Chemistry

Student Solutions Manual

10th Edition

Whitten | Davis | Peck | Stanley

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TENTH EDITION

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Foreword to the Students

This Solutions Manual supplements the textbook, *General Chemistry*, tenth edition, by Kenneth W. Whitten, Raymond E. Davis, M. Larry Peck and George Stanley. The solutions of the 1441 even-numbered problems at the end of the chapters have been worked out in a detailed, step-by-step fashion.

Your learning of chemistry serves two purposes: (1) to accumulate fundamental knowledge in chemistry which you will use to understand the world around you, and (2) to enhance your ability to make logical deductions in science. This ability comes when you know how to reason in a scientific way and how to perform the mathematical manipulations necessary for solving certain problems. The excellent textbook by Whitten, Davis, Peck and Stanley provides you with a wealth of chemical knowledge, accompanied by good solid examples of logical scientific deductive reasoning. The problems at the end of the chapters are a review, a practice and, in some cases, a challenge to your scientific problem-solving abilities. It is the fundamental spirit of this Solutions Manual to help you to understand the scientific deductive process involved in each problem.

In this manual, I provide you with a solution and an answer to the numerical problems, but the emphasis lies on providing the step-by-step reasoning behind the mathematical manipulations. In some cases, I present as many as three different approaches to solve the same problem, since we understand that each of you has your own unique learning style. In stoichiometry as well as in many other types of calculations, the "unit factor" method is universally emphasized in general chemistry textbooks. I think that the over-emphasis of this method may train you to regard chemistry problems as being simply mathematical manipulations in which the only objective is to cancel units and get the answer. My goal is for you to understand the principles behind the calculations and hopefully to visualize with your mind's eye the chemical processes and the experimental techniques occurring as the problem is being worked out on paper. And so I have dissected the "unit factor" method for you and introduced chemical meaning into each of the steps.

I gratefully acknowledge the tremendous help over the years provided by Frank Kolar in the preparation of this manuscript.

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1 The Foundations of Chemistry

1-2. ***Refer to the Introduction to Chapter 1 and a dictionary.***

(a) Organic chemistry is the study of the chemical compounds of carbon and hydrogen and a few other elements.

(b) Forensic chemistry deals with the chemistry involved in solving crimes, including chemical analyses of crime scene artifacts, such as paint chips, dirt, fluids, blood, and hair.

(c) Physical chemistry is the study of the part of chemistry that applies the mathematical theories and methods of physics to the properties of matter and to the study of chemical processes and the accompanying energy changes.

(d) Medicinal chemistry is the study of the chemistry and biochemistry dealing with all aspects of the medical field.

1-4. ***Refer to the Sections 1-1, 1-4, 1-8, 1-13 and the Key Terms for Chapter 1.***

(a) Weight is a measure of the gravitational attraction of the earth for a body. Although the mass of an object remains constant, its weight will vary depending on its distance from the center of the earth. One kilogram of mass at sea level weighs about 2.2 pounds (9.8 newtons), but that same one kilogram of mass weighs less at the top of Mt. Everest. In more general terms, it is a measure of the gravitational attraction of one body for another. The weight of an object on the moon is about 1/7th that of the same object on the earth.

(b) Potential energy is the energy that matter possesses by virtue of its position, condition, or composition. Your chemistry book lying on a table has potential energy due to its position. Energy is released if it falls from the table.

(c) Temperature is a measurement of the intensity of heat, *i.e.* the "hotness" or "coldness" of an object. The temperature at which water freezes is 0°C or 32°F.

(d) An endothermic process is a process that absorbs heat energy. The boiling of water is a physical process that requires heat and therefore is endothermic.

(e) An extensive property is a property that depends upon the amount of material in a sample. Extensive properties include mass and volume.

1-6. ***Refer to the Section 1-1 and the Key Terms for Chapter 1.***

A reaction or process is exothermic, in general, if heat energy is released, but other energies may be released.

