GENERAL CHEMISTRY REVIEW

Review Topics Include:

- Naming Simple Compounds
- Rules for Reporting Significant Figures
- Units of Concentration
- Review of the Equilibrium Expression
- Review of Simple Acid/Base Properties
- Rules for Predicting pH of Solution
- Solubility
- Rules for Assigning Oxidation States
- Review of Thermodynamics
- Review of Electrochemistry
Naming Simple Compounds

**Binary Ionic Compounds**, where the metal ion has only one oxidation state
(Group 1A, alkali metals and group 2A, alkali earth metals)
1. the cation (positive ion) named first using the element name
2. monatomic cations take name from the corresponding element (i.e., cesium)
3. monatomic anions (negative ions) named from the element root and adding -ide suffix
   (i.e., bromide)
For example, CsBr is named cesium bromide.

**Binary Ionic Compounds**, where the metal ion has variable oxidation state (Transition elements)
1. the oxidation state on the metal ion is specified by Roman Numeral in brackets
2. monoatomic anions are named as before
For example, CuCl and CuCl₂ are named as copper (I) chloride and copper (II) chloride, respectively. The older method of using the suffix -ous and -ic to indicate low and high oxidation states respectively is still encountered. Thus, the above compounds could appear as cuprous chloride and cupric chloride.

**Binary Compounds involving only non-metals** (Group 4A - 8A)
1. first element in the formula is named using element name (i.e., nitrogen)
2. second element named as if it were an anion (i.e., oxide)
3. prefixes (mono, di, tri, tetra, penta, hexa etc.) used to denote number of atoms
4. prefix mono- never used for naming the first element (i.e., nitrogen dioxide as opposed to mononitrogen dioxide)
For example, NO₂ is named nitrogen dioxide and P₂O₇ is named diphosphorous heptaoxide.

For compounds involving **polyatomic ions**, similar rules apply. The names of the following common polyatomic ions must be committed to memory.

- \( \text{CO}_3^{2-} \) (carbonate)
- \( \text{HCO}_3^- \) (hydrogen carbonate; aka bicarbonate)
- \( \text{NH}_4^+ \) (ammonium)
- \( \text{NO}_2^- \) (nitrite)
- \( \text{NO}_3^- \) (nitrate)
- \( \text{S}^{2-} \) (sulfide)
- \( \text{SO}_3^{2-} \) (sulfite)
- \( \text{SO}_4^{2-} \) (sulfate)
- \( \text{HSO}_4^- \) (hydrogen sulfate)
- \( \text{PO}_4^{3-} \) (phosphate)
- \( \text{HPO}_4^{2-} \) (hydrogen phosphate)
- \( \text{H}_2\text{PO}_4^- \) (dihydrogen phosphate)
- \( \text{CN}^- \) (cyanide)
- \( \text{ClO}^- \) (hypochlorite)
- \( \text{CH}_3\text{CO}_2^- \) (ethanoate; aka acetate)
- \( \text{MnO}_4^- \) (permanganate)
- \( \text{CrO}_4^{2-} \) (chromate)

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**Sample Exercises:**
1. Name each of the following.
   a) \( \text{Ca}_3(\text{PO}_4)_2 \)
   b) \( \text{Cr}_2\text{O}_3 \)
   c) \( \text{ClO}_2 \)

2. Give the molecular formula for each of the following.
   a) sodium hypochlorite
   b) mercury (II) sulfate
**Rules for Reporting Significant Figures**

1. Nonzero digits always count as significant figures.

2. Zeros are what mix people up. There are three situations in which they can occur.

   * **leading zeros** precede all nonzero digits and are **never** significant (i.e., 0.000182 has three sign. figs.)
   
   * **captive zeros** are between nonzero digits and are **always** significant (i.e., 1008.02 has six sign. figs.)
   
   * **trailing zeros** are significant **only if** a number contains a decimal point (i.e., 1200 has two sign. figs.; 1200.00 has six sign. figs.; 1.200x10^3 has four sign. figs.)

