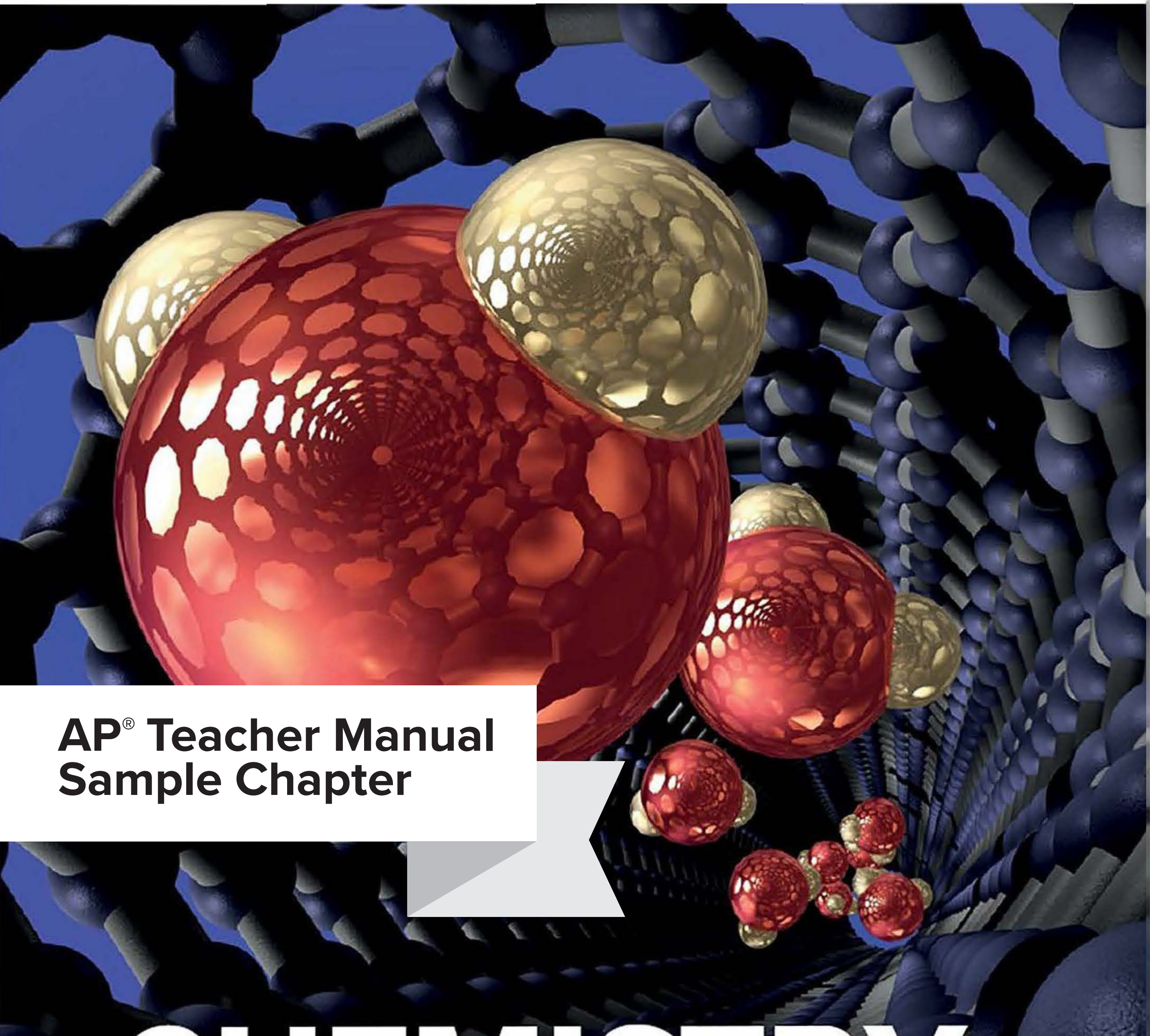

AP® Teacher Manual
Sample Chapter

CHEMISTRY

13e

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McGraw Hill Education

Chapter 4: Reactions in Aqueous Solutions

Chapter 4	Activities and Labs	Topics Covered
4.1 General Properties of Aqueous Solutions		3.7, 3.8
4.2 Precipitation Reactions	• CIA Feature: An Undesirable Precipitation Reaction (SE, p. 129) • Lab Activity: Solutions and Precipitates (Online Chapter Assignment, Chapter 4) • Metathesis Reactions (TM, pp. 12-13)	4.2, 4.3
4.3 Acid-Base Reactions	• Acid-Base Reactions (TM, p. 14)	4.2, 4.3, 4.8
4.4 Oxidation-Reduction Reactions	• CIA Feature: Breathalyzer (SE, p. 146) • Single Replacement Reactions (TM, p. 14)	4.2, 4.3, 4.7, 4.9
4.5 Concentration of Solutions		3.7, 3.8
4.6 Gravimetric Analysis	• Gravimetric Analysis Activity (TM, p. 15)	1.1, 4.5
4.7 Acid-Base Titration	• Lab Activity: Urinalysis (Online Chapter Assignment, Chapter 4)	4.5, 4.6
4.8 Redox Titrations		4.5, 4.6
Review and Assessment	AP A Look Back at the Essential Knowledge (SE, p. 159) Key Equations (SE, p. 160) Questions & Problems (SE, pp. 161-171) AP Chapter Review (SE, pp. AP170-AP171) Focus Review Guide (pp. 35-47) Study Strategies (TM, p.15) Additional AP Practice Questions (TM, pp. 16-18)	

Chapter Overview

Several important concepts are addressed in this chapter. The concept of electrolytes, solubility, molarity, and solution preparation are covered in the initial section. The chapter then goes on to cover writing molecular equations, total ionic equations, and net ionic equations. These types of equations are used when writing equations for precipitation, redox, acid-base neutralization, combination, decomposition, single displacement, and combustion reactions. Solution stoichiometry and other stoichiometric calculations are practiced with each type of reaction. Titration procedures are discussed with both acid-base neutralization and redox reactions. The concepts of oxidation and reduction are defined, along with determination of oxidation numbers.

Concepts in the chapter fall mainly under **Scale, Proportion, and Quantity; Transformations.**

Addressing the Updated Curriculum Framework

As has been stated, reaction predication questions are on the exam, but not stressed as much. Your students will likely encounter a reaction prediction question in the free-response section of the exam, or rarely, in the multiple choice portion of the exam. If the question is located in the free response section, it is generally a “stand alone” question; no subsequent part of the question will depend on how that equation was balanced. Given the “internal consistence” grading method on the AP exam, it is highly unlikely that students would have to balance an equation first and then do a titration or stoichiometry problem with it. A common set of questions is “write the net ionic equation for ____” and the next part of the question is “explain why this reaction is best represented by a net ionic equation”. Students rarely encounter a balancing equation question in the multiple choice section. However, if the students are given a redox reaction to balance in the multiple choice section, it will show the unbalanced equation with blank lines before each species. The question will ask for a coefficient of one of the species. A more common multiple choice question is to identify the type of reaction and provide a justification. Teach your students to identify redox reactions by a change of oxidation numbers. The easiest way is to look for an element in the ground state on one side of the reaction and the same element as an ion or in a compound on the other side – then you know the oxidation number has change, hence redox reaction. This will help identify the type and if they are asked to balance it, to remind the students to break the reaction into its half-reactions.

