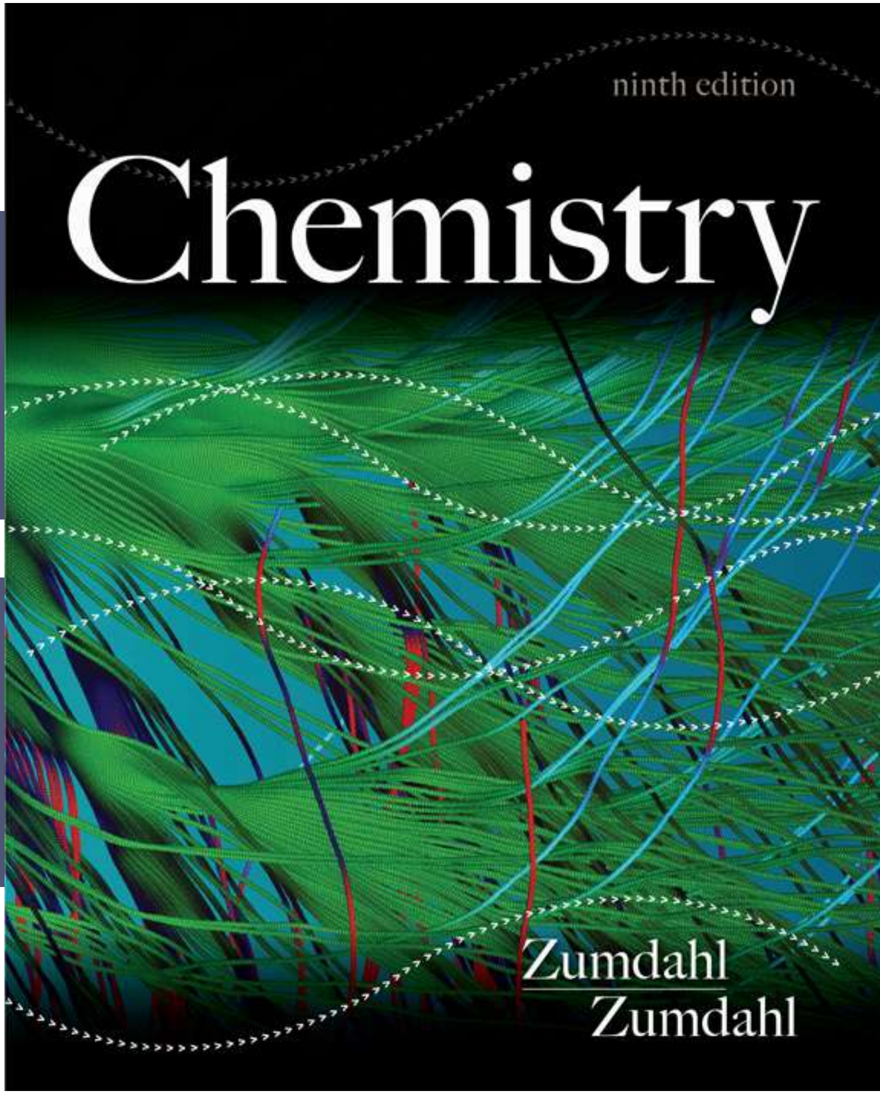


Chapter 4

Types of Chemical Reactions and Solution Stoichiometry

Water:

- One of the most important substances on Earth.
- Can dissolve many different substances.
- It is polar molecule
- H_2O

Aqueous Solutions

- Solute: Substance being dissolved.
- Solvent: Liquid water.
- Electrolyte: The substance that when dissolved in water produces a solution that conducts electricity.

Section 4.2

The Nature of Aqueous Solutions: Strong and Weak Electrolytes

Types of Materials

- Strong Electrolytes – Strong electrical conductors, completely ionized in water (100% dissociation). Ionic compounds are strong electrolytes like HCl, NaCl, $MgCl_2$, …
- NaCl Na+ + Cl-
- Weak Electrolytes – Weak electrical conductors, small degree of ionization in water. Weak salts, weak acids and bases are weak electrolytes. Acetic acid CH_3COOH is an example.
- Nonelectrolytes – no current flows. Dissolves but does not produce any ions. Like sugar.

