Part II

A Review of AP Chemistry

NON-TESTABLE CHEMISTRY CONCEPTS

The following topics, which have long been part of the AP Chemistry curriculum, have been eliminated because they are comprised of knowledge you are expected to have acquired in a prior class. You will not be assessed specifically on this content, but knowing the information may make it easier to answer some exam questions.

- Memorizing exceptions to the Aufbau principle, but you should be able to explain why a specific configuration is an exception
- Interpreting phase diagrams
- Knowledge of colligative properties
- Calculation of molality
- Knowing the definition of Lewis acids, but you are still responsible for understanding the formation of complex ions and their qualitative effect on solubility

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CHEMICAL FOUNDATIONS

In this chapter, you will perform the basic calculations used in chemistry. Chemists use equipment with varying degrees of precision. The measurements recorded in an experiment must reflect the precision of the equipment used. The results of the calculations in an experiment also must reflect the precision of the equipment.

Many of the problems in chemistry use dimensional analysis to convert from one unit to another or to solve stoichiometry problems. Other basic calculations that will be discussed are density and temperature conversions.

You should be able to

- Identify the number of significant figures in a given measurement.
- Perform calculations involving significant figures.
  - Memorize the rules for counting and performing operations with significant figures.
- Differentiate between accuracy and precision as they apply to measurement.
- Determine the density of solids and liquids and calculate volumes or masses using the given density.
- Convert between units of temperature: degrees Celsius and Kelvin.
- Identify the characteristics of the states of matter: solids, liquids, and gases.
- Identify changes as being physical or chemical.
- Name compounds and write formulas for binary compounds, ternary compounds (those with polyatomic ions), and acids.
- Memorize the chemical formulas and charges of the polyatomic ions and the most common transition metal ions.

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AP Tips

Significant figures are important in calculations. A maximum of one point in the free-response sections will be deducted for errors in significant figures that are off by ±1 significant figure.

Units are the key to problem solving. Include units in your calculated work to ensure that your answer has units that agree with the problem. Every measured number or number calculated from measurements has a unit except for equilibrium constants, which by convention are reported without units. When a calculation specifically asks for units to be included, you can be assured that the lack of units or incorrect units will result in a point deduction.

UNCERTAINTY IN MEASUREMENT

(*Chemistry* 8th ed. pages 11–14/9th ed. pages 11–14)

SIGNIFICANT FIGURES

The significant figures of a measurement are all of the certain digits in a measurement, those that represent actual marks on the measuring instrument, and the first uncertain digit (estimated number). Students should be able to read measurements to the proper number of significant figures.

EXAMPLE: Figure 1.7 on page 11 in the 8th edition and Figure 1.7 on page 12 in the 9th edition shows the measurement of a volume of liquid using a buret. The *certain digits* in the measurement are the three digits 20.1. The digit to the right of the one must be estimated by interpolating between the 0.1-mL marks. The measurement with *uncertainty* can be reported as 20.15 mL.

PRECISION AND ACCURACY

Accuracy refers to the agreement of a particular value with a true value.

Precision refers to the degree of agreement among several measurements of the same quantity. The degree of precision refers to the number of digits that a measuring device permits one to measure. In a measuring device, all except the last digit, which is estimated, are certain. For example, a balance which measures to the nearest 0.0001 g is more precise than one that measures to the nearest 0.01 g.

$$\text{Percent Error} = \frac{\text{Experimental Value} - \text{Actual Value}}{\text{Actual Value}} \times 100\%$$

EXAMPLE: In an experiment, the density of aluminum is to be determined. Two students perform the experiment three times and obtain the following results.

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Trial	Student A	Student B
1	2.45 g/mL	2.69 g/mL
2	2.43 g/mL	2.70 g/mL
3	2.44 g/mL	2.71 g/mL

Describe the accuracy and precision of each student's results. For Student A, calculate the average density and the percent error if the actual value is 2.70 g/mL.