Gravimetric analysis is a common laboratory process and is often seen on the free-response section of the exam. It can be on the multiple choice section, but the numbers involved would allow students to solve the problems without calculators. Redox and acid-base titrations are also another common lab based questions that more likely will be free response questions, but may show up as multiple choice (again, if the numbers allow the students to find the answers without a calculator).

For both types of titration a key concept is the equivalence point. Most teachers define the equivalence point in terms of the stoichiometry, i.e. the equivalence point is when the moles of acid and moles of base are equal. However, it is important to develop the concept that the equivalence point is the point at which the limiting reactant and the excess reactant switch rolls. In an acid-base titration of an acid being titrated with a base, before the equivalence point, the acid is always the excess reactant and the base (in the buret) is the limiting reactant. After the equivalence point, the base becomes the excess reactant. This concept leads to being able to identify particles in solution at any point along the titration curve, a skill needed in chapters 15 and 16. It will also help students to eliminate some of the choices in multiple choice questions.

Vocabulary

- Solution
- Solute
- Solvent
- Electrolyte
- Nonelectrolyte
- Hydration
- Reversible reaction
- Chemical equilibrium
- Precipitation reaction
- Precipitate
- Metathesis (double displacement reaction)
- Solubility
- Molecular equation
- Net ionic equation
- Spectator ion
- Brønsted Acid
- Brønsted Base
- Hydronium ion
- Monoprotic acid
- Diprotic acid
- Triprotic acid
- Neutralization reaction
- Salt
- Oxidation
- Reduction
- Oxidation-reduction reaction
- Half-reaction
- Oxidation number (oxidation state)
- Combination reaction
- Decomposition reaction
- Combustion reaction
- Single displacement reaction
- Disproportionation reaction
- Molarity
- Dilution
- Quantitative analysis
- Gravimetric analysis
- Titration
- Standard solution
- Equivalence point
- Indicator

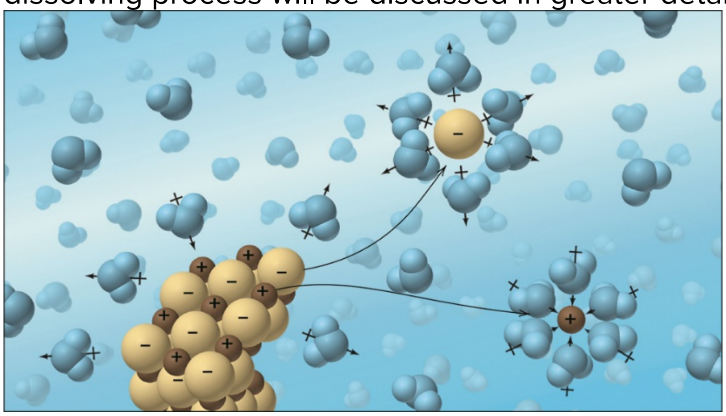
Pacing Guide

Ten days may be necessary to cover this material because of the homework practice necessary, even though much of it should be review for students. This includes two days for labs and one day for a test.

Key Concepts

Section 4.1 General Properties of Aqueous Solutions (EK SAP-5.B.2, SAP-5.B.3, SPQ-3.A.1, SPQ-3.B.1, TRA-1.D.1, TRA-1.D.2, 3.C)
Although it may be review, the concepts of solute vs. solvent must be covered. It is important to mention that, although most solutions are *aqueous*, other solvents can also be used. For example, when a compound is dissolved in an alcohol, the solution is said to be *alcoholic*.

The reasons for water's role as a universal solvent should be discussed. Water's bent shape and polar bonds cause it to have an unequal distribution of charge. This results in the hydrogen end of water having a partial positive charge and the oxygen end of water having a partial negative charge. Ionic compounds dissolve in water because of a strong attraction of the positive end of water for the anions and the negative end of water for the cations in the compound. The energies and attractions involved in this dissolving process will be discussed in greater detail in a later chapter.



Students should be able to correctly orient water molecules towards dissolved ions in a particulate drawing. For example, they could be asked, "Show the association of the ions with several water molecules when one formula unit of KCl dissolves in excess water". The students should be able to draw a K^+ ion surrounded by several water molecules with the oxygen ends of the water molecules oriented toward the K^+ ion. They would draw the Cl^- ion with water molecules surrounding it oriented so that the hydrogen ends are toward the Cl^- ion. Students should also be able to draw the cations as the smaller ions in solution, and the anions as the larger ions (as depicted in the picture above.)

Students need to be familiar with the terms *solvated* and *hydrated*. They also need to understand that compounds that have been termed *insoluble* actually have a very low solubility. This low solubility will be addressed in Chapter 16 when K_{sp} is discussed.

The concepts of strong electrolytes, weak electrolytes, and nonelectrolytes are discussed in this section. If these concepts are new to students, a brief demonstration of electrical conductivity of solids, aqueous solutions of ionic compounds, aqueous solutions of polar covalent compounds, and alcohols may be shown. Students should know that a conductivity test is a helpful tool to understanding the structure or identity of a compound.

Since soluble ionic compounds break into their individual ions when dissolved in water, the total number of moles of each ion present can be calculated if the original mass or moles of compound is known. Students should have practice determining moles of ions. A common mistake that students make is not breaking up compounds correctly. They may need to practice determining the number of ions in a formula unit with examples such as Na_2SO_4, $Mg_3(PO_4)_2$, and NaOH.

They should also recognize that many covalent (molecular) compounds, such as $C_{12}H_{22}O_{11}$, dissolve in water but do not usually ionize. The molecules that have positive and negative ends (polar molecules) and water molecules have mutual attractions that allow the dissolving process. These attractions will be discussed in greater detail in a later chapter.

This section also includes the first reference to chemical equilibrium and reversible reactions. This is a major topic in AP Chemistry (Units 7 and 8) and should be emphasized here. It is important that students understand that it is the rates of the forward and reverse reactions that are equal at equilibrium, not the concentrations of reactants and products. You might find the example of a popular new restaurant helpful. When the restaurant first opens for the day, the rate of people going in is greater than the rate of people coming out (not at equilibrium). When the restaurant fills up, no one can come in unless others come out (at equilibrium). This does not mean that the number of people waiting outside as the same as the number of people eating inside. When it gets close to closing time, the number of people coming out is greater than the number of people going in (not at equilibrium).

