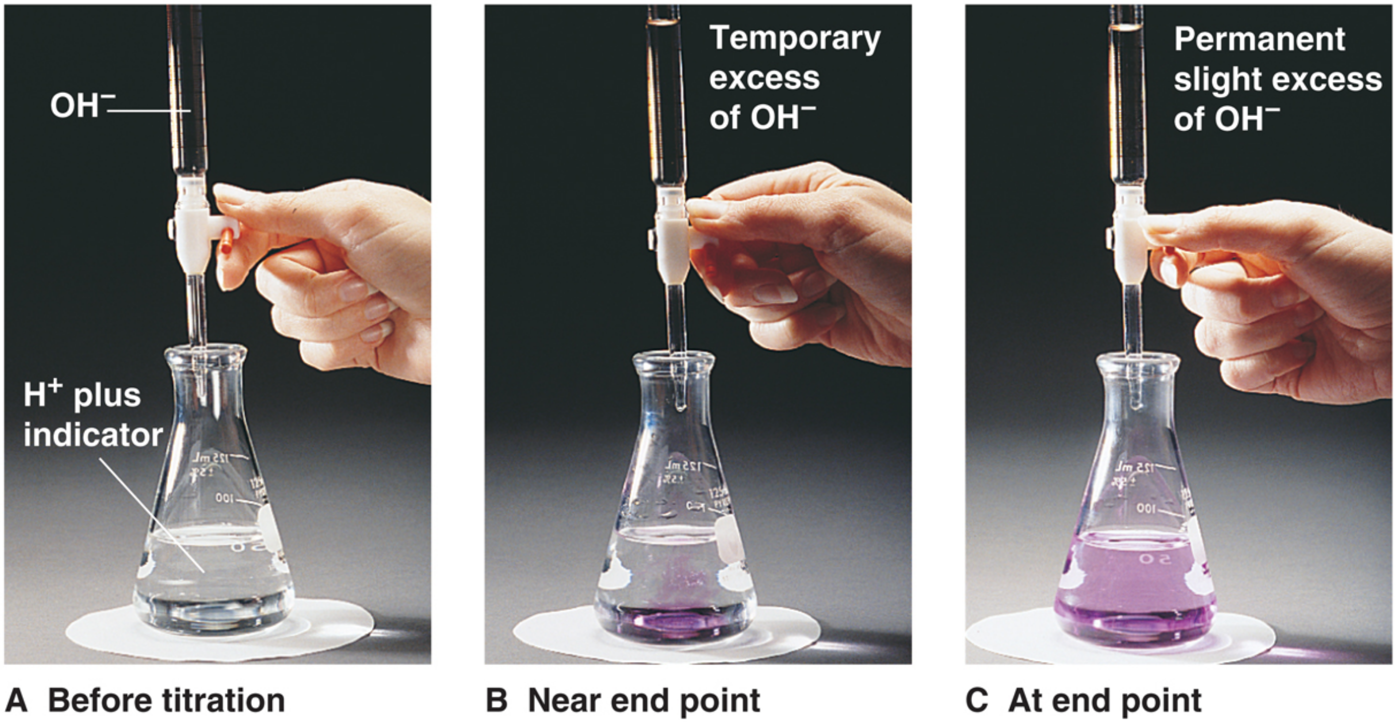
$CH_3CH_2COOH (aq) + OH^- (aq) \rightarrow CH_3CH_2COO^- (aq) + H_2O (l)$

K^+ is the only spectator ion in the reaction.

Acid-Base Titrations

- In a ***titration***, the concentration of one solution is used to determine the concentration of another.
- In an acid-base titration, a standard solution of base is usually added to a sample of acid of unknown concentration.
- An ***acid-base indicator*** has different colors in acid and base, and is used to monitor the reaction progress.
- At the ***equivalence point***, the mol of H^+ from the acid equals the mol of OH^- ion produced by the base.
  - Amount of H^+ ion in flask = amount of OH^- ion added
- The ***end point*** occurs when there is a slight excess of base and the indicator changes color permanently.

Figure 4.11 An acid-base titration.



$$H^+(aq) + X^-(aq) + M^+(aq) + OH^-(aq) \longrightarrow H_2O(l) + M^+(aq) + X^-(aq)$$

Sample Problem 4.8

Finding the Concentration of Acid from a Titration

PROBLEM: A 50.00 mL sample of HCl is titrated with 0.1524 *mol/L* NaOH. The buret reads 0.55 mL at the start and 33.87 mL at the end point. Find the concentration of the HCl solution.

PLAN: Write a balanced equation for the reaction. Use the volume of base to find amount of OH^- (mol), then amount of H^+ (mol) and finally concentration (*mol/L*) for the acid.

volume of base (difference in buret readings)

↓ **multiply by concentration (*mol/L*) of base**

amount of OH^-

↓ **use mole ratio to find amount of acid**

amount of H^+ and acid

↓ **divide by volume (L) of acid**

concentration (*mol/L*) of acid

Sample Problem 4.8

SOLUTION: $NaOH\ (aq) + HCl\ (aq) \rightarrow NaCl\ (aq) + H_2O\ (l)$

volume of base = 33.87 mL – 0.55 mL = 33.32 mL

$$33.32\ \cancel{\text{mL soln}} \times \frac{1\ \cancel{\text{L}}}{10^3\ \cancel{\text{mL}}} \times \frac{0.1524\ \text{mol NaOH}}{1\ \cancel{\text{L soln}}} = 5.078\text{x}10^{-3}\ \text{mol NaOH}$$

Since 1 mol of HCl reacts with 1 mol NaOH, the amount of HCl = $5.078\text{x}10^{-3}$ mol.

$$\frac{5.078\text{x}10^{-3}\ \text{mol HCl}}{50.00\ \cancel{\text{mL}}} \times \frac{10^3\ \cancel{\text{mL}}}{1\ \text{L}} \quad \boxed{= 0.1016\ \textit{mol/L}\ \text{HCl}}$$

Oxidation-Reduction (Redox) Reactions

Oxidation is the ***loss*** of electrons.
The ***reducing agent*** loses electrons and is oxidized.

Reduction is the ***gain*** of electrons.
The ***oxidizing agent*** gains electrons and is reduced.

A **redox reaction** involves ***electron transfer***
Oxidation and reduction occur together.

Figure 4.12 The redox process in compound formation.