SOLUTION: Student A's average density is 2.44 g/mL while Student B's average density is 2.70 g/mL. The values for Student A are precise, but inaccurate. The values are precise because they are consistent, but inaccurate because they are not close to the accepted value. The percent error for Student A is

$$\text{Percent Error} = \frac{\text{Experimental Value} - \text{Actual Value}}{\text{Actual Value}} \times 100\%$$

$$= \frac{2.44\ \text{g/mL} - 2.70\ \text{g/mL}}{2.70\ \text{g / mL}} \times 100\%$$

$$= \frac{0.26\ \text{g/mL}}{2.70\ \text{g/mL}} \times 100\%$$

$$= -9.6\%$$

Significant figure rules are described below.

The values for Student B are both precise and accurate.

Random error (indeterminate error) means that a measurement has an equal probability of being high or low.

Systematic error (determinate error) occurs in the same direction each time; it is either always high or always low.

EXAMPLE: A balance could have a defect causing it to give a result that is consistently 1.000 g too high. What type of error is this?

SOLUTION: A consistently high error is a systematic error. This type of error could be due to miscalibration.

SIGNIFICANT FIGURES AND CALCULATIONS

It is important to know the uncertainty in the final result in an experiment. The final reported result cannot have more certainty than the least precise measurement. The number of significant figures in a single value will be determined. Memorize the rules below and use them to answer the examples that follow.

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Rules for Counting Significant Figures

1. Nonzero integers always count as significant figures.
2. Zeros: There are three classes of zeros.
   a. Leading zeros precede all the nonzero digits and do not count as significant figures. Example: 0.0025 has 2 significant figures.
   b. Captive zeros are zeros between nonzero numbers. These always count as significant figures. Example: 1.008 has 4 significant figures.
   c. Trailing zeros are zeros at the right end of the number.

      Trailing zeros are only significant if the number contains a decimal point. Example: 1.00×10^2 has 3 significant figures.

      Trailing zeros are not significant if the number does not contain a decimal point. Example: 100 has 1 significant figure.
3. Exact numbers, which can arise from counting or definitions such as 1000 mm = 1 m never limit the number of significant figures in a calculation.

EXAMPLE: How many significant figures are in each of the following?

100 L

SOLUTION: There is 1 significant figure. Trailing zeros do not count. Only the "1" is significant.

0.001010 L

SOLUTION: There are 4 significant figures. Leading zeros do not count. The numbers "1010" are significant because one zero is captive and the other zero is trailing, but a decimal is present.

Rules for Significant Figures in Calculations

1. For multiplication and division, the number of significant figures in the result is the same as the measurement with the fewest number of significant figures in the calculation.
2. For addition and subtraction, the result has the same number of decimal places as the measurement with the fewest number of decimal places in the calculation.
3. Rules for rounding:

   In a series of calculations, carry the extra digits to the final result, then round.

   If the digit to be removed

   is less than 5, the preceding digit stays the same. For example, 2.44 rounds to 2.4.

   is greater than or equal to 5, the preceding digit is increased by 1. For example, 2.45 rounds to 2.5.

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It is important to calculate the results of mathematical expressions to the proper number of significant figures. Memorize the rules above and apply them to the examples that follow.

EXAMPLE: Perform the following calculations to the correct number of significant figures.

1) 16.8 g + 3.2557 g

SOLUTION: The calculator answer is 20.0557. The correct answer is 20.1 g. The answer should have one decimal place.

2) 27 g/4.148 mL

SOLUTION: The calculator answer is 6.509161041. The correct answer is 6.5 g/mL. The answer should have 2 significant figures.

DIMENSIONAL ANALYSIS

(*Chemistry* 8th ed. pages 17–20/9th ed. pages 18–22)

Dimensional analysis is used to convert from one unit to another. It is the single most valuable mathematical technique that you will use in general chemistry. The method involves conversion factors to cancel units until you have the proper unit in the proper place. When you are setting up problems using dimensional analysis, you should be more concerned with units than numbers.

EXAMPLE: The density of mercury is 13.6 g/cm³. How many liters is 1500 kg?

SOLUTION:

$$L = 1500\text{ kg} \times \left(\frac{1000\text{ g}}{1\text{ kg}}\right) \times \left(\frac{1\text{ cm}^3}{13.6\text{ g}}\right) \times \left(\frac{1\text{ mL}}{1\text{ cm}^3}\right) \times \left(\frac{1\text{ L}}{1000\text{ mL}}\right)$$

$$= 110\text{ L}$$

Double check that all of your units cancel properly. If they do, your numerical answer is probably correct. If they don't, your answer is definitely wrong.

