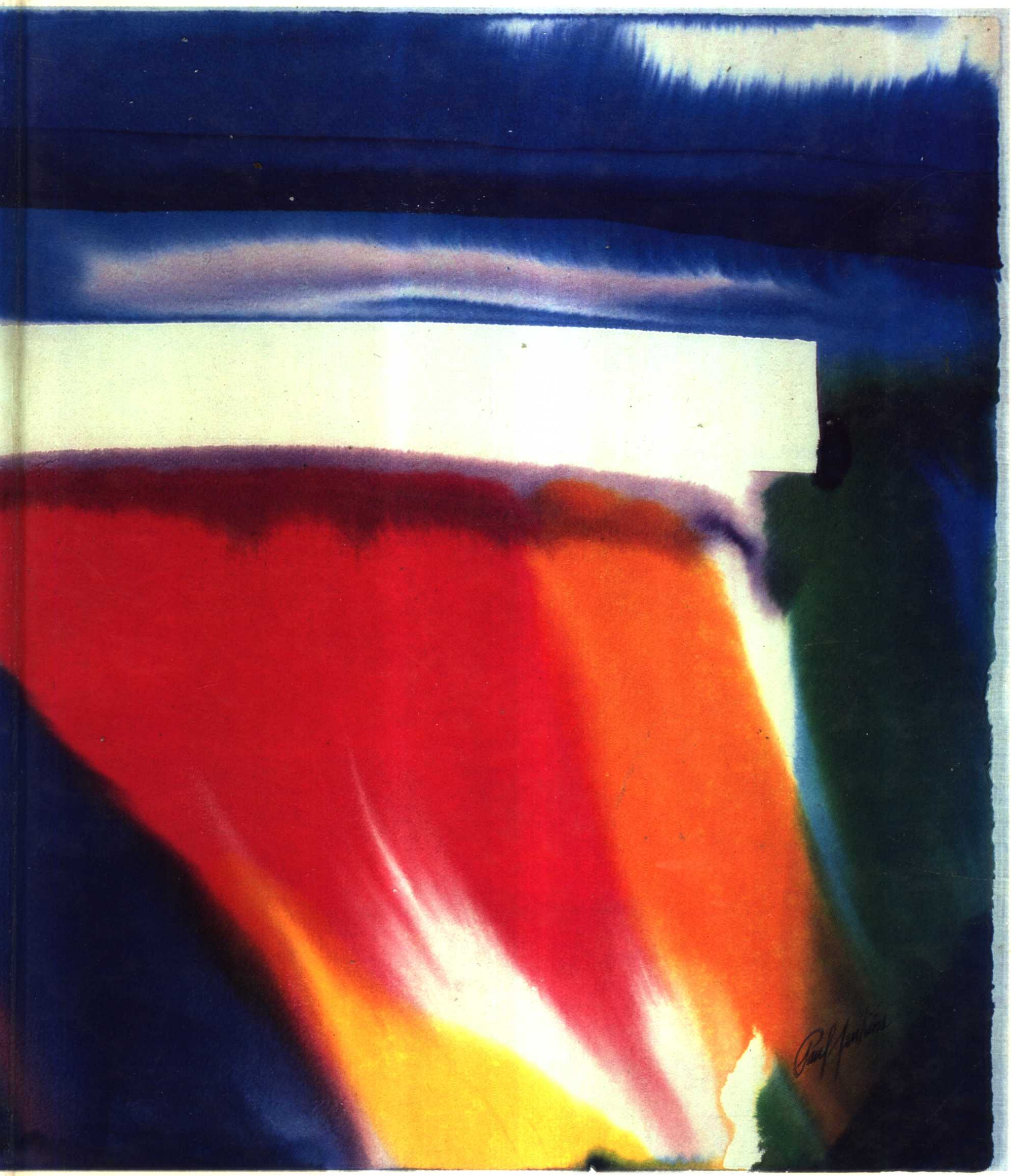
Introduction to

GENERAL, ORGANIC, AND BIOLOGICAL CHEMISTRY



Sally Solomon

Introduction to

GENERAL, ORGANIC, AND BIOLOGICAL CHEMISTRY

Sally Solomon

Drexel University

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INTRODUCTION TO GENERAL, ORGANIC, AND BIOLOGICAL CHEMISTRY

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CALORIC VALUE, COMPOSITION, AND VITAMIN AND MINERAL CONTENT OF COMMON FOODS*