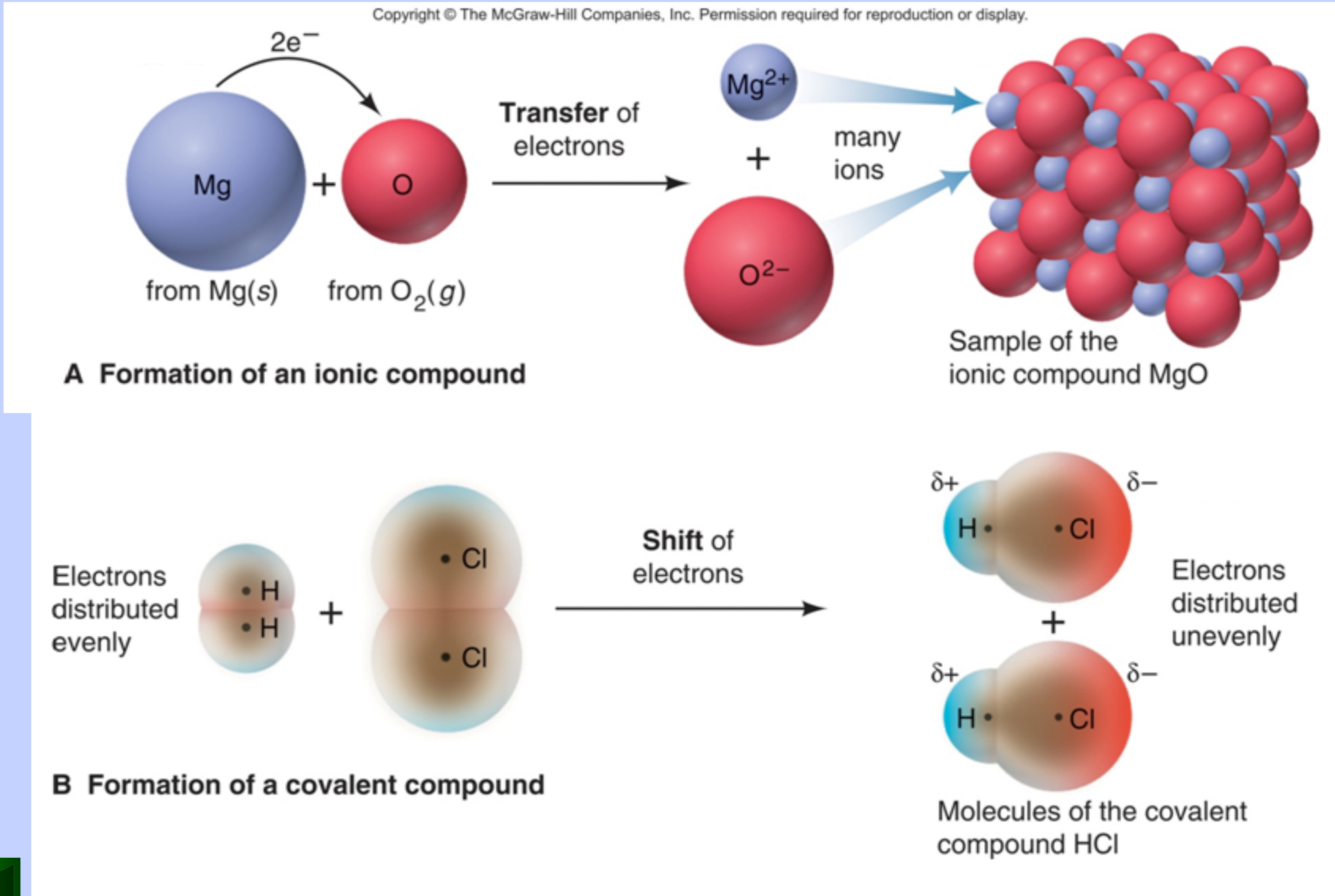


Table 4.3 Rules for Assigning an Oxidation Number (O.N.)

General rules

1. For an atom in its elemental form (Na, O_2, Cl_2, etc.): O.N. = 0
2. For a monoatomic ion: O.N. = ion charge
3. The sum of O.N. values for the atoms in a compound equals zero. The sum of O.N. values for the atoms in a polyatomic ion equals the ion's charge.

Rules for specific atoms or periodic table groups

1. For Group 1:	O.N. = +1 in all compounds
2. For Group 2:	O.N. = +2 in all compounds
3. For hydrogen:	O.N. = +1 in combination with nonmetals
4. For fluorine:	O.N. = -1 in combination with metals and boron
5. For oxygen:	O.N. = -1 in peroxides O.N. = -2 in all other compounds(except with F)
6. For Group 17:	O.N. = -1 in combination with metals, nonmetals (except O), and other halogens lower in the group

Sample Problem 4.9 **Determining the Oxidation Number of Each Element in a Compound (or Ion)**

PROBLEM: Determine the oxidation number (O.N.) of each element in these species:

(a) zinc chloride **(b)** sulfur trioxide **(c)** nitric acid

PLAN: The O.N.s of the ions in a polyatomic ion add up to the charge of the ion and the O.N.s of the ions in the compound add up to zero.

SOLUTION:

(a) $ZnCl_2$. The O.N. for zinc is +2 and that for chloride is -1.

(b) SO_3. Each oxygen is an oxide with an O.N. of -2.
The O.N. of sulfur must therefore be +6.

(c) HNO_3. H has an O.N. of +1 and each oxygen is -2.
The N must therefore have an O.N. of +5.

Sample Problem 4.10 Identifying Redox Reactions

PROBLEM: Use oxidation numbers to decide whether each of the following is a redox reaction or not.

(a) $CaO\,(s) + CO_2\,(g) \longrightarrow CaCO_3\,(s)$

(b) $4\,KNO_3\,(s) \longrightarrow 2\,K_2O(s) + 2\,N_2(g) + 5\,O_2(g)$

(c) $NaHSO_4\,(aq) + NaOH\,(aq) \longrightarrow Na_2SO_4\,(aq) + H_2O\,(l)$

PLAN: Use Table 4.3 to assign an O.N. to each atom. A change in O.N. for any atom indicates electron transfer.

SOLUTION:

(a) $$\underset{+2}{Ca}\underset{-2}{O}(s) + \underset{+2}{C}\underset{-2}{O_2}(g) \longrightarrow \underset{+2}{Ca}\underset{+2}{C}\underset{-2}{O_3}(s)$$

This is not a redox reaction, since no species change O.N.

Sample Problem 4.10

$$\textbf{(b)}\quad 4\,\underset{+1}{\mathrm{K}}\underset{+5}{\mathrm{N}}\underset{-2}{\mathrm{O_3}}(s) \longrightarrow 2\,\underset{+1}{\mathrm{K_2}}\underset{-2}{\mathrm{O}}(s) + 2\,\underset{0}{\mathrm{N_2}}(g) + 5\,\underset{0}{\mathrm{O_2}}(g)$$

This is a redox reaction.