If you wish to teach the two thermochemistry chapters together (chapters 6 and 17), it may be helpful to also develop the concept of an equilibrium constant, K, at this point. This will give the students some context when they encounter the use of K in thermodynamics equations, such as $\Delta G = -RT \ln K$. Combining these two chapters (and giving a single test) does save a couple of days if you have difficulty finishing all of the material in the course.

Topic 3.7; Learning Objective SPQ–3.A: *Calculate the number of solute particles, volume, or molarity of solutions.*
Topic 3.8; Learning Objective SPQ–3.B: *Using particulate models for mixtures to (a) represent interactions between components and (b) represent concentrations of components.*

Section 4.2 Precipitation Reactions (EK SAP-5.B.2, SPQ-3.B.1, SPQ-3.C.2, SPQ-4.A.2, TRA-1.A.2, TRA-1.B.1, TRA-1.B.2, TRA-1.B.3, TRA-1.C.1, TRA-2.A.5, TRA-6.A.1)
Precipitation reactions are introduced as a specific type of metathesis (also known as double displacement or double replacement) reaction. Students should understand that solids are only called *precipitates* if they are formed when ions are removed from solution (solid product of a metathesis reaction). Solid products of other types of reactions are not called precipitates. The redesigned AP® Chemistry exam no longer requires extensive memorization of the solubility rules. Students should know that compounds containing Na^+, K^+, NH_4^+, or NO_3^- are soluble. Other solubility rules needed would be given in words, tables, or could be interpreted from laboratory results given to the students on the test. Students should have practice using solubility tables and interpreting laboratory data.

Precipitates form when the Coulombic attraction between the ions of the precipitate is greater than the attraction of the water molecules for the ions. Since the ions have strong attraction for each other, they tend to form a precipitate, resulting in a cloudy solution or a solid material forming at the bottom of the reaction container.

This section introduces the writing of molecular equations, ionic equations (also known as complete or total ionic equations), and net ionic equations. Students often make mistakes writing net ionic equations and will require extensive practice. When students first start to write net ionic equations, writing all three types of reactions is especially helpful. In the next sections of the chapter, students will gain more practice as they also learn to predict products of simple reactions. Students will have to write net ionic equations on the AP Chemistry exam, so it is a very important skill to master. Practice throughout the year is necessary to maintain this skill.

In figures 4.7 and 4.8, particulate drawings are shown to represent the reactions discussed. Students need to be able to translate between particulate and symbolic (chemical equation) views of the same process.

Although not addressed in any of the 16 College Board labs, it is helpful to do a lab (or mini-lab) where students are performing metathesis reactions, making observations, and writing net ionic equations for the reactions that occurred. See Activity #1.

* Introduce the Lab Activity: Solutions and Precipitates

Topic 4.2; Learning Objective TRA–1.B: *Represent changes in matter with a balanced chemical equation.*

Topic 4.3; Learning Objective TRA–1.C: *Represent a given chemical reaction or physical process with a consistent particulate model.*

Section 4.3 Acid-Base Reactions (EK SAP-9.A.1, SAP-9.B.1, SAP-9.B.2, SAP-9.D.1, TRA-1.A.2, TRA-1.C.1, TRA-2.A.1, TRA-2.B.1, TRA-2.B.2, TRA-2.B.3, TRA-6.A.1)

Another type of metathesis reaction is an acid-base reaction (also called neutralization reaction).

Acid-base chemistry plays an important role in the AP® Chemistry course and will be revisited in chapters 15 and 16. In this type of metathesis reaction, an acid reacts with a base to make a salt and water.

$$\underset{\text{acid}}{HX(aq)} + \underset{\text{base}}{MOH(aq)} \rightarrow \underset{\text{salt}}{MX(aq)} + \underset{\text{water}}{H_2O(l)}$$

The products of this type of reaction do not always have a pH of 7, even though it is called *neutralization*. Later in the course, the acid-base properties of the resulting salt will be discussed.

Students should be comfortable with the idea that $[H^+]$ and $[H_3O^+]$ may be used interchangeably.

Table 4.3 lists the common strong acids. This is the same list that the AP® Chemistry course framework mentions that the students should know. The AP® Chemistry course expects students to treat <u>all</u> Group 1 and 2 hydroxides as strong. The lighter Group 2 hydroxides are not very soluble, but the portion that dissolves ionizes 100%.

The Brønsted-Lowry acid/base definitions are introduced in this chapter and will be treated in greater detail (including identification of conjugate acids and conjugate bases) in Chapter 15. The Brønsted-Lowry acid/base concept is the main focus of the acid/base chemistry on the exam.

Students are expected to remember that when carbonic acid (H_2CO_3) forms in a metathesis reaction, it breaks down immediately to water and carbon dioxide and also that sulfurous acid (H_2SO_3) breaks down into water and sulfur dioxide.

It is important that students understand that the net ionic equation for the reaction of a strong acid with a strong base will always be only $H^+ + OH^- \rightarrow H_2O$. When a weak acid or base is involved in a neutralization reaction, the net ionic equation is more complicated because the weak acid or base is left together and not ionized.

Topic 4.2; Learning Objective TRA–1.B: *Represent changes in matter with a balanced chemical equation.*

Topic 4.3; Learning Objective TRA–1.C: *Represent a given chemical reaction or physical process with a consistent particulate model.*

Topic 4.7; Learning Objective TRA–2.A: *Identify a reaction as acid–base, oxidation–reduction, or precipitation.*

Section 4.4 Oxidation-Reduction Reactions (EK SAP-8.C.1, SPQ-3.B.1, TRA-1.A.2, TRA-1.C.1, TRA-2.A.2, TRA-2.A.3, TRA-2.A.4, TRA-2.C.1, TRA-6.A.1)
Oxidation and reduction are introduced as the transfer of electrons or shifting of electrons from one atom to another. Mnemonic devices such as "LEO the Lion Says GER" (Lose Electrons Oxidation; Gain Electrons Reduction) or "OIL RIG" (Oxidation Is Loss; Reduction Is Gain) may help the student remember the definitions of oxidation and reduction. The terms *oxidizing agent* and *reducing agent* are also introduced. Although these terms will not be used on the AP® Chemistry Exam, students may still be asked which species causes another to be oxidized or reduced.

Simple redox reaction can be balanced by inspection, as was done with precipitation and neutralization reactions. More complex redox reactions require a more involved method. This half-reaction method is addressed in Chapter 18 of this text. It may be helpful to teach this method with Chapter 4 (and review in Chapter 18) as students would have more time to practice and would be able to use it with some of the laboratory experiments that require it throughout the year.

