



SCIENCE

STUDENT BOOK

▶ **11th Grade | Unit 7**

SCIENCE 1107

EQUILIBRIUM SYSTEMS

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Equilibrium Systems

Introduction

Our universe is in a constant state of decay. Since the Fall of man, our universe was locked into the bondage of corruption, which is plainly evident in the physical world.

Equilibrium systems exist throughout nature. You are a walking equilibrium system. When your system is out of equilibrium, you are probably sick. The force of increasing entropy countering the force of decreasing enthalpy is at the heart of our study of equilibrium systems. Thus, after the Fall, the natural balance between these two forces was upset and the balance shifted to the decay, or random motion (entropy), side of the equilibrium. The energy source is running down and is lost to the environment. However, this state will be banished once the great work of redemption has been consummated by Christ, and the “new heavens and new earth” (Isaiah 65:17; 66:22; 2 Peter 3:13; and Revelation 21:1) are established by the power of God.

This LIFEPAK® will help you to understand some of the common equilibriums around us, what affects their state, and how they can be altered. Much of what you have already learned will be applied in this LIFEPAK®.

Objectives

Read these objectives. The objectives tell you what you will be able to do when you have successfully completed this LIFEPAK. When you have finished this LIFEPAK, you should be able to:

1. Calculate gram-formula-weights.
2. Describe the three main characteristics of solutions.
3. Calculate molarity.
4. Define and identify ions.
5. Balance equations.
6. Apply the concepts of equilibrium to the dissolved-undissolved system.
7. Apply the *Law of Chemical Equilibrium* to the dissolved-undissolved system.
8. Predict if a precipitate will form given specific environments.
9. Describe the variables affecting the rate of dissolving.
10. Define and identify an acid.
11. Define and identify a base.
12. Describe neutralization.
13. Define and apply pH.
14. Apply equilibrium concepts to acid-base reactions.
15. Define and identify a salt.
16. Identify anions and cations.
17. Identify oxidizing and reducing agents.
18. Write half-reactions.
19. Balance redox reactions.

1. SOLUTIONS

Sodium chloride, sugar, glycerine, and water are four pure substances. Each is characterized by definite properties, such as vapor pressure, melting point, boiling point, density, crystalline shape, and so on. Suppose we mix some of these pure substances. Sodium chloride dissolves when placed in contact with water. The solid disappears and becomes part of the liquid. Likewise, sugar in contact with water dissolves. When glycerine is added to water, the two pure substances mix to give a liquid similar in appearance to the original liquids. The saltwater mixture, the sugar-water mixture, and the glycerine-water mixture are called solutions. They differ from pure substances in that their properties vary, depending upon the relative amounts of the constituents. The behavior of these solutions

during phase changes is dramatically different from that just described for pure substances. These differences provide a basis for making a distinction between pure substances and solutions and a basis for deciding whether a given material is a pure substance or a solution. If you need more review of these ideas, restudy Science LIFEPAK 1102.

In this section you will review gram-formula-weights, moles, and equations; and you will learn about types and characteristics of solutions. Although many kinds of solutions exist, most of what you will study will involve water solutions. All of your study in this section will be a background for the next three sections of this LIFEPAK; therefore, study the first section carefully and thoroughly.

Section Objectives

Review these objectives. When you have completed this section, you should be able to:

1. Calculate gram-formula-weights.
2. Describe the three main characteristics of solutions.
3. Calculate molarity.
4. Define and identify ions.
5. Balance equations.



MOLES

You have learned that moles are units used to describe a certain fixed number of objects or items. Avogadro's Number, N , is used to define the number of objects in a mole. A mole is 6.02×10^{23} units, 22.4 liters of gas at STP, or one gram-formula-weight of any substance. The mass necessary to produce a mole is equal to the sum of atomic mass units (expressed in grams). The number of units in a mole is always the same, but the mass of a mole of one substance may vary from the mass of one mole of another substance.

This section of the LIFE PAC will review Science LIFE PAC 1103. If you are in doubt, review for added help.

Gram-formula-weight. Study the Periodic Table of Elements that came with Science LIFE PAC 1101. Also review Section 5 of Science LIFE PAC 1103. A mole of any substance is equal to the sum of the atomic

masses of the elements in the substance expressed in grams. Study the following examples.

The gram-formula-weight (g.f.w.) of NaOH is calculated:

1 - Na	=	23	grams/mole	
1 - O	=	16	grams/mole	
1 - H	=	1	grams/mole	
NaOH	=	40	grams/mole	= 1 g.f.w.

The gram-formula-weight of Na_2CO_3 is calculated:

2 - Na	=	46	grams/2 moles	
1 - C	=	12	grams/mole	
3 - O	=	48	grams/3 moles	
Na_2CO_3	=	106	grams/mole	= 1 g.f.w.

Complete these activities.

- 1.1 What is the atomic mass of one mole of Na? _____
- 1.2 What is the atomic mass of one mole of H? _____
- 1.3 What is the atomic mass of one mole of O? _____
- 1.4 What is the g.f.w. of one mole of H_2O ? _____
- 1.5 What is the atomic mass of one mole of C? _____
- 1.6 What is the g.f.w. of one mole of CO_2 ? _____
- 1.7 What is the g.f.w. of one mole of NaHCO_3 ? _____
- 1.8 What is the g.f.w. of one mole of H_2CO_3 ? _____
- 1.9 What is the g.f.w. of one mole of CO? _____

Balancing equations. A balanced equation is one that has the same number of atoms of each kind on both sides of the reaction. A balanced equation

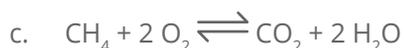
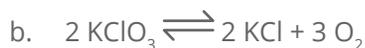
shows a Conservation of Mass. Review Science LIFE PAC 1103, Section 5, to refresh your memory on the techniques for balancing equations.