DENSITY

(*Chemistry* 8th ed. pages 24–26/9th ed. pages 26–27)

Density is the mass of substance per unit volume of substance.

Density = mass/volume

D = m/V

The density of regular shaped objects can be determined by using equations for determining volume such as $l \times w \times h$ for a cube or $\pi r^2 h$ for a cylinder. The density of an object can be determined through the water *displacement method*. The object is massed and then submerged in a measured amount of water in a graduated cylinder. The final volume in the graduated cylinder is read. The volume of water displaced by the object is the volume of the object.

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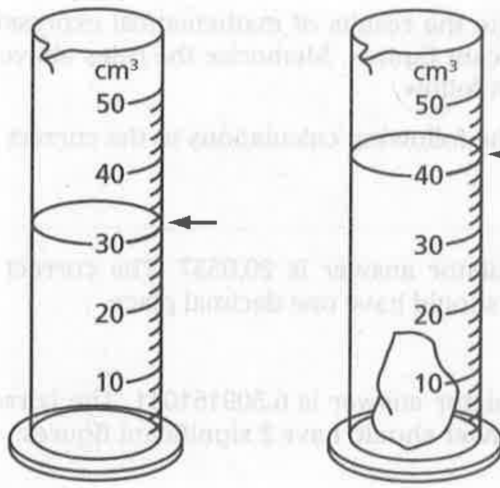

EXAMPLE: A sample containing 33.42 g of metal pellets is poured into a graduated cylinder containing 12.7 mL of water, causing the water level in the cylinder to rise to 21.6 mL. Calculate the density of the metal.

SOLUTION:

Volume of metal = (Volume H_2O metal) – Volume H_2O

= 21.6 mL – 12.7 mL = 8.9 mL

$$\text{Density of metal} = \frac{33.42\text{ g}}{8.9\text{ mL}} = 3.8\text{ g/mL}$$

Note: The answer has 2 significant figures.

TEMPERATURE

(*Chemistry* 8th ed. pages 21–24/9th ed. pages 22–26)

You should be able to interconvert between Celsius and Kelvin. You should also know the freezing and boiling points of water on each scale.

$$T_K = T_C + 273.15$$

EXAMPLE: The boiling point of water on top of Long's Peak in Colorado (14,255 feet above sea level) is about 86.0°C. What is the boiling point in Kelvin?

SOLUTION:

$T_K = T_C + 273.15 = 86.0°C + 273.15 = 359.2$ K (one place after the decimal according to significant figure rules)

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CLASSIFICATION OF MATTER

(*Chemistry* 8th ed. pages 26–29/9th ed. pages 27–31)

Matter exists in three states: solid, liquid, and gas. Properties of these three states of matter are listed below.

State of Matter	Shape	Volume
Solid	Fixed	Fixed
Liquid	Not definite	Fixed
Gas	Not fixed; takes shape of container	Not fixed; takes volume of container

CHANGES IN MATTER

Matter can undergo physical or chemical changes.

Physical changes in matter do not change the original composition of the substance. Changes in state such as boiling or melting are physical changes. Changes involving an alteration in the form of the substance such as grinding or tearing are physical. Physical properties are properties of a substance that can be observed without changing the composition of the substance. During a physical change, intramolecular bonds are not broken and no reaction between atoms occurs. For example, density, color, and boiling point are physical properties.

Chemical changes in matter change the composition of the original substance by breaking and making bonds between atoms. A new substance is produced when a chemical change occurs. Evidence that a chemical change has occurred includes change in color or odor; the production of a gas or a solid (precipitate); and the absorption or release of energy as seen by a temperature change or light given off. Some examples of chemical properties include flammability and reactivity to air.

COMPOUNDS

Elements combine in various ways to form compounds. Chemists generally divide compounds into two classes: ionic (formed by the transfer of electrons) and covalent (formed by the sharing of electrons) and each class has a unique nomenclature system that is used.

IONS

(*Chemistry* 8th ed. pages 52–53, 56–57, 61–62/9th ed. pages 55–57, 57–59, 65–66)

An ion is an atom that has lost or gained electrons and therefore has a positive or negative charge due to an imbalance of protons and electrons. A polyatomic ion is a group of atoms bonded together as a single unit that carries a net charge.