The concept of oxidation numbers is introduced and needs to be practiced by students. A fun and addictive website for practice with oxidation numbers from Stetson University is: http://www2.stetson.edu/mahjongchem/ . Students should be able to assign oxidation numbers to every atom in a chemical reaction and use change in oxidation number to determine which element is oxidized (if any) and which is reduced. If no oxidation numbers change, the reaction is not redox. Oxidation and reduction must both occur. If one element loses electrons, another must gain them.

Several other types of chemical reactions are also redox reactions. In this section, combination (synthesis), decomposition, single displacement (single replacement), and combustion reactions are discussed. Students should be able to predict products of simple reactions of these types. Reaction prediction will not be tested as heavily on the redesigned AP® Chemistry test as it was in past years, but students still need to know the basics of reaction prediction. They will not need to do extensive memorization of products formed in various reactions, but should be able to use their chemical logic (and experimental data) to determine products of fairly simple reactions. Reaction prediction will be found throughout the test, rather than in a separate section (the old Question #4) of the test.

To predict products of combination reactions ($A + B \rightarrow C$), students need to look at the reactants to determine if they are both nonmetals or if one is a metal and the other a nonmetal. If they are both nonmetals, the product will be covalent and determination of the correct product formula may be a little challenging. Students need to think about the common oxidation numbers of the elements involved and also look at their positions on the periodic table. For metal-nonmetal combinations, the product will be ionic and the student simply needs to think about the charges of the ions to put them together correctly to make a neutral compound. Combination reactions that involve one or more pure elements on the reactant side are redox reactions.

For combination reactions, it is also helpful to know:
Metallic oxides + carbon dioxide = metallic carbonate
Metallic oxide + sulfur dioxide = metallic sulfite
Metallic oxide + water = metallic hydroxide

Metallic chlorides + oxygen = metallic chlorate
Nonmetallic oxide + water = weak acid

Decomposition reactions (A $\rightarrow$ B + C) require energy to occur. While this is usually in the form of heat, electrolysis or ultraviolet radiation can also cause some compounds to decompose. If the reactant is a binary compound (2 elements), the products will usually be those two elements. Common mistakes involve using the subscripts from the compounds as subscripts for the elements and forgetting about the diatomic elements. Decomposition reactions that result in the formation of pure elements are redox reactions.

For decomposition reactions, it is also helpful to know:
Metallic carbonate = metallic oxide + carbon dioxide
Metallic sulfite = metallic oxide + sulfur dioxide
Metallic chlorate = metallic chloride + oxygen
Hydrogen peroxide = water + oxygen
Sulfurous acid = water + sulfur dioxide
Carbonic acid = water + carbon dioxide
Hydrated salts = anhydrous salt + water

Combustion reactions are also redox reactions. When a reactant is burned in oxygen, the products are oxides of the elements in the reactants. In first year chemistry, students are usually taught that combustion always involves compounds containing carbon, such as hydrocarbons, alcohols, and sugars. Combustion can also involve other elements. The largest challenge seems to be correctly balancing the reactions.

Single Displacement (single replacement) reactions involve an element reacting with a compound. The element will replace one of the elements in the compound (metal replacing metal, nonmetal replacing nonmetal). In order to predict whether a single displacement reaction will occur, some knowledge of the activity series is helpful. It does not need to be memorized. If students need more information, such as reduction potential values, that information will be provided in the question or set of questions on the AP® Exam. There is no longer a reduction potential table provided for the AP® Chemistry Exam.

The activity series of metals is found in Figure 4.16. It is helpful if students simply break the series into these four parts from top to bottom (rather than memorizing it all).

1. *The most active metals are found in Groups 1,2, & 13 of the periodic table. (most easily oxidized)*
2. *The transition metals are next in reactivity.*
3. *Hydrogen*
4. *The "jewelry" or "inert" metals are least reactive (least easily oxidized)*

For the nonmetals (halogens), students should understand the order of reactivity is the same as their vertical order on the periodic table.

Fluorine is most reactive (most easily reduced)>Cl>Br>I

Topic 4.2; Learning Objective TRA–1.B: *Represent changes in matter with a balanced chemical equation.*

Topic 4.3; Learning Objective TRA–1.C: *Represent a given chemical reaction or physical process with a consistent particulate model.*
Topic 4.7; Learning Objective TRA–2.A: *Identify a reaction as acid–base, oxidation–reduction, or precipitation.*

Section 4.5 Concentration of Solutions (EK SPQ-2.B.1, SPQ-3.A.2, SPQ-3.B.1, SPQ-4.A.1, SPQ-4.A.3, TRA-1.C.1)
Molarity is the most important unit of concentration used in chemistry. Students need to have a strong grasp of the units of molarity (moles of solute/liters of solution) and be able to calculate molarity when given mass of solute and volume of solution. The meaning of the term millimolar (*mM*) should also be mentioned. Emphasize the use of molarity as a conversion factor in stoichiometry problems. Students also need to be able to correctly calculate and thoroughly describe how to prepare a solution of a given molarity starting with either a solid solute or a more concentrated solution. Students often have a great deal of difficulty describing the correct process in words when asked to do this on the AP® Chemistry Exam. It may be helpful to do an activity where students have to calculate and write directions for preparing solutions and then compare with a partner to make sure that the directions are correct and complete. Having students actually prepare their own laboratory solutions is also very beneficial to deepen understanding of the process.

Topic 3.7; Learning Objective SPQ–3.A: *Calculate the number of solute particles, volume, or molarity of solutions.*
Topic 3.8; Learning Objective SPQ–3.B: *Using particulate models for mixtures to (a) represent interactions between components and (b) represent concentrations of components.*

Section 4.6 Gravimetric Analysis (EK SPQ-1.A.1, SPQ-2.B.1, SPQ-2.B.2, SPQ-4.A.1, SPQ-4.A.3, TRA-1.A.2, TRA-2.A.5)
Gravimetric analysis is a lab procedure usually used to quantitatively determine the amount of a substance. It usually involves the formation, filtration, drying, and massing of a precipitate. Types of filtration (gravity and vacuum) should be demonstrated. College Board AP® Chemistry Guided-Inquiry Lab #3 would be appropriate here.

Topic 1.1; Learning Objective SPQ–1.A: *Calculate quantities of a substance or its relative number of particles using dimensional analysis and the mole concept.*

Topic 4.5; Learning Objective SPQ–4.A: *Explain changes in the amounts of reactants and products based on the balanced reaction equation for a chemical process.*

Section 4.7 Acid-Base Titration (EK SAP-9.E.1, SAP-9.E.2, SPQ-1.A.1, SPQ-3.A.2,SPQ-4.A.1, SPQ-4.A.3, SPQ-4.B.1)

Titration is introduced in this section. It is important that students understand that titration is a lab procedure and not a new type of mathematical problem. Proper titration techniques and common errors need to be discussed. Students need to practice titrations in several labs throughout the year. It may need to be clarified that the equivalence point and the end point in a titration are not necessarily the same. If an incorrect indicator is chosen, its end point will occur either before or after the equivalence point. Choosing the correct indicator will be addressed in a later chapter. College Board AP® Chemistry Guided-Inquiry Lab # 4 would be appropriate here.