Concentration, Molarity: molar concentration

Compare solutions (I) and (II):

- (I) 5 g in 50 mL solution & (II) 10 g in 50 mL solution
- (III) 10 g in 100 mL solution
- (II) is more concentrated than (I)
- (I) 5 g in 50 mL solution *and* (III) 10 g in 100 mL solution have the same concentration
- Concentration Units:
- Molarity (M) = moles of solute per volume of solution in liters

$$M = \text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}} \qquad 3\ M\ \text{HCl} = \frac{\text{6 moles of HCl}}{\text{2 liters of solution}}$$

Example:

A 375.0 g sample of potassium phosphate, K_3PO_4, is disolved in enough water to make 500 mL of solution. What is the molarity, M, of the solution? $MM(K_3PO_4)$= 212.3 g/mol.

M = n/V(L)

n = mass/MM

n (K_3PO_4) = 375.0g/(212.3g/mol.) = 1.766 mol.

V(L) = 500/1000 = 0.5 L

M = n/V(L) = 1.766 mol./0.5 L = 3.53 M

$K_3PO_4 \rightarrow 3K^+ + PO_4^{3-}$ (K_3PO_4 is strong electrolyte)

$M(K^+) = 3x3.53 = \boxed{10.59\ M}$

Concentration of Ions:

- For a 0.25 M $CaCl_2$ solution: ($CaCl_2$ is strong electrolyte)

$CaCl_{2\,(s)} \rightarrow Ca^{2+}{}_{(aq.)} + 2Cl^{-}{}_{(aq.)}$

1 mol. of $CaCl_2$ gives 1 mol. of Ca^{2+} ions and 2 moles of Cl^- ions. So:

- $[Ca^{2+}] = [CaCl_2] = 0.25\ M = 0.25$ M Ca^{2+}
- $[Cl^-] = 2[CaCl_2] = 2 \times 0.25\ M = 0.50$ M Cl^-.

Dilution

- The process of adding water to a concentrated or stock solution to achieve the molarity desired for a particular solution.
- Dilution with water does not alter the numbers of moles of solute present.
- Moles of solute before dilution = moles of solute after dilution

$$n = [\quad].V$$

$$M_i V_i = M_f V_f$$

- Example: Dilution

What is the volume of a 2.00 M NaOH solution needed to make 150.0 mL of a 0.800 M NaOH solution?

$$M_i V_i = M_f V_f$$

(discuss the units)

$$V_i = V_f \left(\frac{M_f}{M_i}\right)$$

= (150.0 mL)(0.800 M/2.00 M)

= 60.0 mL

(90 mL water to be added)

Exercise (H.W) : Dilution

How many mL of water are required to be added to 100.0 mL of 0.25 M NaOH solution in order to obtain a solution of 0.20 M?

Answer: 25.0 mL

- *Precipitation Reactions*:

In these reactions, one of the products is insoluble. So, it precipitates.

- *Acid–Base Reactions*: (OR *Neutralization Reactions*)

Reactions in which hydrogen (H^+) and hydroxide (OH^-) ions combine together to produce water.

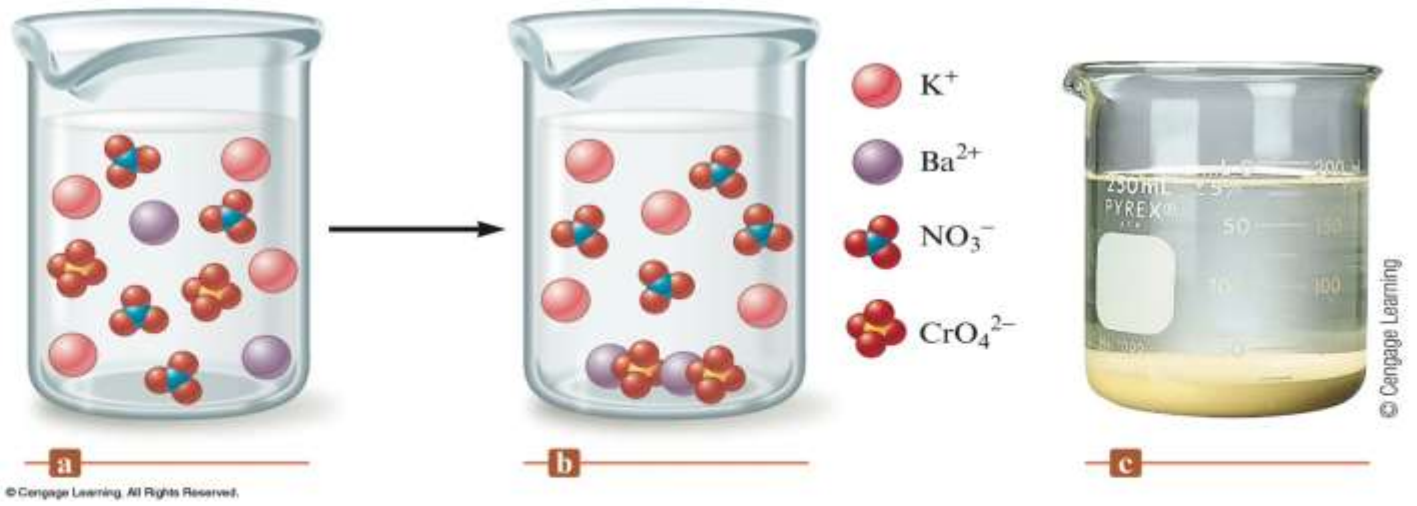
- *Oxidation–Reduction Reactions*: (*Redox Reactions*)

