Solutions:

- Concentrations of ions from salts: If a 0.10 M soluble salt is placed in water, the concentration of each ion depends on the ratio of ions in the salt. For example, a salt with the formula MX₂ dissociates in a 1:2 ratio. The cation will have the same concentration as the salt, but the anion’s concentration will be twice as much.

- Dilution: \( M_1V_1 = M_2V_2 \). Use this to calculate the concentration of a substance after it is mixed with other solutions or diluted with water. Remember: \( V_2 \) = total volume of the entire solution!

- Solution Stoichiometry: When we use aqueous solutions, we know their concentrations and/or volumes. \( M \times V = \) moles of solute in known (starting) solution. Then the mol:mol ratio from the balanced equation can be used to find the moles of the unknown solution. The final step in the calculation is to use moles and volume to find concentration (mol/vol) or to use moles and concentration to find volume (mol/concentration).

Kinetics:

- The main net ionic reaction is: \( 2 \Gamma^- + S_2O_8^{2-} \rightarrow I_2 + 2 SO_4^{2-} \) (reaction 1)

- The generic rate law is written as \( \text{Rate} = k [\Gamma^+]^x [S_2O_8^{2-}]^y \)

- As soon as \( I_2 \) is produced in the main reaction, it immediately reacts with the thiosulfate ion in the clock reaction: \( I_2 + 2 S_2O_3^{2-} \rightarrow 2 I^- + S_4O_6^{2-} \) (reaction 2). Once all the thiosulfate is consumed, the reaction mix turns blue as the \( I_2 \) that continues to accumulate in reaction 1 now reacts with starch. Since the same amount of thiosulfate is added to all runs, it reacts with a constant amount of \( I_2 \); thus the concentration of \( I_2 \) is constant for all runs. \( [I_2] = 1/2 [S_2O_3^{2-}] \)

- The rate of the reaction can then be determined using \( \text{Rate} = \Delta [I_2] / \Delta t \)

- \( M_1V_1 = M_2V_2 \) is used to determine the final concentrations for \( \Gamma^- \) and \( S_2O_8^{2-} \) in the reaction mixtures.

- The values of \( x \) and \( y \) and overall order are determined from the initial rates method using a table of experimental data (Rate of Reaction, \( [\Gamma^+] \) and \( [S_2O_8^{2-}] \) for different runs)

- The rate constant, \( k \), is determined from the experimental rate law: \( \text{Rate} = k [\Gamma^+]^x [S_2O_8^{2-}]^y \) where the known values for \( x \), \( y \), Rate, \( [\Gamma^+] \) and \( [S_2O_8^{2-}] \) are plugged in for each run.

- The average \( k \) for the 5 room temperature trials is calculated.

- The final experimental rate law is reported with the constants plugged in (\( x \), \( y \) and the average \( k \)).

- The activation energy, \( E_a \), is found from an Arrhenius plot of \( \ln k \) versus \( 1/T \), where the slope of the plot = \(-E_a/R\). The effect of temperature change on the reaction rate is examined.

- The effect of a catalyst on the reaction rate is examined.

Le Chatelier’s Principle:

- Equilibrium mixes contain a mixture of both reactants and products. If \( K \) is small there are more reactants present, and if \( K \) is large then more products are present.

- Addition of a reactant will shift the equilibrium position toward products.

- Addition of a product will shift the equilibrium position toward reactant.

- Removal of a reactant (by addition of a substance not in the original equation that causes a side reaction) will shift the equilibrium position toward reactants to replace its concentration.

- Removal of a product (by addition of a substance not in the original equation that causes a side reaction) will shift the equilibrium position toward products to replace its concentration.

- Adding and removing heat and interpreting directions of shift can help determine if a reaction is endothermic or exothermic.
  - Endothermic: heat acts as a reactant  \( \text{Heat} + \text{Reactants} \rightleftharpoons \text{Products} \)
  - Adding heat will shift the equilibrium toward products.
Removing heat will shift the equilibrium toward reactants.
- Exothermic: heat acts as a product: Reactants $\Leftrightarrow$ Products + Heat
- Adding heat will shift the equilibrium toward reactants.
- Removing heat will shift the equilibrium toward products.