* Introduce the Online Lab: Urinalysis.

> ***Topic 4.5; Learning Objective SPQ–4.A:*** *Explain changes in the amounts of reactants and products based on the balanced reaction equation for a chemical process.*
> ***Topic 4.6; Learning Objective SPQ–4.B:*** *Identify the equivalence point in a titration based on the amounts of the titrant and analyte, assuming the titration reaction goes to completion.*

Section 4.8 Redox Titrations (EK SPQ-1.A.1, SPQ-3.A.2, SPQ-4.A.1, SPQ-4.A.3, SPQ-4.B.1)

A redox titration is performed in the same manner as an acid-base titration except that it may be more difficult to balance the equation and different indicators must be used. This would be a good time to perform College Board AP® Chemistry Guided-Inquiry Lab # 8 or another redox titration.

> ***Topic 4.5; Learning Objective SPQ–4.A:*** *Explain changes in the amounts of reactants and products based on the balanced reaction equation for a chemical process.*
> ***Topic 4.6; Learning Objective SPQ–4.B:*** *Identify the equivalence point in a titration based on the amounts of the titrant and analyte, assuming the titration reaction goes to completion.*

Math Skills

- Calculation of number of ions in a formula unit
- Calculation of molarity
- Use of dilution equation
- Use of molarity in stoichiometry
- Assignment of oxidation numbers to elements in an ion or compound

Suggested Problems

Section 4.1: Problems 4.7, 4.8, 4. 9, 4.13
Section 4.2: 4.17, 4.18, 4.21, 4.23
Section 4.3: 4.29, 4.30, 4.31, 4.33
Section 4.4: 4.45, 4.47, 4.57
Section 4.5: 4.61, 4.63, 4.65, 4.67, 4.69, 4.73, 4.75, 4.77
Section 4.6: 4.81, 4.83, 4.84
Section 4.7: 4.87, 4.88, 4.89, 4.91
Section 4.8: 4.95, 4.97, 4.99, 4.102

Mixed Concepts: 4.105, 4.107, 4.111, 4.113, 4.115, 4.119, 4.123, 4.127, 4.130, 4.131, 4.137, 4.140, 4.1424.144, 4.165

Activities

1. **Metathesis Reactions (*Topic 4.2; Learning Objective TRA–1.B)***
   This activity is valuable for learning to make proper observations and to write net ionic equations. Because students are actually mixing the chemicals, they tend to learn more than when simply writing equations from a textbook. While the lab describes using 1 mL portions in test tubes, it can also be done in well plates or on plastic report covers with only a very few drops. If you do not have some of these chemicals, simply omit those equations. This would also be a good opportunity to have students practice making the solutions for the class. Assign each group several solutions to prepare. Check their calculations before they start.

   Prepare approximately 100 mL of each of the following solutions:

   0.1 *M* sodium chromate
   0.1 *M* lead(II) nitrate
   0.05 *M* silver nitrate
   0.1 *M* sodium acetate
   0.1 *M* nickel(II) chloride
   0.1 *M* potassium chloride
   0.1 *M* sodium carbonate
   0.1 *M* copper(II) sulfate
   0.1 *M* sodium sulfide
   0.2 *M* barium chloride
   0.1 *M* sodium hydroxide
   0.1 *M* sodium phosphate
   0.1 *M* ammonium chloride
   1.0 *M* sulfuric acid
   0.1 *M* sodium nitrate
   1.0 *M* hydrochloric acid
   0.1 *M* potassium iodide

Procedure:

19 pairs of chemicals are to be mixed. Use about 1 mL (20 drops) of each of the reagents. Mix the solutions in small test tubes and record your observations. If no reaction is observed (look for precipitates and gases), write NVR. If you suspect an odor, but do not detect it when wafting the gas, pour the solution out into a watch glass and try again. Mix the following combinations:

1. sodium carbonate and hydrochloric acid
2. sodium carbonate and sulfuric acid
3. sodium carbonate and nickel(II) chloride
4. sodium carbonate and copper(II) sulfate
5. sodium nitrate and potassium chloride
6. copper(II) sulfate and barium chloride
7. copper(II) sulfate and sodium phosphate
8. silver nitrate and potassium iodide
9. silver nitrate and potassium chloride
10. silver nitrate and sodium chromate
11. lead(II) nitrate and potassium iodide
12. lead(II) nitrate and potassium chloride
13. lead(II) nitrate and sulfuric acid
14. lead(II) nitrate and sodium sulfide
15. nickel(II) chloride and silver nitrate
16. hydrochloric acid and sodium hydroxide
17. hydrochloric acid and sodium sulfide
18. ammonium chloride and sodium hydroxide
19. sodium acetate and hydrochloric acid

In your laboratory notebook, follow this pattern for each reaction:

Observations____

Molecular Equation____________________________________

Ionic
Equation____

Net ionic
Equation______________________________________

2. **Acid-Base Reactions (Topic 4.6; Learning Objective SPQ–4.B)**
   Simple titrations can be carried out using micro well plates or small test tubes and plastic beral pipettes as burettes. Students should be given dilute concentrations of a strong monoprotic acid and a strong base (HCl and NaOH work well) along with bromothymol blue indicator. The acid or base molarity should be given and they should be asked to calculate the molarity of the unknown. They can count drops as a means of measuring volume and the class data should be pooled to determine an average. The actual concentration of the unknown can be determined using an accurate pH meter, but you need not show them this calculation. Simply report the unknown concentration and have the students calculate the relative (%) error.

   The advantages to this process are that only small volumes of solutions are needed, many trials can be performed quickly, the process is very easy to learn, and students will observe the sudden change in color as one drop is added. Additionally, if the equivalence point is passed the students can be instructed to add several drops of the unknown to bring the solution back to a testable state giving them experience with “shifting equilibrium”. Even though the equipment is rudimentary, this experiment gives excellent results if the two solutions are fairly dilute (e.g. 0.01–0.09 *M*).

