Oxidation-Reduction Reactions

- Reactions resulting in the change of oxidation numbers
- Redox reactions are always associated with transfer of electrons

\[\text{Zn} + 2\text{AgNO}_3 \rightarrow \text{Zn(NO}_3\text{)}_2 + 2\text{Ag}\]

- In this reaction Zn is oxidized (the oxidation number increases) and Ag is reduced (the oxidation number decreases)
- Electrons are transferred from Zn to Ag
Oxidation-Reduction Reactions

- Electrons can be neither created out of nothing nor destroyed.
- In any redox reaction there is an element being reduced and an element being oxidized.
- The total increase in the oxidation numbers must equal the total decrease in the oxidation numbers.

If an element is reduced:
- It gains electrons.
- Oxidation number decreases.
- The substance is classified as an oxidizing agent.

If an element is oxidized:
- It loses electrons.
- Oxidation number increases.
- The substance is classified as a reducing agent.
Balancing Redox Reactions

1. Determine oxidation numbers for all elements in each compound involved in the reaction

2. Separate the oxidation and reduction processes and write them as half-reactions

3. Balance each half-reaction by inspection and add the necessary number of electrons to balance the charge

4. Multiply the half-reactions by integer numbers to equalize the numbers of electrons gained and lost in each

5. Add the half-reactions and cancel any common terms to get the balanced equation
Example 1

- Determine oxidized and reduced species and balance the equation:

\[ \text{F}_2 + \text{NaBr} \rightarrow \text{NaF} + \text{Br}_2 \]
Example 2

 Determine oxidized and reduced species and balance the equation:

\[ \text{Zn} + \text{AgNO}_3 \rightarrow \text{Zn(NO}_3)_2 + \text{Ag} \]
Redox Reactions in Aqueous Solutions

- If a redox reaction is carried out in aqueous solution, we might need to introduce \( H_2O \) molecules, \( H^+ \) or \( OH^- \) ions to balance it.

In acidic solution:
- Add only \( H_2O \) or \( H^+ \)
- Add \( H_2O \) to balance O atoms
- Add \( H^+ \) to balance H atoms

In basic solution:
- Add only \( H_2O \) or \( OH^- \)
- Add \( H_2O \) and \( OH^- \) to balance O and H atoms
Example 3

Determine oxidized and reduced species and balance the equation:

\[ \text{FeCl}_2 + \text{H}_2\text{O}_2 + \text{HCl} \rightarrow \text{FeCl}_3 + \text{H}_2\text{O} \]
Example 4

- Determine oxidized and reduced species, write and balance the net ionic equation, and then derive the formula unit equation:

\[ \text{Bi}_2\text{O}_3 + \text{NaOH} + \text{NaOCl} \rightarrow \text{NaBiO}_3 + \text{NaCl} + \text{H}_2\text{O} \]
Example 5

The citrate ion, $C_2O_4^-$, is oxidized by the permanganate ion, $MnO_4^-$, in the sulfuric acid solution, forming carbon dioxide and $Mn^{2+}$ ion. Write and balance the net ionic equation, and then derive the formula unit equation for this reaction.
Example 6

- The hydrogen sulfate ion, HSO$_4^-$, is reduced by aluminum metal in sodium hydroxide solution, forming aluminum oxide and S$^{2-}$ ion. Write and balance the net ionic equation, and then derive the formula unit equation for this reaction.
Example 7

525 mL of iodine solution was titrated with 7.28 mL of 0.2 M nitric acid solution producing iodic acid and nitrogen(IV) oxide. What is the concentration of the iodine solution?
Example 8

What mass of N$_2$H$_4$ can be oxidized to N$_2$ by 24.0 g K$_2$CrO$_4$, which is reduced to Cr(OH)$_4^-$ in basic solution?
Assignments & Reminders

- Read Chapter 11 completely
- Read Section 4-7 of Chapter 4
- Extra Review Session - 5 to 7 pm TODAY in 105 Heldenfels
- Review Session for Exam #3 - 5:15 to 7:15 pm on Sunday in 100 Heldenfels
CHAPTER 12