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EXAMPLE: Aluminum forms a cation, a positive ion, by losing three electrons:

$Al \rightarrow Al^{3+} + 3\ e^-$

Oxygen forms an anion, a negative ion, by gaining two electrons:

$O + 2\ e^- \rightarrow O^{2-}$

An ammonium cation is formed when ammonia, a neutral compound, gains a hydrogen ion.

$NH_3 + H^+ \rightarrow NH_4^+$

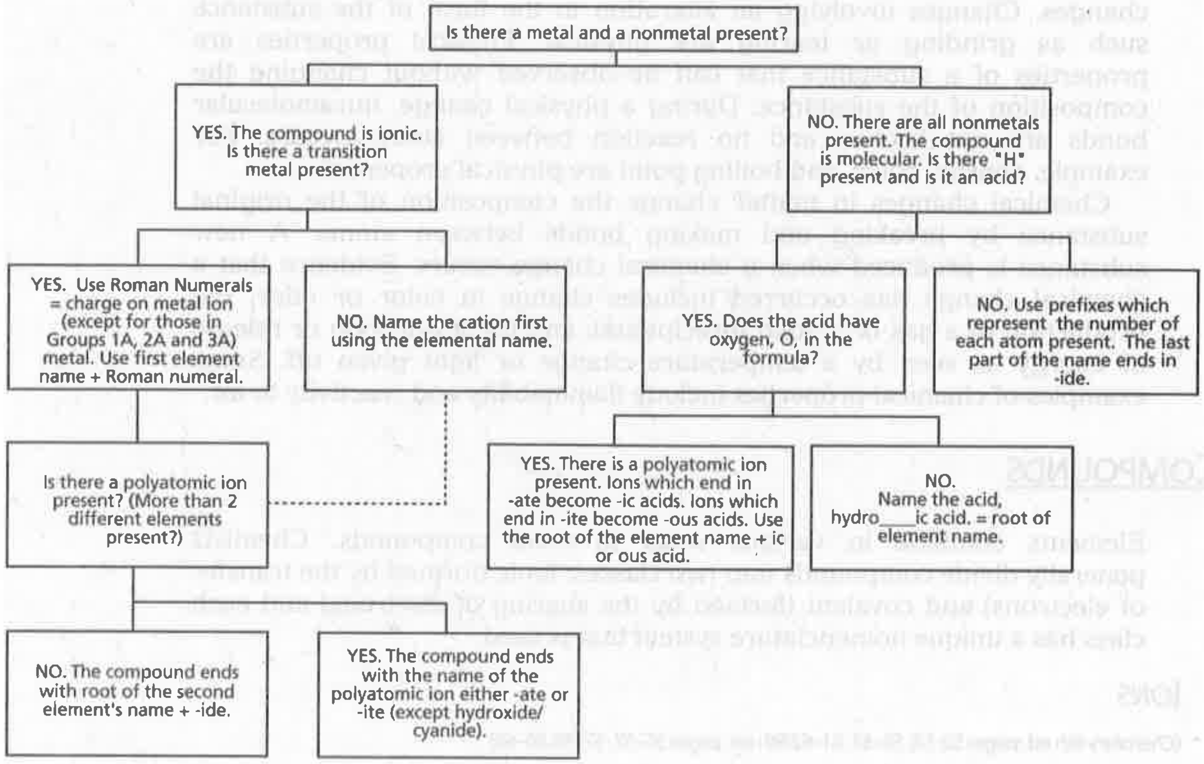
NOMENCLATURE

(*Chemistry* 8th ed. pages 56–67/9th ed. pages 60–70)

When you finish this section you will be able to name, or give formulas for, the following classes of compounds: binary salts, salts with polyatomic ions, binary covalent compounds and acids.

Use the flowcharts that follow to answer the questions below.