   * **Note here the advantage of using exponential notation to clear up these ambiguities!**

3. Exact numbers have no bearing on the number of significant figures in a calculated result. Examples of these are the following:

   - conversion factors such as 1 L = 1000 mL
   - numbers reflecting an exact count such as 8 stones or 16 people
   - stoichiometry in chemical reactions involves exact numbers

**Operations**

**Rules for multiplication/division**

The answer to contain the same **number of sign. figs.** as the least precise measurement used in the calculation.

\[
\begin{array}{ccc}
72.5674 & \text{six sign. figs} \\
\times & 3.34 & \text{three sign. figs (limiting term)} \\
242.3751160 & \text{initial answer (must be rounded off to three sign. figs.)} \\
\end{array}
\]

Final Answer = \(2.42 \times 10^2\)

**Rules for addition/subtraction**

The answer to contain the same **number of decimal places** as the least precise measurement used in the calculation.

\[
\begin{array}{ccc}
456.367963 \\
- 452.1 & \text{least number of decimal places (limiting term)} \\
4.267963 & \text{initial answer (must be rounded off to one decimal place)} \\
\end{array}
\]

Final Answer = 4.3 !

**Rules for logarithms**

In logarithmic values, only those numbers to the right of the decimal place count as significant. For example,

\[
\begin{array}{ccc}
\text{pH} = 10.26 & \text{has only two significant figures and corresponds to } [\text{H}^+] = 5.5 \times 10^{-11} \text{ M} \\
pK_w = 4.730 & \text{has three significant figures and corresponds to } \text{K}_w = 1.86 \times 10^{-5} \\
\end{array}
\]

**Sample Exercises:**

3. What is the pH if the concentration of H\(^+\) is measured to be \(1.25 \times 10^{-6} \text{ M}\)?

4. What is the pK\(sp\) if the solubility product has been determined to be \(2 \times 10^{-30}\)?
UNITS OF CONCENTRATION
Several different units are used to express the concentration of species in solution. You are familiar with the concept of molarity (M) and its usefulness in chemical calculations since it reflects actual numbers of particles in a system. (Recall: moles of solute = mass of solute/MW solute)

\[ M \text{olarity} = \frac{\text{moles of solute}}{\text{litres of solution}} \]

Two common units that you have not used as frequently are ppm (parts per million), and N (normality).

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PART PER MILLION (mass ratio analogous to wt%)
The unit of ppm is the number of milligrams of solute per kg of solvent. For example, a solution that has 1.0 mg of Cl\(^-\) in 1.0 kg of water contains 1.0 mg of Cl\(^-\) in 1000 g of water. As there are 1000 mg in 1 g the concentration is

\[ \frac{1.0 \text{ mg Cl}^-}{1 \text{ kg solvent}} = \frac{1.0 \text{ mg Cl}^-}{1000 \text{ g solvent}} = \frac{1.0 \text{ mg Cl}^-}{1000000 \text{ mg solvent}} \]

i.e. from this last expression there is 1 mg of Cl\(^-\) in 1 million mg of water, hence the term parts per million or ppm. This solution has a Cl\(^-\) concentration of 1.0 ppm.

Assuming the density of water to be 1.0 g/mL the above solution contains 1.0 mg of Cl\(^-\) in 1000 mL (1 L) of water so 1 ppm = 1 mg/L (or 1 μg/mL). The above assumption about the density of water becomes less valid at high temperatures or in highly concentrated solutions such as effluents or in seawater where the density is greater than 1.0 g/mL (density seawater = 1.04 g/mL). Thus, 1 ppm ≠ 1 mg/L for seawater!

Parts per billion: 1 ppb = 1000 x ppm = 1 μg/L (or ng/mL)
The above solution has a Cl\(^-\) concentration of 1 ppm or 1000 ppb

Confusion can be encountered when reporting the concentration of elements that are present in polyatomic species, such as the nitrogen in nitrate form or the phosphorous in phosphates. for example, 1 mg/L of N (nitrogen) corresponds to 4.4 mg/L of NO\(_3\)^- (nitrate) since a nitrate has 4.4 times the mass of a nitrogen (62 g/mol versus 14 g/mol). To clarify what species is being referred to, you may see units reported as mg/L N-NO\(_3\)^- or mg/L NO\(_3\)^- - N. This indicates that it is the mass (in mg) of nitrogen (present in the chemical form of nitrate) per liter of water.