**Introduction to Strong and Weak Acids:**
- pH values of acids are determined by the amount of $H_3O^+$ produced when they dissociate in water.
- $pH = -\log [H_3O^+]$ can be rearranged to $[H_3O^+] = 10^{-pH}$
- The % of dissociation/ionization is calculated from: $\% \text{Ionization} = \frac{[H_3O^+]}{[HA]_{\text{initial}}} \times 100\%$
- Strong acids dissociate to a much larger extent than weak acids. Therefore, two acids with the same concentration (one strong and one weak) will have different pH values. The stronger acid will produce more hydronium in solution and have a lower pH.
- Diluting a weak acid will increase its percent ionization (Le Chatelier’s principle). Diluting the acid does not change whether it’s a strong or weak acid.
- A weak acid can be more acidic (have a lower pH) than a strong acid if the concentration of the weak acid is much larger thus producing more hydronium ions in solution.
- 2 monoprotic acid solutions with approximately the same concentration (and same volume) will require the same volume of strong base to reach the equivalence point. It is the moles of acid present (and number of protons) that determines the volume of base needed, not the strength of the acid!
- Use the balanced neutralization reaction and stoichiometry to solve for the concentration or volume of acid or base in an acid-base titration problem.

**Weak Acids, Weak Bases, and Salts:**
- Salts are made up of cations and anions, and come from combining acids and bases together. The cation comes from a base, the anion comes from an acid. The relative strength of the acid and base will determine if the salt is acidic, basic, or neutral.
  - Acidic salts: weak base + strong acid (ex: NH₄Cl: NH₄⁺ is the conjugate acid of the weak base, NH₃. NH₄⁺ is an acidic cation – it will hydrolyze to produce hydronium ions. Cl⁻ is neutral because it comes from a strong acid and won’t have any further reaction with water.) The acidic cation will hydrolyze – this is a WA hydrolysis reaction.
  - Basic salts: strong base + weak acid (ex: KF: K⁺ is neutral because it comes from a strong base and won’t have any further reaction with water. F⁻ is the conjugate base of the weak acid, HF. F⁻ is a basic anion - it will hydrolyze to produce hydroxide ions.) The basic anion will hydrolyze – this is a WB hydrolysis reaction.
  - Neutral salts: strong base + strong acid (ex: NaCl: both ions are neutral because they come from strong base and strong acid; neither one will react with water.) Neutral cations (group IA and IIA) and neutral anions (Cl⁻, Br⁻, I⁻, NO₃⁻ and ClO₄⁻) will NOT hydrolyze.
- Two chemical equations can be written for a salt solution: dissociation of the salt into its ions and the hydrolysis equation for the ion that is a conjugate of a weak acid or base.
  - Dissociation: salt dissociates into ions. Ex: NaF $\rightarrow$ Na⁺(aq) + F⁻(aq).
  - Hydrolysis: the ion in the salt from the weak acid or weak base will set up equilibrium in water. Ex: F⁻(aq) + H₂O(l) $\Leftrightarrow$ HF(aq) + OH⁻(aq)
- $K_a$ values can be calculated for weak acids, $K_a = \frac{[H_3O^+][A^-]}{[HA]}$
- $K_b$ values can be calculated for weak bases, $K_b = \frac{[BH^+][OH^-]}{[B]}$
• After measuring pH values of acid, base, or salt solutions, the concentration of hydronium \([\text{H}_3\text{O}^+] = 10^{-\text{pH}}\) or hydroxide \([\text{OH}^-] = 10^{-\text{pOH}}\) can be calculated. This calculated value represents the equilibrium \([\text{H}_3\text{O}^+]\) (or \([\text{OH}^-]\)); this is the equilibrium value (x) in an ICE table. The equilibrium constant (\(K_a\) or \(K_b\)) is then calculated by plugging in the equilibrium concentrations for all substances present.
• Equilibrium constants can be calculated for ions depending on whether they produce hydronium (\(K_a\)) after hydrolysis or hydroxide (\(K_b\)) after hydrolysis.
• As acid strength increases, pH decreases, \(K_a\) increases and % ionization increases
• As base strength increases, pH increases, \(K_b\) increases and % ionization increases

**Acid-Base Titrations:**

- **Techniques for a titration include**
  - Conditioning and filling a pipette with an acid
  - Conditioning and filling a burette with a base
  - Be sure to fill the tip and record the volume to 2 decimal places.
  - Record the initial and final volumes for the titration

- Titration curves were obtained for a strong acid and a weak acid. Different regions of the curves were compared and the \(pK_a\) value was obtained for a weak acid.