N changes O.N. from +5 to 0 and is reduced.

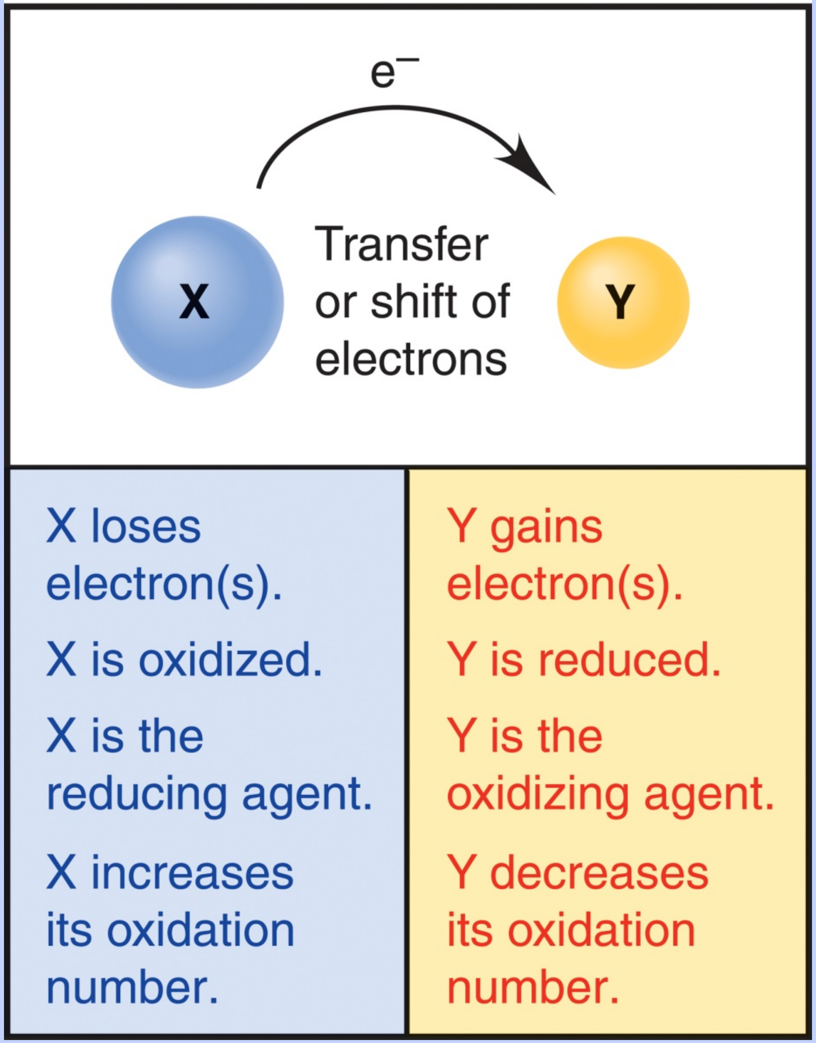
O changes O.N. from -2 to 0 and is oxidized.

Sample Problem 4.10

$$\text{(c)}\quad \underset{+1}{\mathrm{Na}}\,\underset{+1}{\mathrm{H}}\,\underset{+6}{\mathrm{S}}\,\underset{-2}{\mathrm{O}}{}_4\,(aq) + \underset{+1}{\mathrm{Na}}\,\underset{-2}{\mathrm{O}}\,\underset{+1}{\mathrm{H}}\,(aq) \longrightarrow \underset{+1}{\mathrm{Na}}{}_2\,\underset{+6}{\mathrm{S}}\,\underset{-2}{\mathrm{O}}{}_4\,(aq) + \underset{+1}{\mathrm{H}}{}_2\,\underset{-2}{\mathrm{O}}\,(l)$$

This is not a redox reaction since no species change O.N.

Figure 4.13 A summary of terminology for redox reactions.



Sample Problem 4.11 | **Identifying Oxidizing and Reducing Agents**

PROBLEM: Identify the oxidizing agent and reducing agent in each of the following reactions:

(a) $2Al\ (s) + 3H_2SO_4\ (aq) \longrightarrow Al_2(SO_4)_3\ (aq) + 3H_2(g)$

(b) $PbO\ (s) + CO\ (g) \longrightarrow Pb\ (s) + CO_2\ (g)$

(c) $2H_2\ (g) + O_2\ (g) \longrightarrow 2H_2O\ (g)$

PLAN: Assign an O.N. to each atom and look for those that change during the reaction.

The reducing agent contains an atom that is oxidized (increases in O.N.) while the oxidizing agent contains an atom that is reduced (decreases in O.N.).

Sample Problem 4.11

SOLUTION:

(a) $$2\underset{0}{Al}(s) + 3\underset{+1}{H_2}\underset{+6}{S}\underset{-2}{O_4}(aq) \longrightarrow \underset{+3}{Al_2}(\underset{+6}{S}\underset{-2}{O_4})_3(aq) + 3\underset{0}{H_2}(g)$$

-2 **+1**

Al changes O.N. from 0 to +3 and is *oxidized*.
Al is the *reducing* agent.

H changes O.N. from +1 to 0 and is *reduced*.
H_2SO_4 is the *oxidizing* agent.

Sample Problem 4.11

SOLUTION:

(b) $PbO\ (s) + CO\ (g) \longrightarrow Pb\ (s) + CO_2\ (g)$

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

Pb changes O.N. from +2 to 0 and is *reduced*.
PbO is the *oxidizing* agent.

C changes O.N. from +2 to +4 and is *oxidized*.
CO is the *reducing* agent.

Sample Problem 4.11

SOLUTION:

(c) $2H_2\,(g) + O_2\,(g) \longrightarrow 2H_2O\,(g)$

$$\begin{array}{cccc} 2H_2 & O_2 & H \text{ in } H_2O & O \text{ in } H_2O \\ \uparrow & \uparrow & \uparrow & \uparrow \\ \mathbf{0} & \mathbf{0} & \mathbf{+1} & \mathbf{-2} \end{array}$$

H_2 changes O.N. from 0 to +1 and is *oxidized*.
H_2 is the *reducing* agent.

O changes O.N. from 0 to -2 and is *reduced*.
O_2 is the *oxidizing* agent.