3. **Single Replacement Reactions (Topic 4.2; Learning Objective TRA–1.B)**
   Have students perform this activity after providing instruction for single replacement (redox) reactions. Prepare several 0.1 *M* solutions containing metallic cations of the transition elements such as Fe^{3+}, Zn^{2+}, Ag^{+}, Cu^{2+}, Ni^{2+}, and Mg^{2+} nitrates. Prepare small metallic strips or squares that correspond to these cations (iron, zinc, etc.). Give the solutions and the strips to the students and have them construct an activity table illustrating the tendency of an ion to remove an electron from a metal. This can be done in a 24-well plate or a spot plate with the solutions in dropper bottles or beral pipets. Only a few drops of each solution are needed. You may want to demonstrate or state that if a cation removes an electron(s) from the metal, a layer of the new metal will form on the metallic strip and the metal strip will form ions and go into solution. The example given in the book illustrates this:

   $$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$
   $$Cu(s) + Zn^{2+}(aq) \rightarrow \text{No reaction}$$

   Thus you can conclude that zinc metal is a ***more active*** metal than copper since it reacts in solution, while the copper metal does not. Consequently, zinc metal would be above copper metal on an activity table that lists reactivity from highest to lowest. The activity series should indicate the metals from most active to least active as follows: Mg, Zn, Fe, Ni, Cu, and Ag. Have students create their own data table.

4. **Gravimetric Analysis Activity (Topic 4.5; Learning Objective SPQ–4.A)**
   Have students dissolve about 1.5 g of carbonate in 20 mL of water in a 50 mL beaker, and 1.5 g of calcium chloride in 20 mL of distilled water in another 50 mL beaker. Once the solids are completely dissolved, have students combine the two solutions into one beaker, and filter out the precipitate that forms. Students should write out the reaction that just took place, balance it, and use solubility guidelines (or their knowledge of the ions that form compounds that are soluble in water) to figure out the compound that ends up being insoluble (the precipitate). Students should then use their balanced equation and the mass of their reactants to figure out the theoretical yield of calcium carbonate. Students should weigh a piece of filter paper, then filter their precipitate through vacuum filtration, let dry, weigh their precipitate (actual yield), and calculate their percent yield.

Study Strategies

Study Strategies are activities geared to help students struggling with the content or language demands of the AP Chemistry course.

Practice Linguistic Patterns - Beginning or Intermediate

Have students work in small groups. This strategy allows every student to have an opportunity to speak several times. Ask a question or give a prompt about the properties of acids and bases, and then pass a stick or other object to the student. The student speaks, everyone listens, and then passes the object to the next person. The next student speaks, everyone listens then the student passes the object on until everyone has had one or two turns.

Oral Language Development – Advanced

Have students scan the lesson for content vocabulary words in context. Help them pronounce the vocabulary words correctly. Discuss vocabulary meanings with them.

Additional AP Practice Questions

Multiple-Choice:

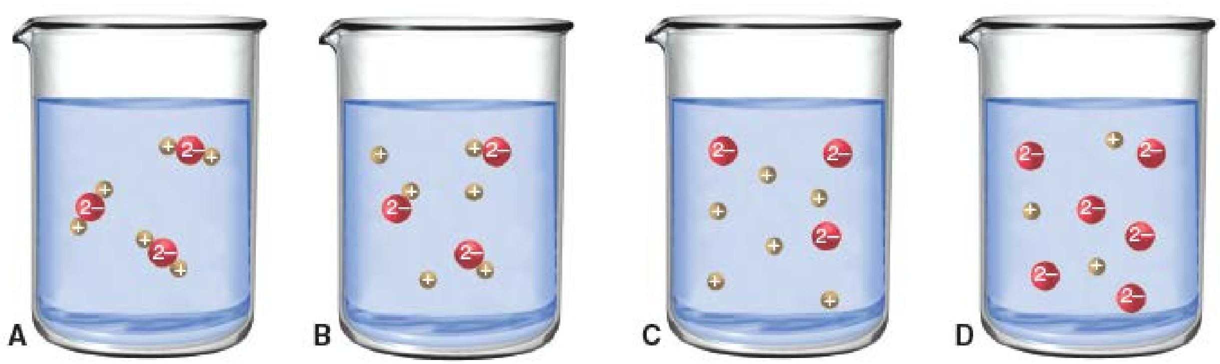
1. Sucrose ($C_{12}H_{22}O_{11}$) is dissolved in pure water and its conductivity is tested. Which of the following is the ***best*** summary of ***and*** correct explanation for the results of this test?

   (A) The solution conducts electricity because the sugar dissolves and dissolved substances conduct electricity.
   (B) The solution conducts electricity because water is a good conductor *and* dissolved substances conduct electricity.
   (C) The solution does not conduct electricity because the conductivity of water is neutralized by the dissolved sugar.
   (D) The solution does not conduct electricity because water is a poor conductor and the sugar does not dissociate.

2. Which of the following ion pairs will ***most likely*** form a precipitate when combined in aqueous solution?

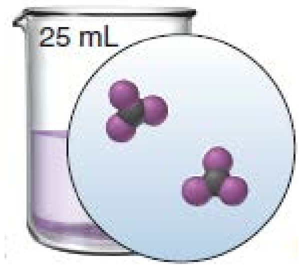
   (A) Na^{+} and Al^{3+}
   (B) Fe^{3+} and S^{2-}
   (C) Al^{3+} and NO_3^{-}
   (D) NO_3^{-} and S^{2-}

3. Using the particle views, which beaker best represents an aqueous solution of sodium sulfate? Water molecules are not shown.



4. Which of the following elements will react with KCl(*aq*)?

   (A) $F_2(g)$
   (B) $Cl_2(g)$
   (C) $I_2(s)$
   (D) $Br_2(l)$



5. The molecular view in the diagram above represents an aqueous solution of ammonia. If each molecule shown represents 0.010 mol of molecules, what is the concentration of the solution? The volume of the solution is shown at the top of the beaker in the diagram.

   (A) 2.0 *M*
   (B) 0.80 *M*
   (C) 0.02 *M*
   (D) 0.00080 *M*

6. What volume of 0.5 *M* H_2SO_4 will be required to exactly neutralize 250 mL of 1 *M* NaOH?

   (A) 125 mL
   (B) 250 mL
   (C) 300 mL
   (D) 500 mL

7. An excess of silver nitrate solution is mixed with a solution containing an unknown concentration of calcium chloride. A precipitation reaction occurs and silver chloride is formed. The precipitate is filtered, dried and its mass is determined. This technique is used to determine

   (A) The amount of silver ions added to the unknown solution
   (B) The concentration of the silver ions in the original solution
   (C) The concentration of the chloride ions in the original solution
   (D) Both the concentration of the silver ions and the chloride ions in the original solution

$$2KMnO_4 + 3H_2O_2 \rightarrow 2MnO_2 + 3O_2 + 2KOH + 2H_2O$$

8. In the above reaction, what has an oxidation state of +4?

(A) Mn in $KMnO_4$
(B) O in H_2O_2
(C) Mn in MnO_2
(D) O in O_2

9. The concentration of Ca^{2+} in impure tap water can be determined by the addition of Na_2CO_3 to a water sample to produce insoluble $CaCO_3$. Excess sodium carbonate solution is added to a 1.0 L sample of tap water and the precipitate that formed was filtered, dried, and had a mass of 0.1001 g. What was the concentration of calcium in the original water sample?