Reactions that involve electron transfer.

The Reaction of $K_2CrO_{4\,(aq.)}$ and $Ba(NO_3)_{2\,(aq.)}$

$$Ba(NO_3)_{2\,(aq.)} + K_2CrO_{4\,(aq.)} \rightarrow BaCrO_{4\,(s)} + 2KNO_{3\,(aq.)}$$

$$Ba^{2+}_{(aq.)} + CrO_4^{2-}{}_{(aq.)} \rightarrow BaCrO_{4\,(s)}$$



Simple Rules for Solubility

1. Nitrate (NO_3^-), alkali metal (group 1A) and NH_4^+ salts are soluble.
2. Most Cl^-, Br^-, and I^- salts are soluble (except Ag^+, Pb^{2+}, Hg_2^{2+}).
3. Most sulfate salts are soluble (except $BaSO_4$, $PbSO_4$, Hg_2SO_4, $CaSO_4$, BaCrO4).
4. Most OH^-, S^{2-}, CO_3^{2-}, CrO_4^{2-}, PO_4^{3-} salts are insoluble, except those containing the cations in Rule 1 above.

Write the molecular, ionic equation and net ionic equations for the reaction between cobalt(II) chloride and sodium hydroxide.

Molecular Equation: (Formula Equation)

$$CoCl_{2\,(aq.)} + 2NaOH_{(aq.)} \rightarrow Co(OH)_{2\,(s)} + 2NaCl_{(aq.)}$$

Ionic Equation:

$$Co^{2+}{}_{(aq.)} + 2Cl^{-}{}_{(aq.)} + 2Na^{+}{}_{(aq.)} + 2OH^{-}{}_{(aq.)} \rightarrow Co(OH)_{2(s)} + 2Na^{+}{}_{(aq.)} + 2Cl^{-}{}_{(aq.)}$$

Net Ionic Equation:

$$Co^{2+}{}_{(aq.)} + 2OH^{-}{}_{(aq.)} \rightarrow Co(OH)_{2(s)}$$

spectator ions

Solving Stoichiometry Problems for Reactions in Solution:

$$n = (M)(V) \quad ; \quad n = \text{mass}/MM$$

- How many <u>grams</u> of $BaSO_4$ are precipitated when 100.00 mL of 0.25 M solution of $BaCl_2.2H_2O$ is mixed with 150.00 mL 0.200 M solution of Na_2SO_4? *MM(g/mol.): $BaSO_4$ = 233.33*

$$BaCl_2.2H_2O_{(aq.)} + Na_2SO_{4(aq.)} \rightarrow BaSO_{4(s)} + NaCl_{(aq.)} + 2H_2O$$

<u>Solution</u>:

n = mass/MM ; M(mol./L) = n/V(L) ; n = M.V(L)

Our case is mixing of reactants, so, we musn’t forget the possibility of having none-stoichmetric mixing, which, of course, implies a limiting reactant.

Stoichiometry: We need Balanced chemical equation.