(a) The discharge of a flashlight battery in which chemical energy is converted to electrical energy is referred to as being exothermic the chemical reaction occurring in the battery releases heat.

(b) An activated light stick produces essentially no heat, but is considered to be exothermic because light is emitted.

1-8. ***Refer to Sections 1-1 and 1-5, and the Key Terms for Chapter 1.***

(a) Combustion is an exothermic process in which a chemical reaction releases heat.

(b) The freezing of water is an exothermic process. Heat must be removed from the molecules in the liquid state to cause solidification.

(c) The melting of ice is an endothermic process. The system requires heat to break the attractive forces that hold solid water together.

(d) The boiling of water is an endothermic process. Molecules of liquid water must absorb energy to break away from the attractive forces that hold liquid water together in order to form gaseous molecules.

(e) The condensing of steam is an exothermic process. The heat stored in water vapor must be removed for the vapor to liquefy. The condensation process is the opposite of boiling which requires heat.

(f) The burning of paper is an exothermic process. The heat generated can be used to light the wood in a fireplace.

1-10. ***Refer to Section 1-1.***

Einstein's equation, written as $E = mc^2$, tells us that the amount of energy released when matter is transformed into energy is the product of the mass of matter transformed and the speed of light squared. From this equation, we see that energy and matter are equivalent. Known as the Law of Conservation of Matter and Energy, we can use this equation to calculate the amount of energy released in a nuclear reaction because it is proportional to the difference in mass between the products and the reactants. The energy released (in joules) equals the mass difference (in kilograms) times the square of the speed of light (in m/s).

1-12. ***Refer to Section 1-1.***

Electrical motors are less than 100% efficient in the conversion of electrical energy into useful work, since a part of that energy is converted into frictional heat which radiates away.

However, the Law of Conservation of Energy still applies:

electrical energy = useful work + heat

1-14. ***Refer to Section 1-3 and Figures 1-7 and 1-8.***

Solids: are rigid and have definite shapes;
they occupy a fixed volume and are thus very difficult to compress;
the hardness of a solid is related to the strength of the forces holding the particles of a solid together; the stronger the forces, the harder is the solid object.

Liquids: occupy essentially constant volume but have variable shape;
they are difficult to compress;
particles can pass freely over each other;
their boiling points increase with increasing forces of attraction among the particles.

Gases: expand to fill the entire volume of their containers;
they are very compressible with relatively large separations between particles.

The three states are alike in that they all exhibit definite mass and volume under a given set of conditions. All consist of some combination of atoms, molecules or ions. The differences are stated above. Additional differences occur in their relative densities:

gases <<< liquids < solids.

Molecular representations of these three phases can be seen in Figure 1-8. Note that water is an exceptional compound. The density of the liquid is greater than the solid phase. That is why solid ice floats in liquid water

1-16. ***Refer to Section 1-6 and the Key Terms for Chapter 1.***

(a) A substance is a kind of matter in which all samples have identical chemical composition and physical properties, e.g., iron (Fe) and water (H_2O).

(b) A mixture is a sample of matter composed of two or more substances in variable composition, each substance retaining its identity and properties, e.g., soil (minerals, water, organic matter, living organisms, etc.) and seawater (water, different salts, dissolved gases, organic compounds, living organisms, etc.).

(c) An element is a substance that cannot be decomposed into simpler substances by chemical means, e.g., nickel (Ni) and nitrogen (N).

(d) A compound is a substance composed of two or more elements in fixed proportions. Compounds can be decomposed into their constituent elements by chemical means. Examples include water (H_2O) and sodium chloride (NaCl).

1-18. ***Refer to Section 1-6.***

(a) Gasoline is a homogeneous liquid mixture of organic compounds distilled from oil.

(b) Tap water is a homogeneous liquid mixture, called an aqueous solution, containing water, dissolved salts, and gases such as chlorine and oxygen.