Naming Compounds from the Given Formula



EXAMPLE:

Name each of the following compounds.

a. CaF_2

b. Cl_2O_7

c. CuO

d. KIO_3

e. $HF(aq)$

f. $HNO_2(aq)$

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SOLUTION:

a. calcium fluoride

b. dichlorine heptoxide

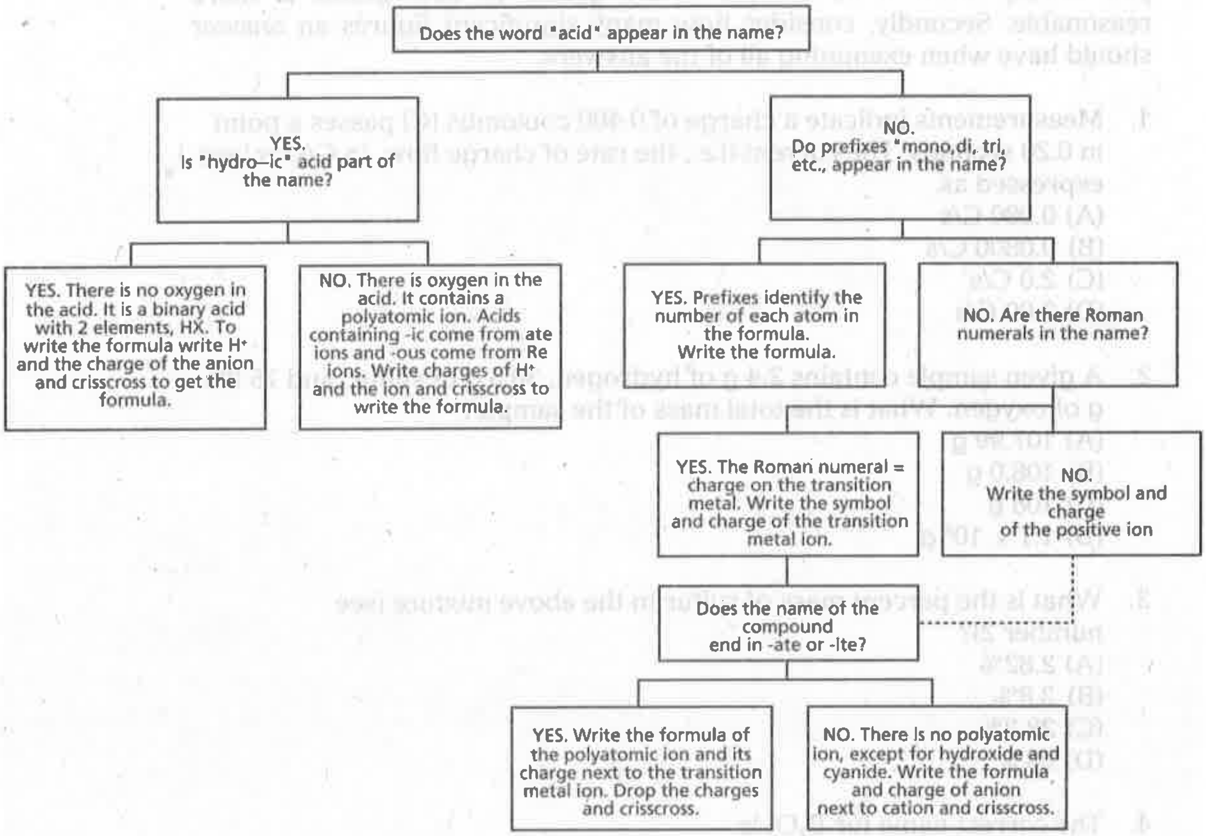
c. copper(II) oxide

d. potassium iodate

e. hydrofluoric acid

f. nitrous acid

Writing Formulas from the Given Name



EXAMPLE:

Write the formulas for each of the following compounds.

a. tin(IV) oxide

b. hydrosulfuric acid

c. nickel(II) nitrate

d. sulfurous acid

e. potassium chloride

f. iodine pentafluoride

SOLUTION:

a. SnO_2

b. $H_2S(aq)$

c. $Ni(NO_3)_2$

d. $H_2SO_3(aq)$

e. KCl

f. IF_5

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MULTIPLE-CHOICE QUESTIONS

Since this chapter is a review of chemistry, most of the questions below are not at the level seen on the AP Chemistry exam. No calculators are to be used in this section.

When choosing your answers, it is helpful to make the general observation that numerical answers tend to fall into one of two areas. Quite often the answers will vary greatly from each other. For example, you usually know if an answer of 3.0 grams or 300. grams is more reasonable. Secondly, consider how many significant figures an answer should have when examining all of the answers.