   (A) 1.0×10^{-1} *M*
   (B) 1.0×10^{-2} *M*
   (C) 1.0×10^{-3} *M*
   (D) 1.0×10^{-4} *M*

10. Which of the following reactions involves the transfer of electrons from one species to another?

   (A) $HCl + NaOH \rightarrow H_2O + NaCl$
   (B) $CaO + CO_2 \rightarrow CaCO_3$
   (C) $Mg + CuSO_4 \rightarrow Cu + MgSO_4$
   (D) $Pb(NO_3)_2 + 2HCl \rightarrow PbCl_2 + 2HNO_3$

Answers Additional AP Practice Questions

1(D), 2(B), 3(C), 4(A), 5(B), 6(B), 7(C), 8(C), 9(C), 10(C)

Chemistry in Action Feature Answers

An Undesirable Precipitation Reaction (SE, p. 129)

1. List the two materials, commonly found on Earth's surface, that are ultimately responsible for hard water
   Limestone (calcium carbonate) and dolomite (calcium carbonate combined with magnesium carbonate)
2. Explain why hard water is undesirable in household appliances.
   When heated, soluble calcium bicarbonate in hard water is converted to solid calcium carbonate precipitate which may collect in appliances and cause decreases in efficiency and durability of pipes and boilers.
3. What is the intermediate acid that is formed when hydrochloric acid is added to hard water that is the source of carbon dioxide gas?
   Carbonic acid forms when a strong acid interacts with carbonate ion but is unstable and decomposes to form carbon dioxide and water.

Breathalyzer (SE, p. 146)

1. Describe how electrons are transferred in the balanced redox reaction utilized by the breathalyzer.
   A total of 12 electrons are transferred from ethanol to chromium, three per chromium ion, to reduce chromium (VI) to chromium (III).
2. State the color change that is read by the breathalyzer detector.
   Dichromate ion is orange-yellow and chromium(III) is green. The degree of this color change is read to measure the concentration of chromium(III) and therefore the amount of ethanol input in the device.
3. What should be the limiting reagent in this process?
   In order to accurately determine its concentration, ethanol should be the limiting reagent and all other reagents should be present in excess.

Student Edition Chapter Review Answers

For the answers to the Questions & Problem section for Chapter 4 see the Answers to Even-Numbered Problems (SE, p. AP-2) & the Instructor's Solution Manual.

The following are the answers to the AP Chapter Review found on SE pages AP170-AP171.

Answers to Applying the Big Ideas Questions

Question 1

A particle dissolved in water is shown below.

Which of the following scenarios would best describe a situation where this particle would interact with water in this manner?

a. Glucose is dissolved to form a non-electrolytic solution
b. Ethanol is dissolved and forms an electrolytic solution
c. Barium hydroxide is dissolved and ionizes to form a cation and anion in an electrolytic solution
d. Sodium chloride is dissolved and ionizes to form a cation and anion in an electrolytic solution

Answer: D

Topic 3.8; Learning Objective SPQ–3.B

Question 2

A highly soluble ionic compound is composed of the ions A^{2+} and X^{3-} and reacts in a metastasis, or double replacement, reaction with another highly soluble compound composed of the ions Y^+ and Z^{2-}. If AZ is an insoluble precipitate, what is the balanced net ionic equation for this reaction?

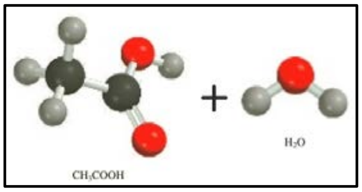
a. $A^{2+}(aq) + Z^{2-}(aq) \rightarrow AZ(s)$
b. $A_3(aq) + 3Z(aq) \rightarrow 3AZ(s)$
c. $A_3X_2(aq) + 3Y_2Z(aq) \rightarrow 2\ Y_3Z(aq) + 3AZ(s)$
d. $3A^{2+}(aq) + 2X^{3-}(aq) + 6Y^+ + 3Z^{2-}(aq) \rightarrow 6Y^+ + 3Z^{2-}(aq) + 3AZ(s)$

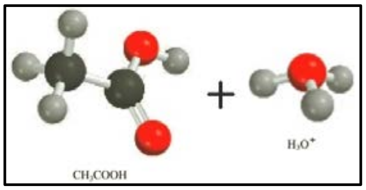
Answer: A

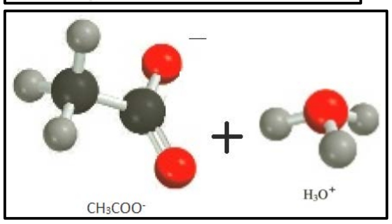
Topic 4.7; Learning Objective TRA–2.A

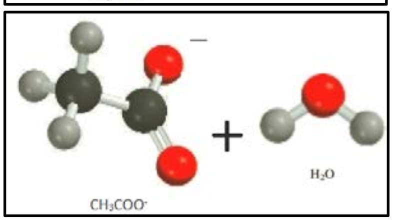
Question 3

Acetic acid, shown below, is a weak acid. Which of the following best illustrates the predominant species present in a solution of acetic acid?

a. 

b. 

c. 

d. 

Answer: A

Topic 8.3; Learning Objective SAP–9.C

Question 4

Which of the following balanced equations is matched correctly with the type of reaction?

a. $HI(aq) + KOH(aq) \rightarrow KI(aq) + H_2O(l)$ Precipitation
b. $Ca(s) + 2H_2O(aq) \rightarrow Ca(OH)_2(s) + H_2(g)$ Acid base neutralization
c. $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$ Oxidation Reduction
d. $2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(g)$ Acid base neutralization

Answer: C

Topic 4.7; Learning Objective TRA–2.A

Question 5

Lead is a useful metal for batteries, coatings and pigments and thus is involved in many chemical reactions. Examine the following and select the reaction that demonstrates the disproportionation of lead.

a. $2PbS(s) + 3O_2(g) \rightarrow 2PbO(s) + 2SO_2(g)$
b. $Ni(s) + Pb(NO_3)_2(aq) \rightarrow Pb(s) + Ni(NO_3)_2(aq)$
c. $2KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$
d. $2PbSO_4(s) + 2H_2O(l) \rightarrow Pb(s) + PbO_2(s) + 2H_2SO_4(aq)$

Answer: D

Topic 4.7; Learning Objective TRA–2.A

Question 6

A student has been assigned the task of mixing four solutions for his lab. Which of the solutions that he made would have the greatest number of dissolved iodide anions?

a. A 1.0 M solution of potassium iodide
b. A 1.0 M solution of lead(II) iodide
c. A 0.5 M solution of aluminum iodide
d. A 0.1 M solution of lithium iodide

Answer: C

Topic: Cannot assign one.