(c) Calcium carbonate is a compound, $CaCO_3$, consisting of the elements Ca, C and O in the fixed atomic ratio, 1:1:3.

(d) Ink from a ball-point pen is a homogeneous mixture of solvent, water and dyes.

(e) Vegetable soup is a heterogeneous mixture of water, vegetables and the compound, NaCl (table salt), depending on the recipe.

(f) Aluminum foil is composed of the metallic element, Al.

1-20. ***Refer to Section 1-6.***

The coin is a heterogeneous mixture of gold and copper because it consists of two distinguishable elements that can be recognized on sight.

1-22. ***Refer to Section 1-4.***

(a) Striking a match, causing it to burst into flames, is a chemical property, since a change in composition is occurring of the substances in the match head and new substances including carbon dioxide gas and water vapor, are being formed.

(b) The hardness of steel is a physical property. It can be determined without a composition change.

(c) The density of gold is a physical property, since it can be observed without any change in the composition of the gold.

(d) The ability of baking soda to dissolve in water with the evolution of carbon dioxide gas is a chemical property of baking soda, since during the reaction, its composition is changing and a new substance is being formed.

(e) The ability of fine steel wool to burn in air is a chemical property of steel wool since a compositional change in the steel wool occurs and heat is released.

(f) The ripening of fruit is a chemical property. When the temperature of the fruit decreases when put into a refrigerator, the rate of the chemical reaction slows. So, the lowering of the fruit's temperature is a physical change, but temperature has a definite effect on the chemical properties of the fruit.

1-24. *Refer to Section 1-5.*

The observations that identify chemical properties are: (c) ultraviolet light converts ozone into oxygen, (e) sodium metal reacts violently with water, and (f) CO_2 does not support combustion.

Some chemists think that dissolution is a chemical process, since it is actually very complex, so some chemists would include (a).

1-26. *Refer to Section 1-1 and the Key Terms for Chapter 1.*

(b), (d) and (e) are examples of potential energy. An inflated balloon (b) possesses energy which will be released if it is popped. The stored chemical energy in a flashlight battery (d) will convert to electrical energy, then into kinetic energy once it is put to use. A frozen lake (e) is stored energy. Once spring comes, the water molecules will be free to move, the lake will be circulating and the energy will convert to kinetic energy. However, a lake can also be a source of potential energy that can be converted into kinetic energy if the water is released via a dam.

(a), (c) and (f) are all examples of kinetic energy due to their motion.

1-28. *Refer to Section 1-5.*

When the sulfur is heated, some of it obviously became a gas. However, there is not enough information to tell whether or not this was the result of a physical or a chemical change.

Hypothesis 1: Solid sulfur could be changing directly into gaseous sulfur. This is a physical change called sublimation.

Hypothesis 2: Solid sulfur could be reacting with oxygen in the air to form a gaseous compound consisting of sulfur and oxygen. This would be a chemical change. The sharp odor may indicate the presence of SO_2, but the smell test is not conclusive.

To verify which hypothesis is correct, we need to identify the gas that is produced.

1-30. *Refer to Appendix A.*

(a) $423.\underline{00}6$ mL = 4.23006×10^2 mL (6 significant figures)

(b) $0.001\underline{0}73\underline{0}4\underline{0}$ g = 1.073040×10^{-3} g (7 significant figures)

(c) $1\underline{0}81.\underline{0}2$ pounds = 1.08102×10^3 pounds (6 significant figures)

1-32. *Refer to Appendix A.*

(a) 50600	(c) 0.1610	(e) 90000.
(b) 0.0004060	(d) 0.000206	(f) 0.0009000

1-34. *Refer to Appendix A.*

? volume (cm^3) = 252.56 cm x 18.23 cm x 6.5 cm = 29927 = $\mathbf{3.0 \times 10^4\ cm^3}$ (2 significant figures based on 6.5 cm)