For air samples, ppm means parts per million by volume rather than by mass and this is often denoted by ppmv. The concentration of CO\(_2\) in air is about 325 ppm or 325 mL of CO\(_2\) in 1 million mL (1000 L) of air.

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Sample Exercises:
5. Convert the following to freshwater concentrations.
   a) 3.2 mmol/L of P to ppm of PO\(_4\)^3-
   b) 1.9 x 10\(^{-6}\) M of copper to ppb Cu

6. Convert the following concentrations to molarity.
   a) 9.8 mg/L of O\(_2\)
   b) 150 ppb Arsenic in seawater
NORMALITY

Normality is defined as the number of equivalents per liter. An equivalent is a multiple of the number of moles of a substance and depends on the context of the reaction (e.g., equivalents of acid, equivalents of electrons or equivalents of charge). Normality units have been retained in water analysis and in industrial analysis because of their advantage is in carrying out titration calculations. Its disadvantage is that it’s definition depends on the type of reaction involved. There are three broad classifications (i) acid/base reactions, (ii) oxidation/reduction reactions and (iii) precipitation or complexation reactions.

\[
\text{Normality}(N) = \frac{\text{equivalents of solute}}{\text{litres of solution}}
\]

The equivalent referred to in this equation is defined as follows:

\[
\text{number of equivalents} = \frac{\text{weight of solute (grams)}}{\text{equivalent weight of solute (grams/equivalent)}}
\]

The tricky part is determining the equivalent weight of the solute.

The equivalent weight, \( EW = \frac{\text{Molecular Weight, } MW}{\text{constant, } K} \) (or = \( \frac{\text{Molar Mass}}{\text{constant, } K} \))

\( K \) is an integer constant \((K \geq 1)\), and is equal to the number of moles of active species (protons, electrons etc) per mole of substance. For example, for \( \text{H}_2\text{SO}_4 \) each mole of sulfuric acid yields two moles of protons, so \( K = 2 \) and the \( EW = \frac{1}{2} \text{MW} \).

In general: \( \text{Normality} = K \times \text{Molarity} \)

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Sample Exercises:

7. A solution contains the following ions. Calculate the number of milliequivalents of positive and negative charge, respectively.
   - 80 mg/L \( \text{Ca}^{2+} \)
   - 22 mg/L \( \text{Na}^+ \)
   - 12 mg/L \( \text{Mg}^{2+} \)
   - 1.2 mg/L \( \text{Al}^{3+} \)
   - 100 mg/L \( \text{HCO}_3^- \)
   - 25 mg/L \( \text{Cl}^- \)
   - 12 mg/L \( \text{SO}_4^{2-} \)

8. A 1.00 L sample reacted with 22.62 mL of 0.1034 M \( \text{H}_2\text{SO}_4 \). How many equivalents of acid were neutralized?
Review of the **Equilibrium Expression**, \( K_{eq} \)

The equilibrium expression for the generalized reaction is determined experimentally by measuring concentrations for solutions (\( K_c \)) or partial pressures for gases (\( K_p \)). It can also be calculated from thermodynamic information, such as \( \Delta G^\circ \) or \( E^\circ \).

\[ a \text{ A} + b \text{ B} \rightleftharpoons c \text{ C} + d \text{ D} \]

\[ K_{eq} = \frac{[C]^c [D]^d}{[A]^a [B]^b} \]

Experimental equilibrium constants are obtained from measured concentrations or pressures and has units of concentration raised to the appropriate power. However, the Equilibrium Expression is actually defined thermodynamically as a quotient of chemical activities rather than concentrations. Activities are numerically similar to concentrations (particularly at low concentrations) but are unitless expressions, thus equilibrium constants often appear without units.