	Amount	Mass, g	Calories	kJ	Protein, g (44–56 g)*	Fat, g	Carbohydrates, g	Cholesterol, mg
Apple (medium)	1	150	70	294	0.3	0.5	18	0
Beef liver	3.5 oz	100	229	958	26.4	10.6	5.3	438
Bread (white)	1 slice	23	60	252	2	1	12	NA
Butter	1 pat	7	50	210	Trace	6	Trace	15.5
Carrots	1 large	100	45	189	1.1	0.2	9.7	0
Cheese	1 oz	28	105	441	7	9	1	30
Chicken (skinless & boneless)	3 oz	85	115	483	20	3	0	70
Chocolate (semisweet)	1 oz	28	132	552	1.5	9	16.5	Trace
Cola	1 cup	240	95	399	0	0	24	0
Corn oil	1 tbsp	14	125	525	0	13.6	0	0
Cottage cheese	1 cup	210	216	904	26	9.4	5.2	32
Egg	1	50	80	336	6	6	0.6	274
Green beans	1 cup	125	30	126	2	Trace	6.8	0
Hamburger	3 oz	85	245	1029	20	17	0	50
Ice cream (vanilla)	1 cup	142	295	1239	6	18	30	59
Macaroni (cooked)	1 cup	130	190	798	6	1	39	0
Margarine	1 pat	7	50	210	Trace	6	Trace	0
Milk	1 cup	244	160	672	9	9	12	35
Orange (medium)	1	136	60	252	1.4	0.1	16.3	0
Orange juice (frozen)	1 cup	249	110	462	1.7	0.1	27	0
Peanut butter	1 tbsp	15	95	399	4	8	3	0
Pork chop	3.5 oz	98	260	1092	16	21	0	130
Potato (baked)	1	99	90	378	3	0.1	21	0
Raisins	½ cup	80	230	962	3	0.3	62	0
Rice (cooked)	1 cup	180	185	777	3.5	0.2	40	NA
Spinach (cooked)	1 cup	180	40	168	5	1	6	0
Sugar	1 tbsp	12	48	201	0	0	12	0
Tomato juice	1 cup	243	45	189	1.9	0.2	10	0
Tuna (in oil)	3 oz	85	170	714	24	7	0.1	20
Yoghurt	1 cup	246	120	504	8	4	13	14

*RDA values are given in parentheses.
NA = not available; 0 = zero or close to zero; trace = small amount present.

B$_1$, thiamin, mg (1–1.4 mg)	B$_2$, riboflavin, mg (1.2–1.6 mg)	Niacin, mg (13–18 mg)	C, mg (60 mg)	A, IU (4000–5000 IU)	D, μg (10 μg)	E, mg (8–10 mg)	K, μg (70–140 μg)	Ca, mg (800–1200 mg)	Fe, mg (10–18 mg)
0.04	0.02	0.1	8	75	0	1.1		8	0.4
0.26	4.2	16.5	27	53,000	0.85	1.4	200	10	8
0.06	0.04	0.5	Trace	Trace	Trace	0.1		16	0.6
0	Trace	Trace	0	230‡	0.16	0.17		1	0
0.06	0.05	0.6	8	12,000	0	0.5		39	0.8
0.05	0.15	Trace	0	265	0.5	0.6		219	0.3
0.05	0.16	7.4	0	25	NA	0.2		8	1.4
0.01	0.03	0.2	0	8	NA	NA		11	1
0	0	0	0	0	0	0	0	3	0.18
0	0	0	0	0	0.03	12		0	0
0.04	0.34	0.2	Trace	342	NA	—		212	0.7
0.04	0.14	Trace	0	260	31	0.8		27	1.1
0.09	0.11	0.6	16	675	NA	NA		62	0.8
0.04	0.2	4	0	25	0	0.1		9	3
0.05	0.33	0.1	1	740	0.1	0.1		175	0.1
0.23	0.14	1.9	0	0	0	—		14	1.4
Trace	Trace	0	0	230‡	NA	4.6	0	1	0
0.08	0.42	0.1	2	350	2.5‡	0.3	2	288	0.1
0.12	0.05	0.5	75	240	0	0.4		49	0.2
0.21	0.03	0.8	112	200	0	NA		22	0.2
0.02	0.02	2.4	0	0	0	NA		11	0.3
0.63	0.18	3.8	0	0	0	0.7		8	2.2
0.1	0.04	1.7	20	Trace	NA	0.06	20	9	0.5
0.1	0.06	0.5	2	8	NA	NA		50	3
0.19	0.1	1.6	0	0	0	0.4		17	1.5
0.12	0.26	1	50	14,600	0	2.4	600	167	4
0	0	0	0	0	0	0	0	0	0
0.13	0.07	1.8	39	1,000	0	2		17	2
0.04	0.1	10.1	0	70	6	0.5		7	1.6
0.09	0.43	0.2	2	150	NA	NA		295	0.1

†Only selected values listed for vitamin K.
‡Vitamins added to fortify these foods.