- **When a strong acid is titrated by a strong base:**
  - The pH is initially low since the acid completely ionizes to form hydronium ions
  - Before the equivalence point the moles of acid are larger than the moles of base thus the pH depends on the concentration of the excess acid present in solution.
  - The pH at the equivalence pt is 7 because the acid and base have completely reacted and we are left with water and a neutral salt in solution.
  - After the equivalence point, the excess sodium hydroxide in solution causes the pH to be high. pH depends on the concentration of excess NaOH present in solution

- **When a weak acid is titrated by a strong base:**
  - The initial pH is a little higher than for the strong acid. The weak acid forms less hydronium ions since it doesn’t ionize completely. (WA calc - use \(K_a\) and the \([\text{HA}]\) to find pH)
  - Half way to the equivalence point, we have a buffer solution containing equal amounts of HA and \(A^-\) so the p\(H = pK_a\). (Buffer region – use change table to find moles of acid remaining and the moles of conj base formed then plug into HH to find the pH)
  - Estimating the pH at the half-eq pt on the WA plot gives us the \(pK_a\), thus \(K_a = 10^{-pK_a}\)
  - The pH at the equivalence pt is > 7 because after the acid and base are neutralized we are left with water and a basic salt in solution. (WB calc – use \(K_b\) and the \([A^-]\) to find p\(OH\) then convert it to pH)
  - The curve for the weak acid is the same as the strong acid after the equivalence point since excess strong base is present.

- Be able to write a balanced reaction for an acid-base neutralization reaction and use provided concentrations/ volumes and/or measured data to obtain the equivalence point volume or M for one of the solutions.
- Be able to calculate the pH at any point for a SA-SB titration and a WA-SB titration.

**Qualitative Analysis:**

- Qualitative analysis is the process of finding out what ions or compounds are contained in a sample based on their chemical reactivity and physical properties.
• In this experiment qualitative analysis methods to separate carbonate, sulfate, chloride, bromide and iodide ions.

• Qualitative analysis involves several different methods that can be used to separate and confirm the presence or absence of different ions.

• Centrifuging a sample greatly accelerates the settling of a precipitate suspended in solution allowing the precipitate to be separated from the supernatant solution above it.

• Litmus paper can be used to test for acidity (blue paper turns red) and basicity (red paper turns blue).

• Carbon dioxide gas is produced when an acid, such as nitric acid, is added to a solution containing carbonate. Thus, acidification allows carbonate ions to be separated from the other anions.

• Barium ions (from barium nitrate) will react with sulfate ions to form barium sulfate. The BaSO₄ precipitate should be separated from the supernatant solution above it (containing the halide ions).

• The halide ions (in the supernatant solution) react with AgNO₃ to form AgCl (white), AgBr (cream) and AgI (yellow) precipitates.

• After 6 M ammonia is added to the silver halide precipitates, the more soluble AgCl dissolves and forms a complex ion, Ag(NH₃)₂⁺ as well as chloride ions. The solid precipitates (AgBr and AgI) are then separated from the supernatant chloride (and complex ion) solution.

• When nitric acid is added to the solution containing chloride and Ag(NH₃)₂⁺ ions, the ammonia molecules are neutralized by the acid, thus freeing the silver ions. The Ag⁺ and Cl⁻ ions now combine to reform the AgCl precipitate.

• After 15 M ammonia is added to AgBr and AgI (and the solution is vigorously stirred), the silver bromide dissolves to form Ag(NH₃)₂⁺ and bromide ions. The solid ppt (AgI) is separated from the supernatant solution containing the bromide ions.

• Acidification of the supernatant solution neutralizes the ammonia molecules and frees the silver ions which now combine with bromide ions to reform the AgBr precipitate.