The numerical value of the Equilibrium constant (\( K_{eq} \));
- is constant at a given temperature
- is independent of initial concentrations
- depends on the form of the balanced equation

The position of heterogeneous equilibria does not depend on the concentrations of solids and pure liquids (i.e., the chemical activity of solids and pure liquids = 1). Therefore, these concentration terms do not appear in the equilibrium expression. For example:

\[ \text{CaCO}_3(\text{s}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g}) \rightleftharpoons \text{Ca}^{2+}(\text{aq}) + 2 \text{HCO}_3^-(\text{aq}) \]

\[ K_{eq} = [\text{Ca}^{2+}][\text{HCO}_3^-]^2/ P_{\text{CO}_2} \]

**Combining chemical reactions and equilibrium constants:**

Equilibrium constants for unknown reactions can be derived from the combination known chemical equilibria.

The equilibrium constant for a reaction that is equivalent to the sum of several other reactions is the product of the equilibrium constants of the constituent reactions.

\[ \text{A} + \text{B} \rightleftharpoons \text{C} + \text{D} \quad K_1 \]
\[ \text{A} + \text{C} \rightleftharpoons \text{E} \quad K_2 \]

\[ 2 \text{A} + \text{B} \rightleftharpoons \text{D} + \text{E} \quad K_{eq} = K_1 K_2 \]

If a reaction is reversed, the equilibrium constant is the inverse of the original constant (1/K)

\[ \text{NH}_3 + \text{H}_3\text{O}^+ \rightleftharpoons \text{NH}_4^+ + \text{H}_2\text{O} \quad K_{eq} = 1/K_{d(\text{NH}_4^+)} \]
**Predicting the direction of a reaction using \( Q \), the reaction quotient:**

The reaction quotient, \( Q \) has the same form as the equilibrium expression, but involves any set of reactant and product concentrations, not necessarily those at equilibrium. The value of \( Q \) can be compared to that of \( K \) to predict which way an equilibrium will shift given a set of initial concentrations. For instance,

- if \( Q > K \) the reaction will shift to the left
- if \( Q < K \) the reaction will shift to the right
- if \( Q = K \) the reaction is at equilibrium

**Calculating equilibrium concentrations/pressures given initial conditions:**

Using the balanced chemical equation and the initial concentrations, set up an expression for the equilibrium concentration of reactants and products in terms of \( x \), the extent of reaction. Substitute these equilibrium concentrations into the expression for the equilibrium constant and solve for \( x \). Once \( x \) is known it can be substituted into the earlier equations to give the equilibrium concentrations. These types of equilibrium problems are often complicated by cumbersome polynomial equations. Here are several short cuts:

- To solve a quadratic equation of the form \( ax^2 + bx + c = 0 \), use the quadratic formula.

  - Under certain circumstances simplifications are possible which greatly reduce the mathematical difficulty of the problem. Such a situation arises when treating a system with a very small equilibrium constant (see examples). After solving for \( x \), check to see if the approximation was a valid one!

**Samples Exercises:**

9. Write an expression for the equilibrium constant for the dissolution of calcium carbonate in the presence of carbon dioxide gas (see rxn above) from the following information.

   \[
   \text{CaCO}_3(s) \rightleftharpoons \text{Ca}^{2+}(aq) + \text{CO}_3^{2-}(aq) \quad K_{sp}
   \]

   \[
   \text{HCO}_3^-(aq) + \text{H}_2\text{O} \rightleftharpoons \text{CO}_3^{2-}(aq) + \text{H}_3\text{O}^+ \quad K_{a2(\text{H}_2\text{CO}_3)}
   \]

   \[
   \text{CO}_2(g) \rightleftharpoons \text{CO}_2(aq) \quad K_{\text{H}}
   \]

   \[
   \text{CO}_2(aq) + \text{H}_2\text{O} \rightleftharpoons \text{HCO}_3^-(aq) + \text{H}_3\text{O}^+(aq) \quad K_{a1(\text{H}_2\text{CO}_3)}
   \]

10. Consider the following reaction between the powerful disinfectant hypochlorous acid (HOCl) and water to form a weaker disinfectant hypochlorite ion (OCl\(^-\)).