Answers to Free Response Questions

Question 7

A student in Flint, Michigan is asked to find the concentration of lead(II) ions in 4.0 L of a tap water sample and decides to use a precipitation reaction with iodide ions and gravimetric analysis to do it. Knowing that his accuracy will be increased by doing three trials, he divides his assigned solution into four equal aliquots by pouring 1.0 L of the solution in four labeled graduated cylinders and then tests the first three. The student's data is shown below.

Aliquot	Volume of 2.0M HCl solution used (mL)	Mass of filter paper (g)	Mass of filter paper with dried precipitate (g)
1	100.0	0.090	0.119
2	100.0	0.101	0.124
3	100.0	0.092	0.123

a. Write a balanced, net ionic equation for this reaction.
b. What was the mass of the precipitate for each aliquot? What was the average mass of the precipitates from the three aliquots?
c. Calculate the moles of lead(II) ions from the three averaged trials.
d. What was the concentration of the lead solution used in the experiment?
e. If the actual concentration of the lead solution was 1.82×10^{-5}M, what was the student's

percent error?

f. The student measured four aliquots from the assigned solution, but only tested the first three he poured out. Why would he choose to not test the fourth aliquot?

g. Hydrochloric acid was an excess reagent in this experiment. Why was such an excess of this reactant needed for this experiment?

Answer

a. Write a balanced, net ionic equation for this reaction.

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

b. What was the mass of the precipitate for each aliquot? What was the average mass of the precipitates from the three aliquots?

1 = 0.029g, 2 = 0.023g, 3 = 0.031g   Average mass = 0.028g

c. Calculate the moles of lead(II) ions from the three averaged trials.

6.1×10^{-5} mol Pb^{2+} ions

d. What was the concentration of the lead solution used in the experiment?

2.0×10^{-5} M Pb^{2+}

e. If the actual concentration of the lead solution was 1.82×10^{-5}M, what was the student's percent error?

10.%

f. The student measured four aliquots from the assigned solution, but only tested the first three he poured out. Why would he choose to not test the fourth aliquot?

The fourth aliquot would not be exactly 1.0 L since some small volume is lost in the process of pouring/measuring.

g. Hydrochloric acid was an excess reagent in this experiment. Why was such an excess of this reactant needed for this experiment?

In order to ensure that all of the lead ions precipitated, excess HCl was used.

Topics 1.1, 4.2, 4.5 and 4.7; Learning Objectives SPQ–1.A, TRA–1.B, SPQ–4.A and TRA–2.A

Question 8

Coconut water has been a recent addition to the beverage market as a quick way to rehydrate and replace needed electrolytes after an intense workout. A high school cross country runner and budding chemist, Callie, has read about the drink and wishes to find out how acidic it is. She performs a titration using the coconut water as the analyte and 0.010M NaOH as the titrant. Her data is listed below.

Trial	Volume of coconut water tested (mL)	Initial Buret Volume (mL)	Final Buret Volume, measured at the equivalence point (mL)
1	100.2	0.12	3.47
2	99.7	0.04	3.19
3	101.6	0.02	3.22

a. The primary acid in coconut water is ascorbic acid, $C_4H_7NO_4$, a weak acid that is monoprotic. Write a dissociation equation for the dissolving of ascorbic acid.
b. Write the net ionic equation for the reaction that occurred when Callie titrated the coconut water with the sodium hydroxide titrant.
c. Using the data in trial 2, find the concentration of the hydrogen ions in the sample of coconut water.
d. In a later trial, Callie decided to test the concentration of another of her favorite athletic drinks, but finds that it requires a 1.5 M solution of sodium hydroxide. Describe below how she should make 100 mL of this solution if she already has a stock solution of sodium hydroxide that is 5.0 M. Be sure to include descriptions of appropriate glassware and safety techniques.
e. Another student in class becomes interested in Callie's experiment, but wants to test the electrolytic properties of coconut water. Based on what you know about it, would you expect this to be a strong, weak or non-electrolyte? Explain your reasoning.

Answer

a. The primary acid in coconut water is ascorbic acid, $C_4H_7NO_4$, a weak acid that is monoprotic. Write a dissociation equation for the dissolving of ascorbic acid.

$C_4H_7NO_4(aq) \rightarrow C_4H_6NO_4^-(aq) + H^+(aq)$

b. Write the net ionic equation for the reaction that occurred when Callie titrated the coconut water with the sodium hydroxide titrant.

$C_4H_7NO_4(aq) + OH^-(aq) \rightarrow C_4H_6NO_4^-(aq) + H_2O(l)$

c. Using the data in trial 2, find the concentration of the hydrogen ions in the sample of coconut water.

3.2×10^{-4} M H^+

d. In a later trial, Callie decided to test the concentration of another of her favorite athletic drinks, but finds that it requires a 1.5 M solution of sodium hydroxide. Describe below how she should make 100 mL of this solution if she already has a stock solution of sodium hydroxide that is 5.0 M. Be sure to include descriptions of appropriate glassware and safety techniques.

Use $M_1V_1 = M_2V_2$ to find that she should measure out 30. mL of the 5.0 M NaOH in a graduated cylinder. The measured NaOH should then be added to a 100 mL volumetric flask, filled with some distilled water, swirled to mix, then filled to the etched line with distilled water and labeled.

e. Another student in class becomes interested in Callie's experiment, but wants to test the electrolytic properties of coconut water. Based on what you know about it, would you expect this to be a strong, weak or non-electrolyte? Explain your reasoning.

Coconut water should have strong electrolytic properties because it is used to replace lost electrolytes from athletes. It would definitely have some electrolytic properties because of the presence of the weak ascorbic acid.

Topics 3.7, 4.5, 4.6 and 4.7 and 8.3; Learning Objectives SPQ–3.A, SPQ–4.A, SPQ–4.B, TRA–2.A and SAP–9.C

Answers to Applying the Science Practices Questions

Think Critically Questions
Question 1: Use the data in the table to make a graph of atmospheric pressure versus altitude.
Answer: Accept student line graphs with correctly labeled axes and a line showing a negative slope.
Question 2: Calculate your actual diving depth if your depth gauge reads 18 m, but you are at an altitude of 1800 m and your gauge does not compensate for altitude.
Answer: 20 m
Question 2: Dive tables are used to determine how long it is safe for a diver to stay under water at a specific depth. **Infer** why it is important to know the correct depth of a dive.
Answer: The amount of time that it is safe to stay under water is directly related to the diving depth. If you do not know your actual diving depth, you cannot determine how long it is safe to stay at a particular diving depth.