• The basifications steps illustrated how formation of complex ions can cause insoluble salts to dissolve.

• The flow chart is a useful tool that can be used to visualize the stepwise addition of the reagents and the separation and identification of the ions/precipitates as the scheme proceeds.

**Thermodynamics:**

• The law of conservation of energy states that heat/energy is always conserved.

  - \( q_{sys} = -q_{surr} \)
  - \( q_{sys} = -(\text{mass})\times(\text{specific heat capacity})\times(\Delta T) \)
  - \( \Delta H_{\text{rxn}} = \frac{q_{sys}}{\text{moles}} \)

• \( \Delta H_{\text{rxn}} \) - The change in enthalpy describes the heat of reaction. This is the heat gained or lost by the system during a reaction (@ constant P).

  - Exothermic (\( \Delta H \): heat is released by the system.
  - Endothermic (\( \Delta H \)): heat is gained by the system.
  - Breaking bonds absorbs heat/energy while making bonds releases heat/energy.
  - If T decreases when a salt dissolves in water, then the salt is absorbing heat from the water causing the water T to decrease = endothermic dissolution of a salt (\( -\Delta T \) but +\( \Delta H \) due to negative sign in equation from heat transfer between the solution and the salt)
  - If T increases when a salt dissolves in water then the salt is releasing heat to the water causing the T to rise = exothermic dissolution of a salt (+\( \Delta T \) but -\( \Delta H \))

• Entropy describes the amount of randomness in a system. The universe tends to moves towards disorder.

\[ \Delta S_{\text{Total}} = \Delta S_{\text{System}} + \Delta S_{\text{Surroundings}} \]
\[ \Delta S_{\text{Total}} > 0 \text{ (positive) moves towards disorder = spontaneous} \]
\[ \Delta S_{\text{Total}} < 0 \text{ (negative) moves towards order = nonspontaneous} \]

- Gibbs free energy predicts whether a reaction is spontaneous at a specific temperature.
  \[ \Delta G = \Delta H - T \Delta S \]
  - Spontaneous when \( \Delta G < 0 \) (negative)
  - Nonspontaneous when \( \Delta G > 0 \) (positive)
  - At equilibrium when \( \Delta G = 0 \)

- Not all reactions are spontaneous at all temperatures. The combination of \( \Delta H \), \( \Delta S \), and \( T \) (in Kelvin) in the equation \( \Delta G = \Delta H - T \Delta S \) and their magnitude determines the value of \( \Delta G \) and therefore predicts spontaneity.

**Electrochemistry:**

- Oxidation is the process of losing electrons/losing hydrogens/gaining oxygens
- Reduction is the process of gaining electrons/gaining hydrogens/losing oxygens
- An oxidizing agent oxidizes another species while it is reduced.
- A reducing agent reduces another species while it is oxidized.
- Reactions between solid metals and metal solutions can be used to place metals in order of activity.  The activity series is based on how easily a metal will be oxidized to a cation.
- The more active metal will be oxidized = anode
- Voltages of Galvanic Cells can be measured by setting up two compartments (in well plates or beakers) and connecting a wire to the two metals. Key components of the Galvanic Cell include:
  - Each metal is placed in a 1.0 M solution of the metal nitrate.
  - A salt bridge is made by soaking a piece of string or paper in a soluble ionic compound (i.e., \( \text{NaNO}_3 \)). The ends of the salt bridge must reach into the metal nitrate solution in each compartment of the cell.
  - Leads are connected to each metal piece. A voltmeter is connected to the wire and leads to read voltage. A negative value of voltage indicates that the leads are switched. (The black electrode is the anode, the red electrode is the cathode.)
  - Reduction and oxidation potentials of metals can be determined after the overall cell potential (\( E^\circ_{\text{cell}} \)) is measured and a standard half reaction potential is known. For example, in the Electrochemistry lab, the zinc half-reaction was used as the standard. If zinc was oxidized, the sign of its reduction potential was changed to represent the oxidation potential. \( E^\circ_{\text{cell}} = E^\circ_{\text{ox}} + E^\circ_{\text{red}} \). This equation can be manipulated to find the unknown metal’s half-reaction potential.
  - From the \( E^\circ_{\text{cell}} \) value, \( \Delta G^\circ \) can be determined: \( \Delta G^\circ = -nFE^\circ_{\text{cell}} \).
  - The equilibrium constant can then be calculated using \( E^\circ_{\text{cell}} = \frac{RT}{nF} \ln K \)
    - In both equations, \( F \) is Faraday’s constant: 96,500 J/mol e\(^-\)K
    - \( n \) = moles electrons (total electrons when the half-reactions are balanced)
    - \( R = 8.314 \text{ J/mol}\cdot\text{K} \)
    - \( T \) = temperature in Kelvin
  - Solid metals can not be reduced, they can only be oxidized. Metals are good reducing agents.
  - Metal cations in solution are reduced. They make good oxidizing agents.
  - A positive cell potential indicates a spontaneous reaction and will have a negative value of \( \Delta G^\circ \) as well as a large equilibrium constant, \( K > 1 \).
  - As the cell potential increases, \( \Delta G^\circ \) decreases and \( K \) increases.
Nuclear:
- Nuclear reactions produce entirely different substances as products, unlike chemical reactions in which particles are conserved.
  - Notations for nuclear particles are written with mass numbers as a superscript to the left of the element symbol. Atomic numbers (number of protons) may be written as a subscript to the left of the element symbol.
    