   \[
   \text{HOCl} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OCl}^- \quad K_{eq} = 2.5 \times 10^{-8}
   \]

   a) Compute the equilibrium ratio of HOCl to OCl\(^-\) and the percent of the total ‘OCl’ present as HOCl at pH 8.0.

   b) At a particular instant, the activities (concentrations) of H\(_3\)O\(^+\), HOCl and OCl\(^-\) in solution are 10\(^{-7}\), 10\(^{-3}\) and 10\(^{-4}\), respectively. Determine whether the reaction above is at equilibrium and if not, predict what the concentration of HOCl will increase or decrease as the reaction proceeds.
Review of Simple Acid/Base Properties

Where HA is used to represent any generic Bronsted acid:

\[
\text{HA (aq)} + \text{H}_2\text{O(l)} \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{A}^-(aq)
\]

\[K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}\]

STRONG ACIDS completely dissociate and therefore have very large values of \(K_a\). Examples include HCl, HNO₃, HClO₄ and H₂SO₄. The conjugate base of strong acids are very weak bases. Therefore, Cl⁻, NO₃⁻, ClO₄⁻ and SO₄²⁻ are examples of weak bases and are said to be neutral anions.

WEAK ACIDS remain largely undissociated and have relatively small values of \(K_a\). Examples include HF, HNO₂, HOC₁ and H₂S. The conjugate bases of weak acids are relatively strong bases. Therefore, F⁻, NO₂⁻, OCl⁻ and S²⁻ are examples of basic anions.

STRONG BASES are quantitatively converted into hydroxide ion in aqueous solution. NaOH, KOH, Ca(OH)₂ and Mg(OH)₂ are examples of strong bases, although the latter two have limited solubility in water.

WEAK BASES partially react with water to produce hydroxide ions.

\[
\text{NH}_3(aq) + \text{H}_2\text{O(l)} \rightleftharpoons \text{NH}_4^+(aq) + \text{OH}^-(aq)
\]

\[K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}\]

Water is amphoteric. It can both donate and accept a proton, thus it has the ability to act either as an acid or a base.

\[
\text{H}_2\text{O} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OH}^-
\]

\[K_w = [\text{H}_3\text{O}^+][\text{OH}^-] \quad K_w = 1.0 \times 10^{-14} \text{ at } 25^\circ\text{C}\]

In general: \(K_b \text{ (base)} = K_w/K_a \text{(conjugate acid)}\) and \(K_a \text{ (acid)} = K_w/K_b \text{(conjugate base)}\)

Thus, \(K_a \text{ (NH}_4^+) = K_w/K_b \text{(NH}_3)\)

\[\text{pH} = -\log [\text{H}^+] \approx -\log [\text{H}^+]\]

In general: \(\text{pH} + \text{pOH} = \text{pK}_w\)

In pure water at 25°C; \(\text{pH} = \text{pOH} = 7.0\)
Rules for Predicting the pH of Aqueous Solutions

In general: Aqueous solutions of covalent oxides are acidic (i.e., CO$_2$, SO$_2$, SO$_3$ and NO$_2$).
Note: rainwater is naturally acidic (pH ~ 5.8 - 6.2) due to the presence of atmospheric CO$_2$

Aqueous solutions of ionic oxides are basic (i.e., CaO, MgO and K$_2$O)

Neutral ions: All Group 1A and 2A cations (e.g., Na$^+$, Ca$^{2+}$ etc)
Anions which are conjugates of strong acids (Cl$^-$, NO$_3^-$ etc)

Acids: Uncharged strong or weak acids (e.g., HCl, H$_2$CO$_3$ etc.)
Metal cations other than those of Groups 1A and 2A
Cationic conjugates of nitrogen bases (e.g., NH$_4^+$, C$_5$H$_5$NH$^+$ etc)
The two anions HSO$_4^-$ and H$_2$PO$_4^-$

Bases: All anions not listed above
Uncharged nitrogen bases, such as NH$_3$, C$_3$H$_5$N etc

Sample Exercises:
11. State whether the following aqueous solutions are expected to be acidic, basic or neutral
   a) CuCl$_2$
   b) Na$_3$PO$_4$
   c) NH$_4$Br
   d) MgS

12. Calculate the concentration of ammonium chloride in a solution whose pH is measured as 5.13, given that the $K_a$ for the ammonium ion is $5.6 \times 10^{-10}$ M.