Balancing Redox Equations (oxidation number method)

1. Assign O.N.s to all atoms.
2. Identify the reactants that are oxidized and reduced.
3. Compute the numbers of electrons transferred, and draw tie-lines from each reactant atom to the product atom to show the change.
4. Multiply the numbers of electrons by factor(s) that make the electrons lost equal to the electrons gained.
5. Use the factor(s) as balancing coefficients.
6. Complete the balancing by inspection and add states of matter.

Sample Problem 4.12

Balancing Redox Equations by the Oxidation Number Method

PROBLEM: Use the oxidation number method to balance the following equations:

(a) $Cu\,(s) + HNO_3\,(aq) \longrightarrow Cu(NO_3)_2\,(aq) + NO_2\,(g) + H_2O\,(l)$

SOLUTION:

Assign oxidation numbers and identify oxidized and reduced species:

$$\textbf{(a)}\quad \underset{0}{Cu}\,(s) + \underset{+1}{H}\underset{+5}{N}\underset{-2}{O_3}\,(aq) \longrightarrow \underset{+2}{Cu}(\underset{+5}{N}\underset{-2}{O_3})_2\,(aq) + \underset{+4}{N}\underset{-2}{O_2}\,(g) + \underset{+1}{H_2}\underset{-2}{O}\,(l)$$

Sample Problem 4.12

loses 2e⁻; oxidation

$$Cu(s) + HNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + NO_2(g) + H_2O(l)$$

gains 1e⁻; reduction

Multiply to make e^- lost = e^- gained:

$$Cu\ (s) + \mathbf{2}HNO_3\ (aq) \longrightarrow Cu(NO_3)_2\ (aq) + \mathbf{2}NO_2\ (g) + H_2O\ (l)$$

Balance other atoms by inspection:

$$\boxed{Cu\ (s) + \mathbf{4}HNO_3\ (aq) \longrightarrow Cu(NO_3)_2\ (aq) + \mathbf{2}NO_2\ (g) + \mathbf{2}H_2O\ (l)}$$

Sample Problem 4.12

(b) $PbS\ (s) + O_2\ (g) \longrightarrow PbO\ (s) + SO_2\ (g)$

SOLUTION:

Assign oxidation numbers and identify oxidized and reduced species:

$$\textbf{(b)}\ \underset{+2}{\mathrm{Pb}}\underset{-2}{\mathrm{S}}\ (s) + \underset{0}{\mathrm{O}_2}\ (g) \longrightarrow \underset{+2}{\mathrm{Pb}}\underset{-2}{\mathrm{O}}\ (s) + \underset{+4}{\mathrm{S}}\underset{-2}{\mathrm{O}_2}\ (g)$$

+5

Sample Problem 4.12

$$\overbrace{PbS\ (s) + O_2\ (g) \longrightarrow PbO\ (s) + SO_2\ (g)}^{\textit{loses 6e}^-\textit{; oxidation}}$$

gains 2e⁻ per O; reduction

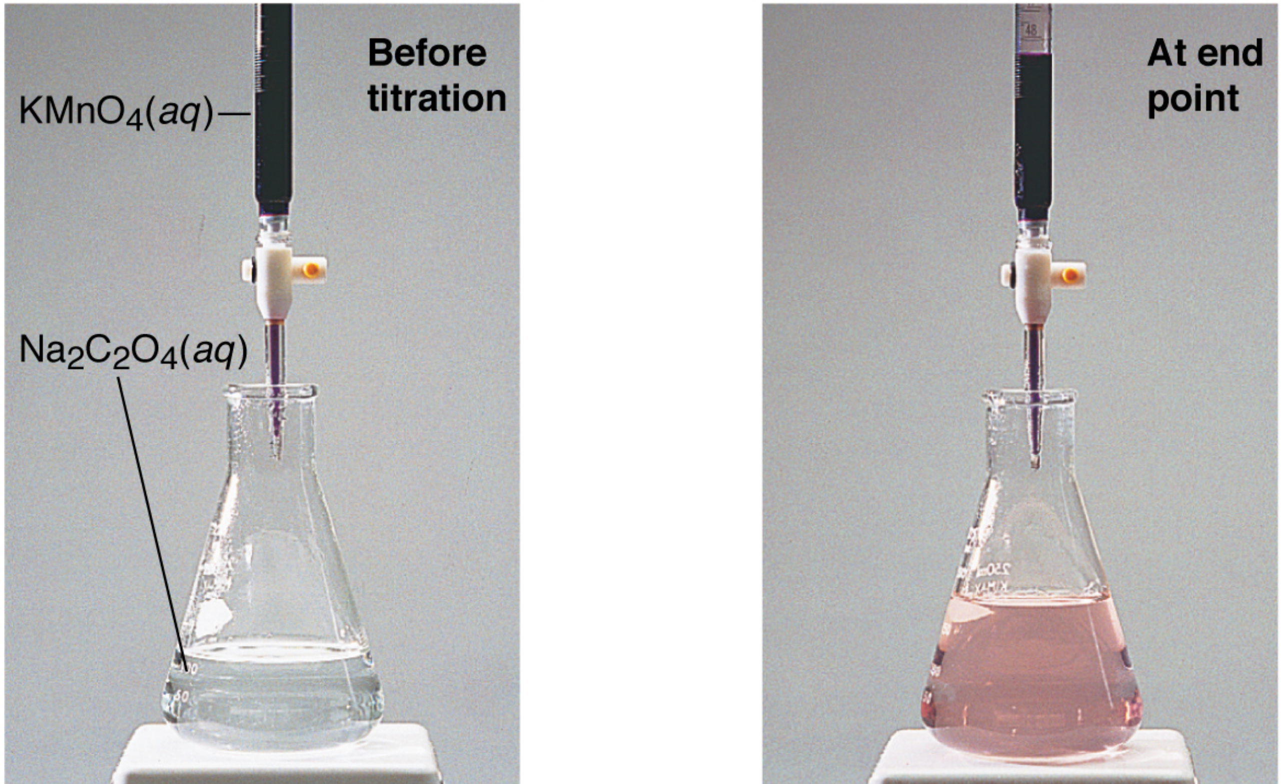
Multiply to make e^- lost = e^- gained:

$$PbS\ (s) + \frac{3}{2} O_2\ (g) \longrightarrow PbO\ (s) + SO_2\ (g)$$

Balance other atoms by inspection,
and multiply to give whole-number coefficients:

$$\boxed{2PbS\ (s) + 3O_2\ (g) \longrightarrow 2PbO\ (s) + 2SO_2\ (g)}$$

Figure 4.14 The redox titration of $C_2O_4^{2-}$ with MnO_4^-



Net ionic equation:

$$\overset{+7}{2MnO_4^-(aq)} + \overset{+3}{5C_2O_4^{2-}(aq)} + 16H^+(aq) \longrightarrow \overset{+2}{2Mn^{2+}(aq)} + \overset{+4}{10CO_2(g)} + 8H_2O(l)$$