Molecular equation:

$$BaCl_2.2H_2O_{(aq)} + Na_2SO_{4\,(aq)} \rightarrow BaSO_{4\,(s)}\downarrow + 2NaCl_{(aq)} + 2H_2O_{(aq)}$$

0.0250 (L.R) 0.0300

n = M.V(L)

n ($BaCl_2.2H_2O$) = (0.100 L) (0.250 M) = 0.0250 mol.

n (Na_2SO_4) = (0.150 L) (0.200 M) = 0.0300 mol.

The stoichiometric ratio of the reactants is 1:1 so the L.R is $BaCl_2.2H_2O_{(aq)}$ because it has the lowest number of moles.

The stoichimetric ratio between the L.R and the precipitate ($BaSO_4$) is 1:1, so,

n ($BaSO_4$) = n ($BaCl_2.2H_2O$) = 0.0250 mol.

Mass ($BaSO_4$) = (n)(MM) = 0.0250 mol.x 233.33 = 5.83 g.

Homework:

10.0 mL of a 0.30 M sodium phosphate, Na_3PO_4, solution mixed with 20.0 mL of a 0.20 M lead(II) nitrate, $Pb(NO_3)_2$, solution. How many grams of $Pb_3(PO_4)_2$ precipitated?

MM(g/mol.): $Pb_3(PO_4)_2 = 811.54$.

Answer: 1.1 g $Pb_3(PO_4)_2$

The molecular equation:

$$2Na_3PO_{4\,(aq.)} + 3Pb(NO_3)_{2\,(aq.)} \rightarrow Pb_3(PO_4)_{2\,(s)} + 6NaNO_{3\,(aq.)}$$

1. Find the number of moles of the reactants?
2. Find the limiting reactant?
3. Find the number of moles of the precipitate, $Pb_3(PO_4)_2$?
4. Find the mass of the precipitate?

Acid–Base Reactions

Brønsted–Lowry Definitions of acids and bases:

- Acid: proton donor
- Base: proton acceptor

I. Strong acid/strong base reactions:

Example is the reaction between HCl and NaOH,

Molecular eqn.: $HCl_{(aq.)} + NaOH_{(aq.)} \rightarrow H_2O_{(l)} + NaCl_{(aq.)}$

Ionic eqn.:

$$H^+_{(aq.)} + Cl^-_{(aq.)} + Na^+_{(aq.)} + OH^-_{(aq.)} \rightarrow H_2O_{(l)} + Cl^-_{(aq.)} + Na^+_{(aq.)}$$

Spectator Ions: $Cl^-_{(aq.)}$ and $Na^+_{(aq.)}$

<u>Net Ionic Eqn.</u>: *(remove the spectator ions)*

$$H^+_{(aq.)} + \cancel{Cl^-}_{(aq.)} + \cancel{Na^+}_{(aq.)} + OH^-_{(aq.)} \rightarrow H_2O_{(l)} + \cancel{Cl^-}_{(aq.)} + \cancel{Na^+}_{(aq.)}$$

$$H^+_{(aq.)} + OH^-_{(aq.)} \rightarrow H_2O_{(l)}$$

Consider the reaction of HNO_3 and KOH. Strong acid and strong base respectively. Here, K^+ and NO_3^- are <u>spectator ions</u>. Consequently, the net ionic equation is the same as
the one obtained for the reaction of HCl and NaOH. So,

For a strong acid/strong base reaction the net ionic equation is always:

$$H^+_{(aq.)} + OH^-_{(aq.)} \rightarrow H_2O_{(l)}$$

II. Weak acid/strong base or weak base/strong acid reactions:

Example: reaction of acetic acid and NaOH

$$CH_3COOH_{(aq.)} \;+\; NaOH_{(aq.)} \rightarrow H_2O_{(l)} + CH_3COO^-Na^+_{(aq.)}$$

weak acid — strong base

weak dissociation

Na^+ is the only spectator ion and the net ionic eqn. is:

$$CH_3COOH_{(aq.)} \;+\; OH^-_{(aq.)} \rightarrow H_2O_{(l)} + CH_3COO^-_{(aq.)}$$

<u>Example</u>:

Reaction of 25.00 mL of unknown HCl solution required 26.60 mL of 0.100 M NaOH solution to consume all HCl. What is the concentration of the HCl solution?

$$HCl_{(aq.)} + NaOH_{(aq.)} \rightarrow H_2O_{(l)} + NaCl_{(aq.)}$$

The stoichiometric ratio of HCl : NaOH is 1:1, so:

n (HCl) = n (NaOH)