13. Calculate the value of $K_b$ for cyanide ion, given that a $1.6 \times 10^{-2}$ M solution of potassium cyanide has a pH of 10.76. (Compare your answer with that obtained using the $K_a$ for HCN).

14. Commercial bleach is a 5% (wt/wt) solution of sodium hypochlorite. The value of $K_a$ for hypochlorous acid (HOCl) is $2.8 \times 10^{-8}$.
   a) Do you expect bleach to be acidic or basic?
   b) Calculate the $K_a$ of hypochlorite ion.
   c) Calculate the pH of a 5% bleach solution.
Solubility (Dissolution and Precipitation)

You are expected to be able to:

• use the solubility rules to predict solubility behaviour of binary compounds.
• calculate the solubility product, $K_{sp}$, given the solubilities and vice versa.
• explain the effect of pH and a common ion on solubility.

Solubility rules for ionic compounds
1. Almost all salts of the Group 1A (alkali metals) and of NH$_4^+$ are soluble in water
2. All nitrates are soluble in water
3. Most chlorides, bromides and iodides are soluble, the exceptions are those of Ag$^+$ and Hg$_2^{2+}$
4. Most sulfates are soluble, the major exceptions being those of barium, lead, mercury, bismuth and tin. Calcium sulfate is sparingly soluble.
5. Most carbonates, oxides, hydroxides, phosphates and sulfides are insoluble. Exceptions are the alkali metals (rule 1). Calcium and barium hydroxides are fairly soluble.

For the general case where MX(s) represents an ionic compound;

$$MX(s) \rightleftharpoons M^+(aq) + X^-(aq)$$

$$K_{sp} = [M^+][X^-]$$

And solubility (s) is equal to the number of moles of solid dissolved per liter of solution.

In general, ionic compounds are less soluble in a solution that contains either ion from another source. Thus calcium sulfate is less soluble in a solution that contains calcium from another source, such as calcium carbonate.

In general, ionic compounds that contain a basic anion become more soluble in acidic solution as the protons present will associate with the basic anion thus effectively lowering their concentration. Thus, hydroxides, sulfides, carbonates, fluorides and phosphates all become more soluble at low pH.

Sample Exercises:

15. In which of the following cases would you expect precipitation to occur when equal volumes 0.1 M solutions are mixed?
   a) calcium chloride and silver nitrate
   b) potassium sulfide and cadmium chloride
   c) barium hydroxide and copper (II) sulfate

16. Given that the $K_{sp}$ of CaF$_2$ is $5.3 \times 10^{-9}$ M$^3$
   a) Calculate the solubility of CaF$_2$
   b) Is calcium fluoride more or less soluble in basic solution?
   c) Is calcium fluoride more or less soluble in ‘hard’ water?
Rules For Assigning Oxidation States

1. For free elements the oxidation state is zero.
   e.g.  Fe(s), O₂(g), O₃(g), H₂(g), Hg(l), Hg(g), S(s) etc.

2. For monoatomic ions, the oxidation state is given by the charge on the ion.
   e.g.  Cl⁻ (-1), Fe²⁺ (+2), Fe³⁺ (+3), S²⁻ (-2), Ca²⁺ (+2), H⁺ (+1) etc

3. Certain elements when present in compounds have common oxidation states.
   a)  alkali metals (Li⁺, Na⁺, K⁺) are always +1
   b)  alkali earth metals (Mg²⁺, Ca²⁺, Sr²⁺, Ba²⁺) are always +2
   c)  hydrogen is +1 (except in metal hydride compounds such as LiH)
   d)  oxygen is -2 (except in peroxides such as H₂O₂)
   e)  halogens (F⁻, Cl⁻, Br⁻, I⁻) are usually -1 (exceptions include interhalogen compounds
       e.g., in Cl₂F₇, chlorine is +7 and fluorine is -1 and oxyanions e.g., in ClO₃⁻ chlorine is +5)

4. The sum of the oxidation states in a molecule is zero.
   e.g.  H₂O  (+1) + (+1) + (-2) = 0  Fe(OH)₂  (+2) + 2(-2) + 2(+1) = 0

5. The sum of the oxidation states in an ion is equal to the charge on the ion.
   e.g.  OH⁻ (-2) + (+1) = -1  SO₄²⁻ (+6) + 4(-2) = -2

Note: Oxidation corresponds to an increase in the oxidation state and reduction corresponds to a a
decrease in the oxidation state.