Sample Problem 4.13 | **Finding the Amount of Reducing Agent by Titration**

PROBLEM: To measure the Ca^{2+} concentration in human blood, 1.00 mL of blood was treated with $Na_2C_2O_4$ solution to precipitate the Ca^{2+} as CaC_2O_4. The precipitate was filtered and dissolved in dilute H_2SO_4 to release $C_2O_4^{2-}$, which was titrated with $KMnO_4$ solution. The solution required 2.05mL of 4.88×10^{-4} *mol/L* $KMnO_4$ to reach the end point. The balanced equation is

$$2\,KMnO_4\,(aq) + 5\,CaC_2O_4\,(s) + 8\,H_2SO_4\,(aq) \longrightarrow$$
$$2\,MnSO_4\,(aq) + K_2SO_4\,(aq) + 5\,CaSO_4\,(s) + 10\,CO_2\,(g) + 8\,H_2O\,(l)$$

Calculate the amount (mol) of Ca^{2+} in 1.00 mL of blood.

Sample Problem 4.13

PLAN: Calculate the mol of $KMnO_4$ from the volume and concentration of the solution. Use this to calculate the mol of $C_2O_4^{2-}$ and hence the mol of Ca^{2+} ion in the blood sample.

volume of $KMnO_4$ soln

↓ **convert mL to L and multiply by *mol/L***

mol of $KMnO_4$

↓ **molar ratio**

mol of CaC_2O_4

↓ **ratio of elements in formula**

mol of Ca^{2+}

Sample Problem 4.13

SOLUTION:

$$2.05\ \cancel{\text{mL soln}} \times \frac{1\ \cancel{\text{L}}}{10^3\ \cancel{\text{mL}}} \times \frac{4.88\text{x}10^{-4}\ \text{mol KMnO}_4}{1\cancel{\text{L soln}}} = 1.00\text{x}10^{-6}\ \text{mol KMnO}_4$$

$$1.00\text{x}10^{-6}\ \cancel{\text{mol KMnO}_4} \times \frac{5\ \text{mol CaC}_2\text{O}_4}{2\ \cancel{\text{mol KMnO}_4}} = 2.50\text{x}10^{-6}\ \text{mol CaC}_2\text{O}_4$$

$$2.50\text{x}10^{-6}\ \cancel{\text{mol CaC}_2\text{O}_4} \times \frac{1\ \text{mol Ca}^{2+}}{1\ \text{mol}\ \cancel{\text{CaC}_2\text{O}_4}} \boxed{= \mathbf{2.50x10^{-6}\ mol\ Ca^{+2}}}$$

Elements in Redox Reactions
Types of Reaction

- Combination Reactions
  - Two or more reactantants combine to form a new compound:
  - X + Y → Z
- Decomposition Reactions
  - A single compound decomposes to form two or more products:
  - Z → X + Y
- Displacement Reactions
  - double diplacement: AB + CD → AC + BD
  - single displacement: X + YZ → XZ + Y
- Combustion
  - the process of combining with O_2

Figure 4.15 Combining elements to form an ionic compound.

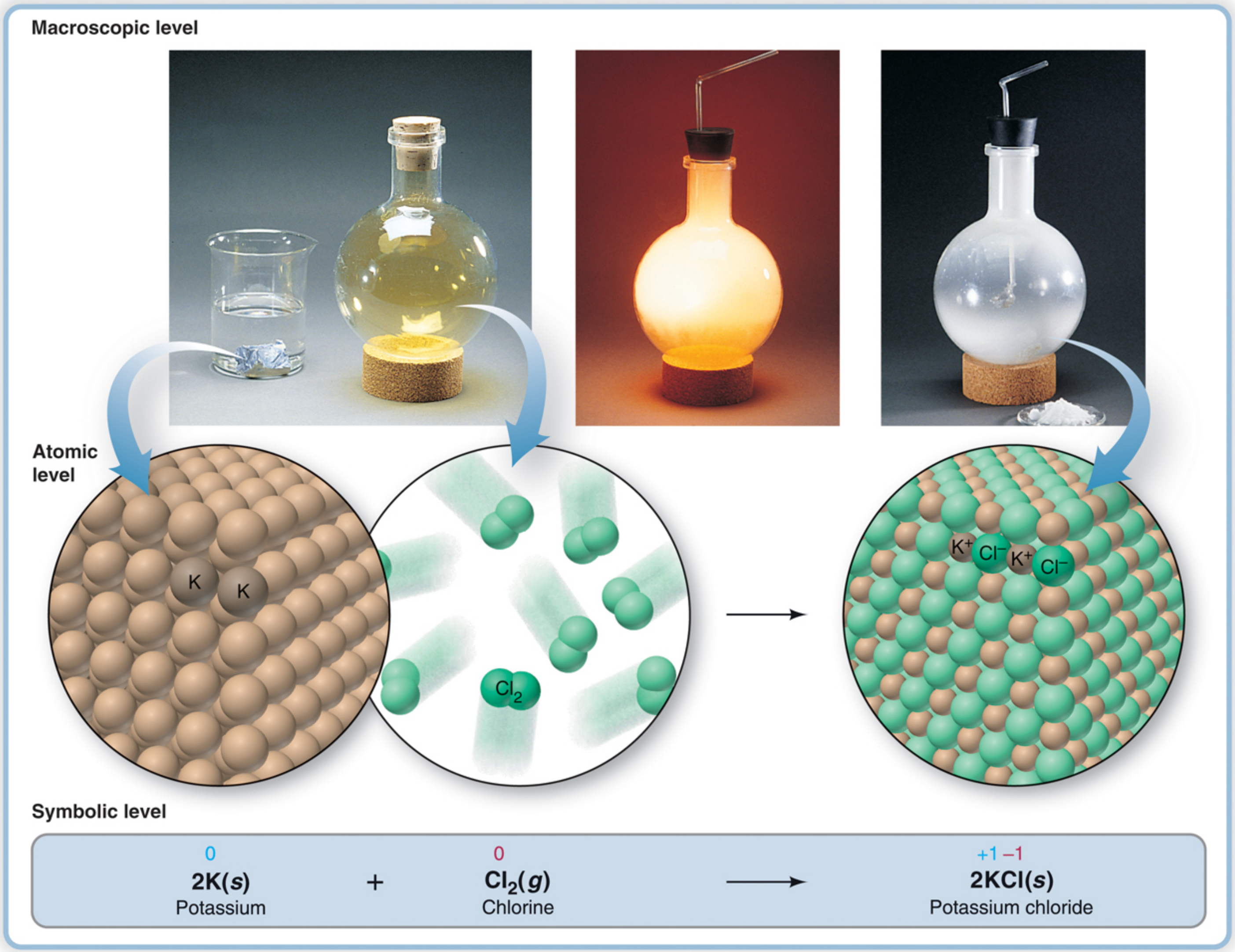


Figure 4.16 Decomposition of the compound mercury(II) oxide to its elements.