- Gases and the Kinetic-Molecular Theory
Gases vs. Liquids & Solids

**Gases**
- Weak interactions between molecules
- Molecules move rapidly
- Fast diffusion rates
- Low densities
- Easy to compress

**Liquids & Solids**
- Strong interactions between molecules
- Molecules move slowly
- Slow diffusion rates
- High densities
- Hard to compress
Pressure

- Force per unit area
- Units of pressure:
  - pounds per square inch (psi)
  - mm Hg = torr
  - atmospheres (atm)
  - pascals (Pa)
- Normal atmospheric pressure - the pressure of air at the sea level at 0°C (=32 F)
  - $1 \text{ atm} = 760 \text{ mm Hg} = 760 \text{ torr} = 101,325 \text{ Pa} \approx 101.3 \text{ kPa}$
  - (Evangelista Torricelli - 1608-1647)
Kinetic-Molecular Theory

- Explains the behavior of gases in terms of molecular motion.
- The kinetic energy of gas molecules depends on their velocities:
  \[ E = \frac{mv^2}{2} \]
- The gas exerts pressure due to the molecular motion: many molecules have to strike the surface to produce this macroscopic effect.
Boyle's Law

- $p \times V = \text{const} = k$

- $p$ - pressure
- $V$ - volume

(at constant temperature and amount of gas)

$p_1 \times V_1 = p_2 \times V_2$
Boyle’s Law - Molecular Picture

\[ p_1 \times V_1 = p_2 \times V_2 \]

- The amount of gas (the number of gas molecules) remains constant.
- The temperature is constant and therefore the kinetic energy of gas molecules remains about the same.
- If the volume is decreased, then higher number of gas molecules strike a unit area, therefore the pressure increases.
- If the volume is increased, the reverse effect takes place - the pressure decreases.
Boyle’s Law - Example

- A 1.00 L sample of gas at 760 mm Hg is compressed to 0.800 L at constant temperature. Calculate the final pressure of the gas.
Charles’ Law

- $V \propto T$ or $V = kT$
  (at constant pressure and amount of gas)
- This equation defines a straight line
- Extrapolating this line to $V=0$ results in the \textbf{absolute zero} of temperature on the Kelvin temperature scale
Charles' Law - Molecular Picture

- The amount of gas (the number of gas molecules) remains constant.
- As the temperature increases, the thermal energy is converted into the kinetic energy and gas molecules move faster.
- The gas molecules strike the surface more vigorously and, if the pressure is to be kept constant, the gas has to expand.
- If the temperature is decreased, the volume also decreases.

\[
\frac{V_1}{T_1} = \frac{V_2}{T_2}
\]
Charles’ Law - Example

- A sample of gas at 1.20 atm and 27°C is heated at constant pressure to 57°C. Its final volume is 4.75 L. What was its original volume?
Combined Gas Law

- For a constant amount of gas

\[ \frac{p_1 V_1}{T_1} = \frac{p_2 V_2}{T_2} \]

- Both Boyle’s Law and Charles’ Law can be derived from the Combined Gas Law
Combined Gas Law - Example

- A 4.00 L sample of gas at 30°C and 1.00 atm is changed to 0°C and 800 mm Hg. What is its new volume?
Avogadro’s Law

- At the same temperature and pressure, equal volumes of all gases contain the same number of molecules.

- At constant $T$ and $p$, the volume $V$ occupied by a sample of gas is directly proportional to the number of moles $n$

$$V \propto n \quad \text{or} \quad V = kn$$
The standard molar volume of an ideal gas is equal to 22.414 liters per mole at standard temperature and pressure.

Standard temperature and pressure (STP)

- $T = 273.15 \, \text{K} = 0^\circ \text{C} = 32^\circ \text{F}$
- $p = 760 \, \text{torr} = 1 \, \text{atm} = 101,325 \, \text{Pa}$

1 mole of an ideal gas occupies 22.414 L volume ONLY at standard temperature and pressure.

To find the volume of 1 mole at different conditions we have to use other gas laws.
What volume will be occupied by 32.0 g of oxygen at STP? How will this volume change if the pressure is increased to 3 atm and the temperature is raised to 100°C?