    \[
    \text{mass number} = A \\
    \text{atomic number} = Z
    \]
  - Isotopes are substances that differ in the number of neutrons. Ex: \(^{14}\text{C}\) vs \(^{12}\text{C}\).
- Nuclear particles include:
  - Alpha: \(^4\text{He}^2\)
  - Beta: \(^0\text{e}^- (-1\beta)\)
  - Gamma: \(^0\gamma\)
- The main types of nuclear decay are:
  - **alpha (a) emission (or decay):** a helium nucleus, \(^4\alpha\) or \(^4\text{He}^2\), is emitted. In this case, the parent and daughter are atoms of different elements and also have different masses.
  - **beta (b) emission (or decay):** an electron, \(^0\beta\) or \(^0\text{e}^-\), is emitted when a neutron inside an atom decays to produce a proton and an electron. In this case, the parent and daughter are of different elements but the mass remains the same.
  - **gamma (\gamma) emission (or decay):** high energy photons or gamma rays, \(^0\gamma\), are emitted. This generally accompanies the emission of a particle.
  - **positron (\text{e}) emission (or decay):** a positron, \(^0\text{e}\), is emitted when a proton inside an atom decays to produce a neutron and a positron. In this type of radioactive decay, the parent and daughter are different elements but have the same mass.
  - **electron capture:** an electron in the 1s orbital "falls" into the nucleus, which fuses with a proton to form a neutron.
- Nuclear equations are balanced by equalizing the mass numbers and atomic numbers of all reactants and products. A coefficient will count for both the mass number and atomic number.
  - Ex: \(^{238}\text{U}\) \(\rightarrow\) \(6\text{e}^-\) + \(8\text{He}^4\) + _______ The unknown product can be predicted by balancing the equation. It should have an atomic number of 82 and a mass number of 206. The symbol is \(^{206}\text{Pb}\).
- Radioactive decay reactions all follow first order kinetics:
  \[
  t_{1/2} = \frac{0.693}{k} \quad \ln \left( \frac{A_t}{A_0} \right) = -kt
  \]
- If the half-life (t/2) of a particle is known, the rate constant (k) can be calculated. Since radioactive decay is always first order, the units of k will always be \(\text{time}^{-1}\) (ex: \(\text{s}^{-1}\), \(\text{min}^{-1}\), \(\text{yr}^{-1}\), etc.).
- Once the rate constant is determined, the percentage or amount of material left at the end of a length of time can be calculated. The equation on the right above can be rewritten as \(\ln (A_t) = -kt + \ln (A_0)\) where \(A_t\) represents the amount of substance left after time \(t\), \(k\) is the rate constant, \(t\) is the amount of time that has passed (how long the sample has been decaying), and \(A_0\) is the initial mass or amount of substance.
- Radioactive substances used for medical purposes tend to have short half-lives, whereas substances used for geological and radioactive dating have long half-lives.