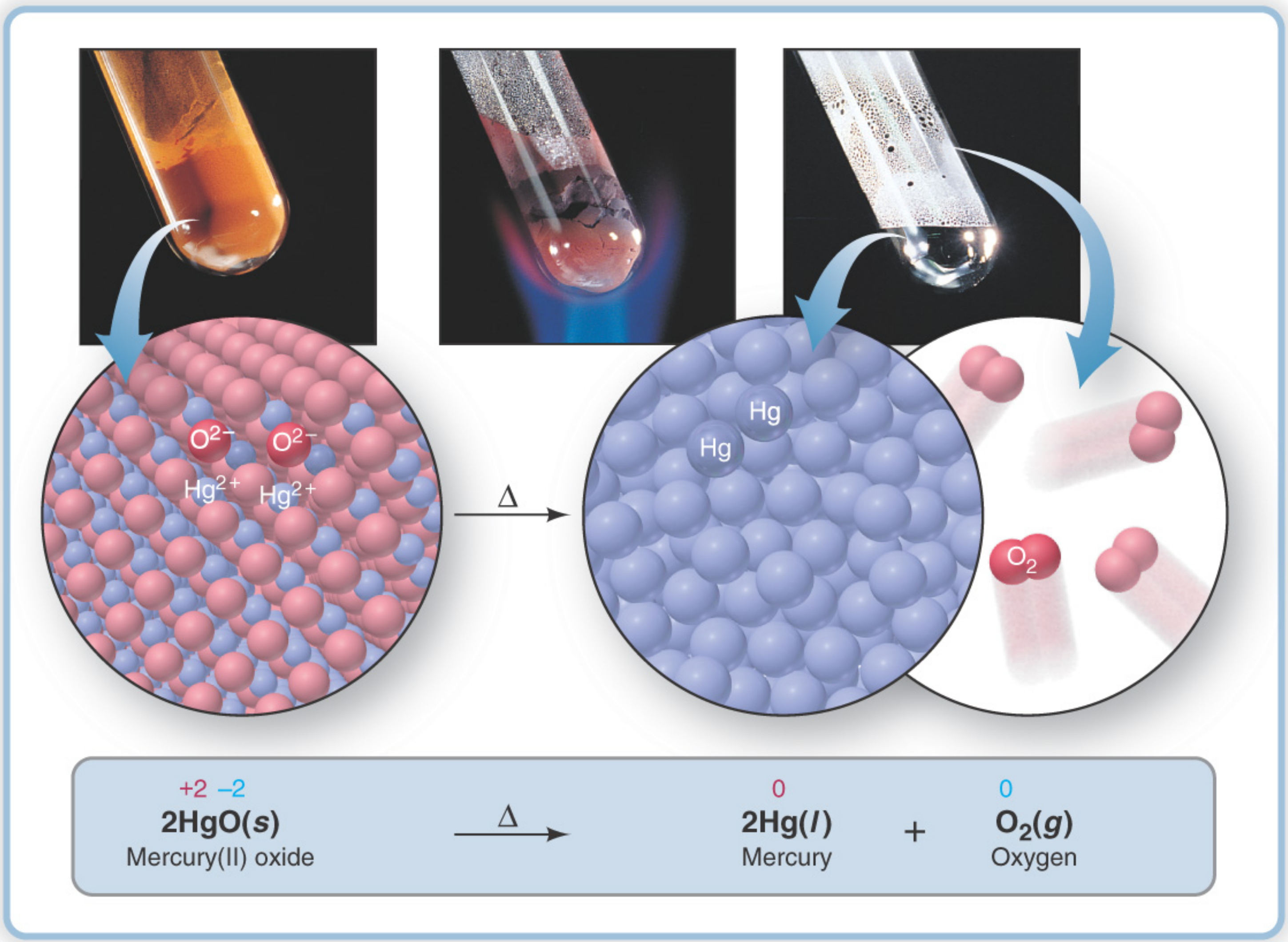
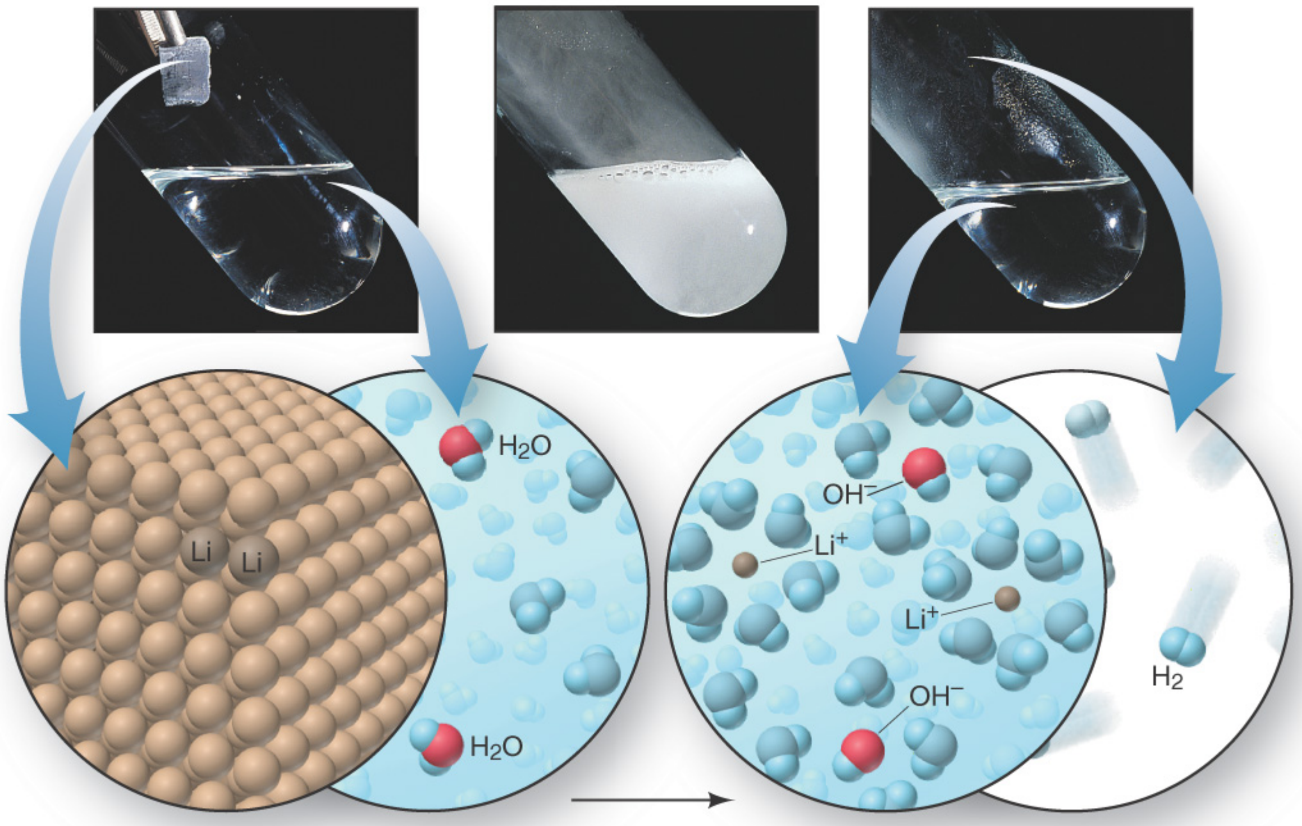