1. Measurements indicate a charge of 0.400 coulombs (C) passes a point in 0.20 seconds. The current (i.e., the rate of charge flow, in C/s) is best expressed as
(A) 0.080 C/s
(B) 0.0800 C/s
(C) 2.0 C/s
(D) 2.00 C/s

2. A given sample contains 2.4 g of hydrogen, 30.5 g of sulfur, and 75.09 g of oxygen. What is the total mass of the sample?
(A) 107.99 g
(B) 108.0 g
(C) 108 g
(D) 1.1×10^2 g

3. What is the percent mass of sulfur in the above mixture (see number 2)?
(A) 2.82%
(B) 2.8%
(C) 28.2%
(D) 28%

4. The correct name for B_2O_3 is
(A) boron oxide
(B) diboron trioxide
(C) boron(II) oxide(III)
(D) beryllium oxide

5. The correct name for $Mg(OH)_2$ is
(A) magnesium hydroxide
(B) magnesium(I) hydroxide
(C) magnesium(II) hydroxide(I)
(D) magnesium hydrogen oxide

6. The correct formula for iodine monobromide is
(A) $I_2(Br_2)_2$
(B) IBr
(C) I_2Br
(D) I_2Br_2

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7. The correct formula for copper(II) phosphate is
(A) $CuPO_4$
(B) Cu_3P_2
(C) $Cu_3(PO_4)_2$
(D) Cu_2PO_4

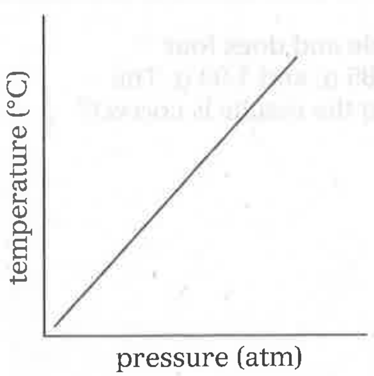
8. A student was trying to find the density of an unknown substance by measuring the mass of specific volumes of liquid. She used 3 different volumetric pipets and an electronic balance to obtain the data below. The actual density is 1.50 g/mL. Which of the following would account for her results?

Volume (mL)	**Mass (g)**	**Density (g/mL)**
10.00	16.091	1.609
20.00	32.106	1.605
15.00	24.072	1.605

(A) Random error; the pipets were poor quality.
(B) Random error; the balance was dirty.
(C) Systematic error; the pipets delivered too much volume.
(D) Systematic error; the mass indicated on the balance was too large.

9. The density of copper is approximately 9.0 g/cm³. What is the mass of 20.0 cm³ of pure copper?
(A) 2.2 g
(B) 200 g
(C) 180. g
(D) 180 g

10. An experiment that tested the effect of temperature on pressure was performed and data indicated a direct relationship between pressure and temperature. The student forgot to convert his temperatures from Celsius to Kelvin when he graphed temperature vs. pressure, however. How would this affect his results?



(A) The slope of the line is 273 times too large.
(B) The slope of the line is 273 times too small.
(C) The y-intercept of the line is 273 units too high.
(D) The y-intercept of the line is 273 units too low.

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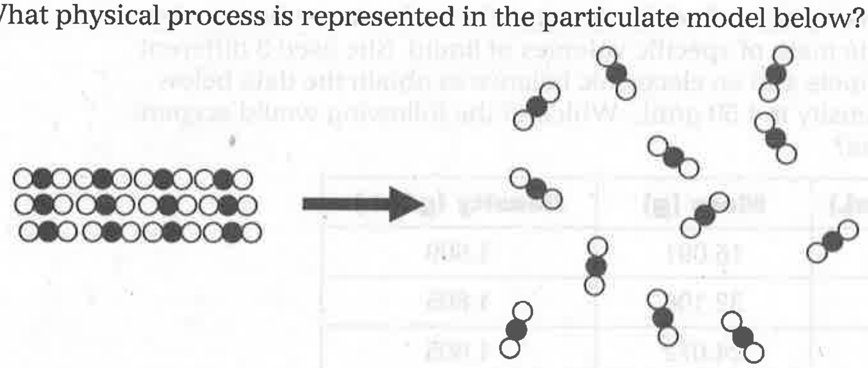
11. Which of the following processes represents a chemical change?
(A) water boiling
(B) iodine subliming
(C) sugar dissolving in water
(D) natural gas burning