1-36. ***Refer to Section 1-9, the conversion factors from Tables 1-6 and 1-8, and Examples 1-3 and 1-4.***

(a) $? \text{ km} = 453.4 \text{ m} \times \frac{1 \text{ km}}{1000 \text{ m}} = \mathbf{0.4534 \text{ km}}$

(b) $? \text{ m} = 36.3 \text{ km} \times \frac{1000 \text{ m}}{1 \text{ km}} = \mathbf{3.63 \times 10^4 \text{ m}}$

(c) $? \text{ g} = 487 \text{ kg} \times \frac{1000 \text{ g}}{1 \text{ kg}} = \mathbf{4.87 \times 10^5 \text{ g}}$

(d) $? \text{ mL} = 1.32 \text{ L} \times \frac{1000 \text{ mL}}{1\text{L}} = \mathbf{1.32 \times 10^3 \text{ mL}}$

(e) $? \text{ L} = 55.9 \text{ dL} \times \frac{1 \text{ L}}{10 \text{ dL}} = \mathbf{5.59 \text{ L}}$

(f) $? \text{ cm}^3 = 6251 \text{ L} \times \frac{1000 \text{ cm}^3}{1 \text{ L}} = \mathbf{6.251 \times 10^6 \text{ cm}^3}$ (Note: $1 \text{ cm}^3 = 1 \text{ mL}$)

1-38. ***Refer to Section 1-9, the conversion factors listed in Table 1-8, and Example 1-9.***

$? \text{ cents/L} = \frac{\$3.119}{1 \text{ gal}} \times \frac{1 \text{ gal}}{4 \text{ qt}} \times \frac{1.057 \text{ qt}}{1 \text{ L}} \times \frac{100 \text{ cents}}{\$1} = \mathbf{82.42 \text{ cents/L}}$

1-40. ***Refer to Section 1-10, the conversion factors from Table 1-8, and Examples 1-7 and 1-9.***

(a) $? \text{ L} = 0.750 \text{ ft}^3 \times \frac{(12 \text{ in})^3}{(1 \text{ ft})^3} \times \frac{(2.54 \text{ cm})^3}{(1 \text{ in})^3} \times \frac{1 \text{ L}}{1000 \text{ cm}^3} = \mathbf{21.2 \text{ L}}$

(b) $? \text{ pints} = 1.00 \text{ L} \frac{1.057 \text{ qt}}{1 \text{ L}} \times \frac{2 \text{ pt}}{1 \text{ qt}} = \mathbf{2.11 \text{ pt}}$

(c) $? \frac{\text{km}}{\text{L}} = \frac{1 \text{ mile}}{1 \text{ gal}} \times \frac{1.609 \text{ km}}{1 \text{ mile}} \times \frac{1 \text{ gal}}{4 \text{ qt}} \times \frac{1.057 \text{ qt}}{1 \text{ L}} = \mathbf{0.4252 \frac{km}{L}}$

Therefore, to convert miles per gallon to kilometers per liter, one multiplies the miles per gallon by the factor, 0.4252.

1-42. ***Refer to Appendix A.***

$\text{Average} = \frac{58.2 + 56.474}{2} = 57.337 = \mathbf{57.3\ \%}$ since the answer must be rounded to the tenths place

1-44. ***Refer to Section 1-9, Appendix A, the conversion factors from Table 1-8 and Example 1-9.***

(a) $18 \text{ pints} \times \frac{1 \text{ qt}}{2 \text{ pints}} = \mathbf{9.0 \text{ qt}}$

(b) $\frac{55.0 \text{ miles}}{\text{hr}} \times \frac{1.609 \text{ km}}{1 \text{ mile}} = \mathbf{88.5 \text{ km/hr}}$

(c) $15.45 \text{ s} + 2.2 \text{ s} + 55 \text{ s} = 72.65 \text{ s} = \mathbf{73 \text{ s}}$ since the answer must be rounded to the one's place.