(M) (V) = (M) (V)

(M) (25.00 mL) = (0.100)(26.60 mL)

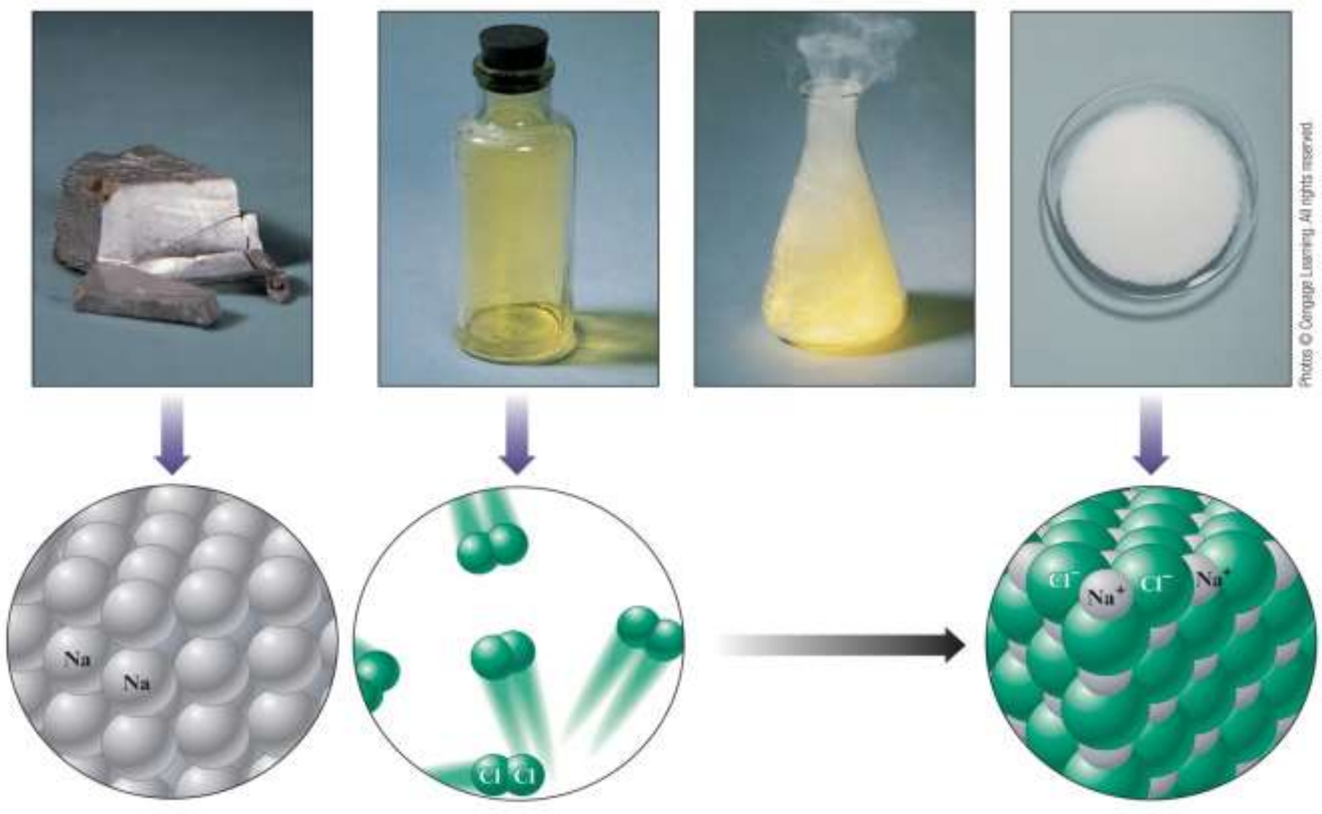
☐ M (HCl) = (0.100)(26.60 mL)/(25.00 mL) = 0.106 M

$$H_2SO_{4\,(aq.)} + 2NaOH_{(aq.)} \rightarrow 2H_2O_{(l)} + Na_2SO_{4\,(aq.)}$$

Oxidation-Reduction Reactions: (Redox Reactions)

They are reactions that involve one or more electron transfer.