12. What physical process is represented in the particulate model below?



(A) $CO_2(s) \rightarrow CO_2(l)$
(B) $CO_2(s) \rightarrow CO_2(g)$
(C) $CO_2(l) \rightarrow CO_2(g)$
(D) $(CO_2)_{12} \rightarrow 12\ CO_2$

13. What is the formula of hydrosulfuric acid?
(A) HS
(B) H_2S
(C) H_2SO_3
(D) H_2SO_4

14. A pure solid is heated and it decomposes into two substances, one a liquid and the other a gas. One can conclude with certainty that:
(A) the two products are elements
(B) one of the products is an element
(C) the original solid is not an element
(D) both products are compounds

15. A student measures the mass of silver in a sample and does four determinations. The results are 1.75 g, 1.71 g, 1.85 g, and 1.93 g. The true value is 1.81 g. Which statement concerning the results is correct?
(A) high precision and accurate results
(B) high precision and poor accuracy
(C) poor precision and poor accuracy
(D) poor precision and accurate results

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FREE-RESPONSE QUESTIONS

1. (a) Describe how, in the laboratory, you might determine experimentally the density of a solid, such as a sugar cube, which is water soluble. Indicate what equipment you might best use in the process.
   (b) Then describe a second experimental method for determining the density of this same object, so that you might verify the results of the first method.

2. Standard deviation is a measure of reproducibility of data. Low standard deviation means the data are precise. Percent error measures accuracy. Low percent error means accurate data. Use the data below to make a claim about why appropriate graduated cylinders should be used for accurate and precise volume measurements. Support your claim with evidence.

Glassware Used	Approximate Volume Measured (mL)	% Error	Standard Deviation
10-mL graduated cylinder	9	−1.80	0.0110
100-mL graduated cylinder	45	−2.43	0.0347
1000-mL graduated cylinder	45	−19.09	0.1357
50-mL Erlenmeyer flask	45	−3.83	0.0306
50-mL beaker	45	−4.13	0.0754
100-mL beaker	45	−7.12	0.0250
2000-mL beaker	50	−12.62	0.1921

Answers

MULTIPLE-CHOICE QUESTIONS

1. C Even if you do not understand what is meant by a term (in this case, current) you can often successfully attack a problem by careful attention to units. In the problem, you are seeking an answer in C/s, so divide the 0.400 C by 0.20 s (and keep your answer to 2 significant figures) (*Chemistry* 8th ed. pages 14–15/9th ed. pages 14–15).

2. B Each of these values is known to one-tenth of a gram (0.1 g) or more, so the total may be known to no more than the nearest 0.1 g.

   2.4 g + 30.5 g + 75.09 g = 107.99 g = 108.0 g

   Remember, when adding or subtracting, you do not count the number of significant figures but instead look at the position of the decimal point in each of the data used in the calculation (*Chemistry* 8th ed. pages 14–15/9th ed. pages 14–15).

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3. C Remember that no calculators are allowed on the multiple-choice portion of the AP exam. You should be able to do "mental math" to solve the problem.

$$\frac{30.5\text{ g}}{108.0\text{ g}} \times 100 \approx \frac{30}{100} \times 100 \approx 30\%$$

Now you are dividing, so the answer is known to the number of significant figures which is the least number of significant figures in data used to obtain the answer (here, 3 significant figures in the mass of sulfur limits the answer to 3 significant figures even though you know the total mass to 4). Also note that the 100% is known by definition, that is, it is considered an "exact number" (*Chemistry* 8th ed. pages 14–15/9th ed. pages 14–15).

4. B Prefixes are used when naming covalent (nonmetal–nonmetal) compounds (*Chemistry* 8th ed. pages 63–64/9th ed. pages 66–67).

5. A Roman numerals are used with transition metal compounds to indicate the charge or oxidation number of the metal ion in this compound. Since transition metal ions can often have more than one charge (oxidation number), this identification is necessary. There is no need for such nomenclature when the magnesium ion used for it can have only a 2+ charge. Note also that even the transition metal ions Ag^+, Zn^{2+}, and Cd^{2+} generally are found with only the oxidation number indicated, and since there is only one form of ion, Roman numerals are not used with these three transition metal ions (*Chemistry* 8th ed. pages 56–61, esp. Table 2.4/9th ed. pages 59–64, esp. Table 2.4).