Sample Exercises:
17. Determine the oxidation states for all of the atoms in each of the following:
   a)  NO₃⁻, NH₃, NH₄⁺, N₂
   b)  Na₂S, Na₂SO₃, Na₂S₂O₃, Na₂SO₄
   c)  ClO₄⁻, ClO₃⁻, ClO₂⁻, ClO⁻
   d)  CO₂, H₂CO₃, C₂H₅OH, CH₃CHO

18. Indicate whether the following processes involve oxidation or reduction
   a)  SO₄²⁻ ---> H₂S
   b)  NH₄⁺ ---> NO₃⁻
   c)  NaClO ---> Cl⁻
   d)  2 Cu⁺ ---> Cu²⁺ + Cu

19. In the following reactions identify the species that is oxidised and that being reduced.
   a)  IO⁻¹ + I⁻ + H⁺ ---> I₂ + H₂O
   b)  NO₃⁻ + H⁺ + Cl⁻ ---> NO + Cl₂ + H₂O
   c)  NO₃⁻ + Cu + H⁺ ---> NO₂⁻ + Cu²⁺ + H₂O
Thermodynamics Review

Gibbs free energy changes: $\Delta G$

Reactions spontaneously proceed in a direction to lower free energy.

In general:
- If $\Delta G < 0$, then the reaction is spontaneous in the direction written
- If $\Delta G > 0$, then the reaction is spontaneous in the opposite direction
- If $\Delta G = 0$, then the reaction is at equilibrium

When all species are present at their standard state (1 atm for gases and 1 M for species in solution), free energy changes are denoted as $\Delta G^\circ_{\text{rxn}}$ for any given reaction. Standard free energy changes can be determined by any one of four different approaches depending on the type of experimental data available.

1. From standard enthalpies of formation ($\Delta H^\circ_f$) and entropy changes ($\Delta S^\circ$)
   \[ \Delta G^\circ_{\text{rxn}} = \Delta H^\circ_{\text{rxn}} - T \Delta S^\circ_{\text{rxn}} \]

2. From standard free energies of formation, ($\Delta G^\circ_f$) of reactants and products
   \[ \Delta G^\circ_{\text{rxn}} = \sum \Delta G^\circ_f (\text{products}) - \sum \Delta G^\circ_f (\text{reactants}) \]

3. From equilibrium constants (K).
   \[ \Delta G^\circ_{\text{rxn}} = -RT \ln K_{\text{eq}} \quad \text{note: } K_{\text{eq}} = \exp \{-\Delta G^\circ/RT\} \]
   where $R$ is the universal gas constant (R = 8.314 J mol$^{-1}$ K$^{-1}$) and $T$ is the Kelvin temperature

4. From electrochemical data ($E^\circ$)
   \[ \Delta G^\circ_{\text{rxn}} = -nFE^\circ \quad \text{note: } E^\circ = -\Delta G^\circ/nF \]
   where $n$ is the number of moles of $e^-$ transferred and $F$ is Faraday’s constant ($F=96,485$ C mol$^{-1}$)

To calculate free energy changes for systems not at standard state conditions ($\Delta G$), several approaches can be taken.

1. Using the reaction quotient (experimental species concentrations, i.e., non-standard state)
   \[ \Delta G = \Delta G^\circ + RT \ln Q \]
   where $Q$ is the reaction quotient

2. Using the experimental (non-standard state) electrochemical cell potential
   \[ \Delta G = -nFE \]
Review of Electrochemistry

Summary of definitions of electrochemical terms.