Example: $2Na_{(s)} + Cl_{2\,(g)} \rightarrow 2Na^{+}Cl^{-}_{(s)}$



Rules for Assigning Oxidation States

1. Oxidation state of an atom in an element = 0
2. Oxidation state of monatomic ion = charge of the ion
3. Oxygen = –2 in covalent compounds (except in peroxides where it = –1) H_2O_2
4. Hydrogen = +1 in covalent compounds (except in hydrides, -1)
5. Fluorine = –1 in compounds
6. Sum of oxidation states = 0 in neutral compounds
7. Sum of oxidation states = charge of the ion in ions
8. LiH, NaH, H2CO3, MnO4-

Find the oxidation states for each of the elements in each of the following compounds:

(go back to chapter 2 and look at the charges of polyatomic ions)

- $K_2Cr_2O_7$ K = +1; Cr = +6; O = –2
- CO_3^{2-} C = +4; O = –2
- MnO_2 Mn = +4; O = –2
- PCl_5 P = +5; Cl = –1
- SF_4 S = +4; F = –1

Redox Reactions’ Characteristics

- Transfer of electrons
- Ions are formed upon the electron transfer.
- Oxidation – increase in oxidation state, it happens by loss of electrons). Reducing agent
- Reduction – decrease in oxidation state it happens by gain of electrons); oxidizing agent

Oxidizing agent and reducing agent:

Example: Reaction of sodium and chlorine

$2Na_{(s)} + Cl_{2\,(g)} \rightarrow 2NaCl_{\,(aq.)}$ Oxidation/Reduction Reaction

$2Na_{(s)} \rightarrow Na^{+}_{(aq.)} + 2e^{-}$ Oxidation half

$Cl_{2\,(g)} + 2e^{-} \rightarrow 2Cl^{-}_{(aq.)}$ Reduction half

- $Na_{(s)}$: Reducing agent, itself oxidized, it reduced chlorine.
- $Cl_{2\,(g)}$: Oxidizing agent, itself reduced, it oxidized sodium.

Example:

Is the following an oxidation-reduction reaction? If yes, identify the oxidizing agent and the reducing agent.

$$Zn_{(s)} + 2HCl_{(aq.)} \rightarrow ZnCl_{2\,(aq.)} + H_{2\,(g)}$$

$Zn_{(s)} \rightarrow Zn^{+2}_{(aq.)} + 2e^{-}$ Oxidation half

$2H^{+}_{(aq.)} + 2e^{-} \rightarrow H_{2(g)}$ Reduction half

Zn is reducing agent and HCl is oxidizing agent.

- Balance the reaction between solid zinc and aqueous hydrochloric acid to produce aqueous zinc(II) chloride and hydrogen gas.

$$Zn_{(s)} + HCl_{(aq.)} \rightarrow ZnCl_{2\,(aq.)} + H_{2\,(g)}$$

- Find the oxidation state for the reactants and products.
- Predict the electron loss and gain, then equalize them.

Electron loss and gain?
What coefficients are needed to equalize the electrons gained and lost?

1 e^- gained (each atom) **× 2**

- $Zn_{(s)} + HCl_{(aq.)} \rightarrow Zn^{2+}_{(aq.)} + Cl^-_{(aq.)} + H_{2(g)}$

  0 +1 −1 +2 −1 0

  2 e^- lost

- The oxidation state of chlorine remains unchanged.

- $Zn_{(s)} + 2HCl_{(aq.)} \rightarrow Zn^{2+}_{(aq.)} + Cl^-_{(aq.)} + H_{2(g)}$

Mass and Charge Balances:

- $Ce^{+4}{}_{(aq.)} + Sn^{+2}{}_{(aq.)} \rightarrow Ce^{+3}{}_{(aq.)} + Sn^{+4}{}_{(aq.)}$
- $2 \times \{ Ce^{+4}{}_{(aq.)} + e^{-} \rightarrow Ce^{+3}{}_{(aq.)} \}$ reduction
- $Sn^{+2}{}_{(aq.)} \rightarrow Sn^{+4}{}_{(aq.)} + 2e^{-}$ oxidation

$$\boxed{2\,Ce^{+4}{}_{(aq.)} + Sn^{+2}{}_{(aq.)} \rightarrow 2\,Ce^{+3}{}_{(aq.)} + Sn^{+4}{}_{(aq.)}}$$

End of Chapter 4