Figure 4.17 The active metal lithium displaces H_2 from water.



$$\overset{0}{2Li(s)} + \overset{+1\ -2}{2H_2O(l)} \longrightarrow \overset{+1\ -2\ +1}{2LiOH(aq)} + \overset{0}{H_2(g)}$$

Lithium + Water → Lithium hydroxide + Hydrogen

Figure 4.18 The displacement of H_2 from acid by nickel.

O.N. increasing

oxidation occurring

reducing agent

O.N. decreasing

reduction occurring

oxidizing agent

$$\overset{0}{Ni}\,(s) + 2\overset{\boxed{+1}}{H^+}\,(aq) \longrightarrow \overset{+2}{Ni^{2+}}\,(aq) + \overset{\boxed{0}}{H_2}\,(g)$$

Figure 4.19 A more reactive metal (Cu) displacing the ion of a less reactive metal (Ag^+) from solution.

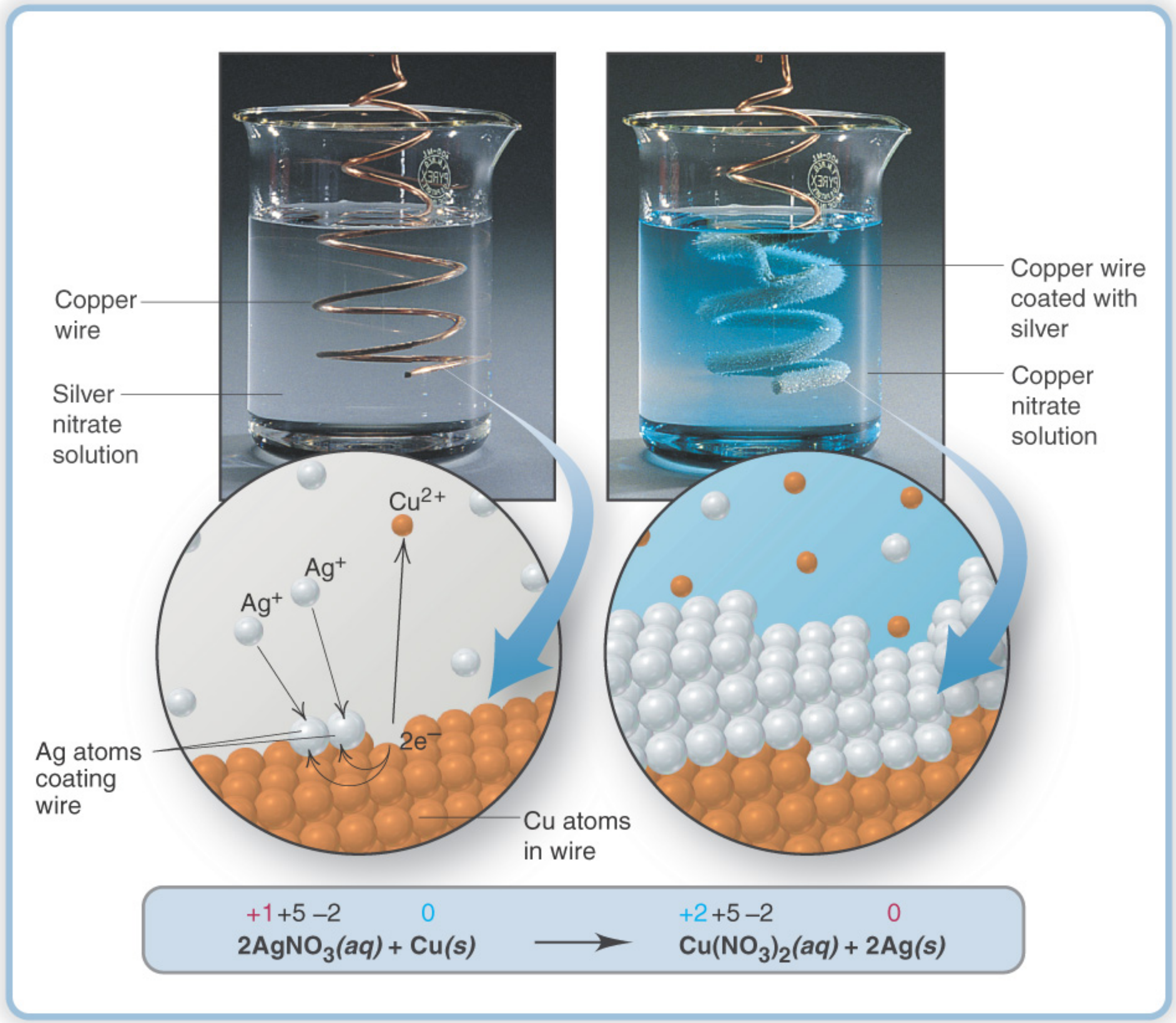
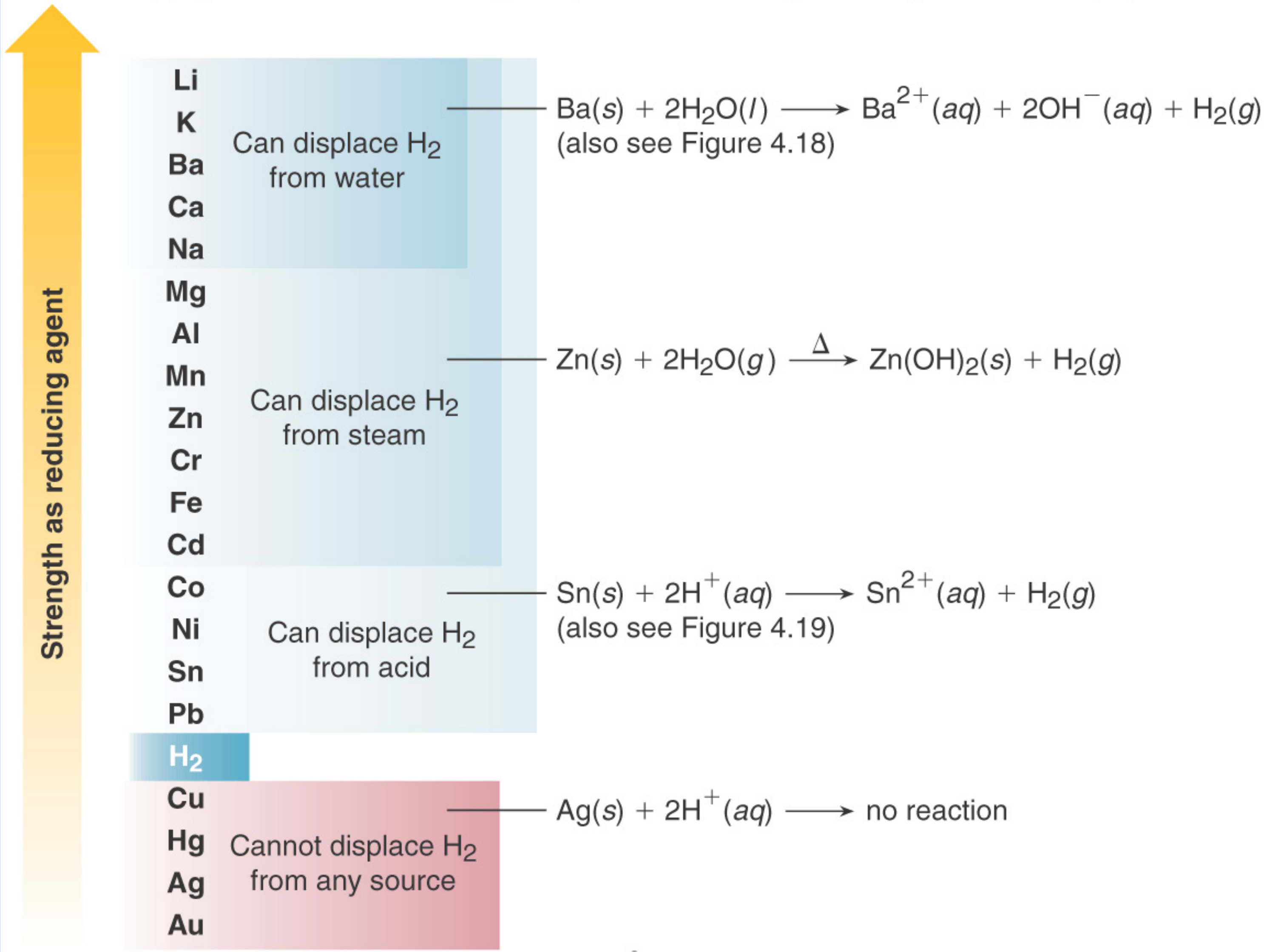