Oxidation: a chemical process in which the reactant loses electrons (involves an increase in oxidation state)

Reduction: a chemical process in which the reactant gains electrons (involves a decrease in oxidation state)

Anode: the electrode at which oxidation occurs

Cathode: the electrode at which reduction occurs

Galvanic Cell: an electrochemical cell which produces electricity, such as a flashlight battery

Electrolytic Cell: an electrochemical cell which is driven by the input of electricity, such as in the chlor-alkali process

Standard reduction potentials are thermodynamic parameters related to the driving force ($\Delta G$) for the reaction to proceed. They are tabulated for a great many half-reactions (written as reductions) in standard reference sources. These half-reactions are listed in order of decreasing values of $E^\circ$ (i.e., decreasing driving force to occur as reductions). Therefore, the most powerful oxidizers are those species at the top of standard reduction tables. Remember, large $E^\circ$ the greater the driving force to occur in the direction written. Hence species such as $F_2$, $O_3$, $H_2O_2$ and $Cl_2$ are strong oxidizing agents. At the other extreme, metals such as Li, Na and Mg are powerful reducing agents.

The potential for the complete cell is calculated from the individual half-reactions.

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} + E^\circ_{\text{anode}}$$

where $E^\circ_{\text{anode}}$ has the sign of the standard reduction potential reversed.

The greater the value of the cell potential, the greater the driving force for the reaction to proceed in the direction written. A more positive $E$ corresponds to a more negative $\Delta G$, the Gibbs free energy, since

$$\Delta G = -nFE$$

where, $n$ = number of electrons involved in a redox process
$F$ = Faraday constant (96,487 Coulombs/mol)
$E$ = cell potential

Recall, spontaneous reactions have negative $\Delta G$, but positive $E$

$\Delta G^\circ$, $E^\circ$ the values of $\Delta G$ or $E$ when all reactants and products are at unit activity (in practice these are the values found in tables or calculated form tabulated data).
Nernst Equation:

\[ E = E^\circ - \frac{RT}{nF} \ln Q \]

where,  \( n, F, E \) and \( E^\circ \) defined as before.
- \( R = \) universal gas constant \((8.314 \text{ J mol}^{-1} \text{ K}^{-1})\)
- \( T = \) temperature \((\text{Kelvin})\)
- \( Q = \) reaction quotient

Since \( \frac{RT}{F} = 0.0257 \) at 25°C, the Nernst equation can be written as at this temperature.

\[ E = E^\circ - \frac{0.0257}{n} \ln Q \]

At equilibrium, \( \Delta G = 0 \) and therefore \( E = 0 \)

Hence, \( E^\circ = \frac{0.0257}{n} \ln K \) 
where \( K \) replaces \( Q \) at the equilibrium condition.

This can be rearranged such that; \( K = \exp \{nE^\circ/0.0257\} \)

Sample exercises:

20. In each of the following, state whether oxidation or reduction is occurring.
   a) \( \text{CH}_4 \quad \longrightarrow \quad \text{CO}_2 \quad + \quad \text{other products} \)
   b) \( \text{Cl}_2(\text{g}) \quad + \quad 2e^- \quad \longrightarrow \quad 2 \text{Cl}^-(\text{aq}) \)
   c) \( \text{ClO}_2^- \quad \longrightarrow \quad \text{ClO}_2 \quad + \quad e^- \)

21. Using the standard reduction potentials, list the following species in order of increasing oxidizing ability. \( \text{ClO}_2 \quad \text{Cl}_2 \quad \text{O}_3 \quad \text{O}_2 \)

22. Use the standard reduction potentials to predict if a reaction will occur if copper metal will react with a solution containing silver ions, \( \text{Ag}^+ \). Calculate the \( E^\circ \) for the complete electrochemical cell.

23. Identify the half-reactions (neglecting spectator ions) and calculate the equilibrium constant, \( K_{eq} \).
   \( \text{Br}_2(\text{l}) \quad + \quad 2 \text{FeBr}_2(\text{aq}) \quad \longrightarrow \quad 2 \text{FeBr}_3(\text{aq}) \)