6. B As indicated in the discussion for question #4, prefixes in covalently bonded compounds are used to correctly write the formula (*Chemistry* 8th ed. pages 63–64/9th ed. pages 66–67).

7. C Both copper with a 2+ and a 1+ charge are common ions. It is necessary to assign copper the 2+ charge in this case and then determine the ratio of Cu^{2+} and PO_4^{3-} to make the unit neutral; hence $Cu_3(PO_4)_2$ (*Chemistry* 8th ed. pages 61–62, esp. Table 2.5/9th ed. pages 64–65, esp. Table 2.5).

8. D The calculated densities were all precise but too large. This means that systematic error was present. It is very unlikely that 3 different volumetric pipets would have the same error so the error is most likely in the balance (*Chemistry* 8th ed. pages 12–14/9th ed. pages 12–14).

9. D From density = *m*/*v*, solving for mass and using mass = (*d*) × (*v*) = 9.0 g/cm^3 × 20.0 cm^3. The answer can only have 2 significant figures, 180 g (*Chemistry* 8th ed. pages 24–26/9th ed. pages 26–27).

10. D Since the relationship between Celsius and Kelvin is K = °C + 273.15, graphing the temperature in Celsius will cause the intercept to be –273°C instead of 0 K (*Chemistry* 8th ed. pages 21–23/9th ed. pages 22–25).

11. D When natural gas burns, at least two new substances are produced, CO_2 and H_2O. Choices (A), (B), and (E) are all examples of phase

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changes or changes of state, which are physical changes. Choice (C) is an example of a mixture, which can be separated by physical means back into the original substances (*Chemistry* 8th ed. pages 25–29/9th ed. pages 27–31).

12. **B** The model represents solid carbon dioxide subliming into gaseous carbon dioxide (*Chemistry* 8th ed. page 27/9th ed. page 29).

13. **B** Acids that are made from nonoxygen-containing anions are named as *hydro- -ic* (*Chemistry* 8th ed. pages 66–67, and Appendix 6/9th ed. pages 69–70).

14. **C** The original solid is pure and therefore it must be a compound because it can be chemically decomposed into simpler substances. You are not given specific information about the composition of the products and thus you can not make any definite conclusions about them (*Chemistry* 8th ed. pages 28–29/9th ed. pages 30–31).

15. **D** The average of the measurements is 1.81 g, which is the same as the true value. Therefore, the accuracy is high. The range of measured values is 1.75 g to 1.93 g, which is a difference of 0.18 g, or 10% of the average value, which is poor precision. Precision is a measure of how close the measurements are to each other (*Chemistry* 8th ed. pages 13–14/9th ed. pages 13–14).

FREE-RESPONSE QUESTIONS

1 (a) Since density is a ratio of the mass of the object compared to the volume of the object, you might first determine the mass of the sugar cube using a balance. Greater precision is possible with an analytical balance. Then determine the volume of this regular cube by measuring its length, width, and depth (they should all be the same, of course!) and calculating the volume [$V = (l) \times (w) \times (d) = (\text{side})^3$]. It is then a simple step to divide the mass by the volume just calculated (*Chemistry* 8th ed. pages 24–25/9th ed. pages 26–27).

(b) The second method suggests a liquid displacement method for determining volume. However the liquid used may not be water since the object is water soluble. Select a nonpolar liquid to keep this polar solid from dissolving, perhaps, 1,1,1-trichloroethane (*Chemistry* 8th ed. pages 24–25 and 130–132/9th ed. pages 26–27 and 139–141).

2. The 10-mL graduated cylinder is both accurate and precise for volumes of less than 10 mL. The evidence for this is the low percent error (–1.80%) and the low standard deviation. These values indicate measured values that are close to theoretical and values that are reproducible. The 100-mL graduated cylinder is the second most accurate measuring device with a percent error of –2.43%, but the precision is actually not as good as the 100-mL beaker or 50-mL Erlenmeyer flask. The graduated cylinder has a standard deviation of 0.0347, which is higher than the beaker (0.0250) and the flask (0.0306) (*Chemistry* 8th ed. pages 12–14 and A10–12/9th ed. pages 12–14 and A10–A12).

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