Figure 4.20 The activity series of the metals.



Sample Problem 4.14 **Identifying the Type of Redox Reaction**

PROBLEM: Classify each of the following redox reactions as a combination, decomposition, or displacement reaction. Write a balanced molecular equation for each, as well as total and net ionic equations for part (c), and identify the oxidizing and reducing agents:

(a) magnesium (*s*) + nitrogen (*g*) $\longrightarrow$ magnesium nitride (*aq*)

(b) hydrogen peroxide (*l*) $\longrightarrow$ water (*l*) + oxygen gas

(c) aluminum (*s*) + lead(II) nitrate (*aq*) $\longrightarrow$ aluminum nitrate (*aq*) + lead (*s*)

PLAN: Combination reactions combine reactants, decomposition reactions involve more products than reactants and displacement reactions have the same number of reactants and products.

Sample Problem 4.14

SOLUTION:

(a) This is a combination reaction, since Mg and N_2 combine:

$$\underset{0}{3Mg}\,(s) + \underset{0}{N_2}\,(g) \longrightarrow \underset{+2}{Mg_3}\underset{-3}{N_2}\,(s)$$

Mg is the reducing agent; N_2 is the oxidizing agent.

Sample Problem 4.14

(b) This is a decomposition reaction, since H_2O_2 breaks down:

$$2\,\underset{+1}{H_2}\underset{-2}{O_2}\,(l) \longrightarrow + 2\underset{+1}{H_2}\underset{-2}{O}\,(l) + \underset{0}{O_2}\,(g)$$

H_2O_2 is *both* the reducing and the oxidizing agent.

Sample Problem 4.14

(c) This is a displacement reaction, since Al displaces Pb^{2+} from solution.

$$2\underset{0}{\mathrm{Al}}\ (s) + 3\underset{+2}{\mathrm{Pb}}(\underset{+5}{\mathrm{N}}\underset{-2}{\mathrm{O}}_3)_2\ (aq) \longrightarrow 2\underset{+3}{\mathrm{Al}}(\underset{+5}{\mathrm{N}}\underset{-2}{\mathrm{O}}_3)_3\ (aq) + 3\underset{0}{\mathrm{Pb}}\ (s)$$

Al is the reducing agent; $Pb(NO_3)_2$ is the oxidizing agent.

The total ionic equation is:

$$2Al\ (s) + 3Pb^{2+}\ (aq) + 2NO_3^-\ (aq) \longrightarrow 2Al^{3+}\ (aq) + 3NO_3^-\ (aq) + 3Pb\ (s)$$

The net ionic equation is:

$$2Al\ (s) + 3Pb^{2+}\ (aq) \longrightarrow 2Al^{3+}\ (aq) + 3Pb\ (s)$$