Complete these activities.

1.10 Draw a circle around each coefficient (mole ratio number) and underline each subscript.



1.11 Explain the meanings of each of the following equations.



1.12 In the following equations list the reactants and products.



reactants: _____

products: _____



reactants: _____

products: _____

1.13 Balance the following equations.

**Balanced Equation**

a. _____

b. _____

c. _____

d. _____

e. _____

f. _____

g. _____

**CHECK**

Teacher

Date

TYPES OF SOLUTIONS

Some parts of the earth are heterogeneous—they have many unlike parts. Some of the parts are uniform throughout—they are homogeneous. Familiar examples of heterogeneous materials are granite, which consists of various minerals suspended in another mineral; an oil and vinegar salad dressing, which consists of droplets of oil suspended in aqueous acetic acid; and black smoke, which consists of particles of soot suspended in air. Examples of homogeneous materials are diamond, water, saltwater, and clear air. Heterogeneous materials are hard to describe and classify but we can describe homogeneous materials rather precisely.

Pure substances and solutions are both homogeneous. A homogeneous material that contains only one substance is called a pure substance. A solution is a homogeneous material that contains more than one substance.

We have used the terms gas phase, liquid phase, and solid phase. A phase is a homogeneous part of a system—a part that is uniform and alike throughout. A system is any region, and the material in that region is that which we wish to consider. A system may include only one phase or more than one phase. For a further review of solutions, reread Science LIFE PAC 1102.

Gaseous solutions. All gas mixtures are homogeneous, hence all gas mixtures are solutions. Air is an example. Air is only one phase—the gas phase—and all the molecules, whatever the source, behave as gas molecules. The molecules themselves may have come from gaseous substances, liquid substances, or solid substances. Whatever the source of the constituents, this gaseous solution, air, is a single, homogeneous phase. As with other solutions the constituents of air are separated by phase changes.

Solid solutions. Solid solutions are more rare. Crystals are stable because of the regularity of the positioning of the atoms. A foreign atom interferes with this regularity and decreases the stability of the crystal. Therefore, as a crystal forms, it tends to exclude foreign atoms. Crystallization, as a result, provides a good method for purification of substances from a mixture.

Only in metals are solid solutions relatively common. The atoms of one element may enter the crystal of another element if their atoms have similar



| SiO₂ crystals

size. Gold and copper form such solid solutions. The gold atoms can replace copper atoms in the copper crystal, and in the same way copper atoms can replace gold atoms in the gold crystal. Such solid solutions are called *alloys*. Some solid metals dissolve hydrogen or carbon atoms. For example, steel is iron containing a small amount of dissolved carbon.

Liquid solutions. In your laboratory work, you will deal mostly with liquid solutions. Liquid solutions can be made by mixing two liquids (for example, water and glycerine), by dissolving a gas in a liquid (for example, carbon dioxide and water), or by dissolving a solid in a liquid (for example, sugar and water). The result is a homogeneous system containing more than one substance: a solution. In such a liquid, each component is diluted by the other component. In saltwater the salt dilutes the water and the water dilutes the salt. The solution is only partly made up of water molecules and the vapor pressure of the solution is correspondingly lower than the vapor pressure of pure water. Whereas water must be heated to 100°C to raise the vapor pressure to 760 mm, a salt solution must be heated above 100°C to reach this vapor pressure. Therefore, the boiling point of saltwater is above the boiling point of pure water. The amount the boiling point is raised depends upon the relative amounts of water and salt. As more salt is added, the boiling point increases.

In a similar way, a lower temperature is required to crystallize ice from saltwater or from a glycerine/water solution than from pure water. “Antifreeze” substances added to automobile radiators act on this principle—they dilute the water in the radiator and lower the temperature at which ice can crystallize from the solution. Putting salt on ice around the container of an ice cream freezer lowers the freezing point of the ice-water bath so that the ice

cream freezes. Again the amount that the freezing temperature is lowered depends upon the relative amounts of water and the “antifreeze” compound.

In general, the properties of a solution depend upon the relative amounts of the components. We should be able to specify quantitatively what is present in a solution, that is, to specify its composition. The composition of a solution can be specified in many ways, but one method will suffice for our purposes.

Answer true or false.

- 1.14 _____ All gaseous mixtures are solutions.
- 1.15 _____ Solid solutions are common.
- 1.16 _____ Vapor pressure of water is not affected by the amount of salt added.
- 1.17 _____ Alloys are examples of solid solutions.
- 1.18 _____ Crystallization is a good way to purify solids.
- 1.19 _____ Salt acts as an antifreeze.

CHARACTERISTICS OF SOLUTIONS

The components of a solution are the pure substances that are mixed to form the solution. If the solution contains two components, one is sometimes called the *solvent* and the other the *solute*. These terms are used merely for convenience. Since both must intermingle to form the final solution, we cannot make any important distinction between them. When chemists make a liquid solution from a pure liquid and a solid, they usually call the liquid component the solvent.

Concentrations. To indicate the composition of a particular solution, we must show the relative amounts as well as the kind of components. These relative amounts chemists call *concentrations*. Chemists use different ways of expressing concentration for various purposes. You will study two different ways to calculate concentrations in this LIFEPAK, one in this section and one in the third section.

Chemists often indicate the concentration of a substance in water solution by the number of moles dissolved per liter of solution. This concentration is called the molar concentration. A one-molar (1 M)

solution contains one mole of the solute per liter of total solution, a 2-molar solution (2 M) contains two moles of solute per liter, and a 0.1-molar solution (0.1 M) contains one-tenth mole of solute per liter.





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