PREFACE

Introduction to General, Organic, and Biological Chemistry will appeal particularly to anyone planning to enter one of the allied health fields such as nursing, medical technology, physical therapy, or nutrition. No previous chemistry course is needed to understand the material.

As suggested by the title of the book, the topics covered belong to three main areas:

General Chemistry: Chapters 1 through 11 discuss the fundamental principles of chemistry. The behavior of matter is understood in terms of atoms and molecules.
Organic chemisty: Chapters 12 through 20 cover the chemistry of carbon-containing compounds.
Biochemistry: Chapters 21 through 28 are devoted to the structures and functions of biomolecules.

Throughout this text the excitement and usefulness of chemistry are conveyed by making clear connections between chemical principles, the surrounding world and the human body. Practical applications are presented along with their underlying chemical principles. For instance, acid-base properties of blood are found in the acid-base chapter, fat-soluble vitamins are discussed in a chapter on lipids, and sickle-cell anemia is covered in the nucleic acids chapter, where the disease's molecular basis is best explained. The practical applications of chemistry are not isolated by setting them aside from the rest of the material, rather they are integrated into the text to emphasize the close correlation between chemical theory and its biological application.

An important feature of this text is the great number of learning aids intended to make chemistry more accessible and enjoyable.

Sample exercises: Every chapter contains sample exercises that are solved in detail.
Exercises: Accompanying each sample exercise is an unsolved exercise. The answers to all the unsolved in-chapter exercises are given at the end of the book along with the answers to the odd-numbered end-of-chapter questions.
Margin comments: The wide margins of this text include comments that summarize ideas or supply interesting information about the topic in the adjacent paragraph.
Highlighted margin comments: Marginal comments printed in color are set next to paragraphs that discuss medical and biological applications. These notes can be used to quickly locate material that is of special interest to students of the allied health sciences.
Keywords: New terms are shown in boldface. At the end of every chapter these are listed as keywords, all of which are defined in the glossary found at the end of the text.

Summary: Each chapter is accompanied by a concise summary.
Appendexes: For students who need extra help in mathematical skills, there are appendexes that cover exponentials, significant figures, and algebraic expressions. Like the chapters that precede them, each appendix has both sample exercises and unsolved exercises.
Color photographs: In the middle of the text you will find a section of color photographs, most of which were taken by Paul Winkfield, a photographer with incredible patience who welcomed the challenge of bringing chemical experiments to life on film. The color plates are not restricted to simply pretty pictures of colorful compounds but are designed to demonstrate chemical principles discussed throughout the text. Most of the experiments that we photographed can be performed in the classroom, laboratory, or lecture hall using instructions found in the *Instructor's Manual* for this text.

There are a great many people without whom this book could never have been produced. I wish to thank the following reviewers for their constructive comments and for the many positive remarks that helped provide me with the inspiration that I needed to finish this project.

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Seymour Yolles	University of Delaware

The people at McGraw-Hill were patient and helpful. I was continually amazed by the ability of Randi Kashan, editor, to make the right suggestion about whatever it was we were trying to do, to spot my mistakes, and to keep me on schedule (almost). The efforts of Jack Maisel, editing supervisor, who coordinated the production process, seem to me just short of miraculous. Jo Jones, the designer, intuitively understood what inside and cover design would best lend itself to my writing and teaching style. Finally, the enthusiastic response from my students at Drexel University to my chemistry courses encouraged me to create my own textbook.

Sally Solomon

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1. MEASUREMENT

1.1 INTRODUCTION

To the untrained eye it may seem that chemists are able to do magic in the laboratory, producing something where once there was nothing. From simple chemicals the modern chemist can synthesize a drug with the ideal structural features to treat a particular disease or create a remarkable plastic with just the right properties to replace a worn body part. Very rarely does a sudden, almost magical, discovery lead the way to this sort of success. In most cases careful, occasionally tedious, experimentation must come first.

Performing experiments in chemistry and interpreting their results is what chemists do. It is with the devices used to produce measured quantities, the units in which they are expressed, and the techniques used to do calculations upon them that the study of chemistry begins.

1.2 EXPERIMENTS

A formal term for the experimental approach to the solution of problems is the **scientific method**. The scientific method can be defined as an approach by which scientists combine an idea, or *hypothesis,* with experimental observations to arrive at *theories,* which not only explain the experiments but can also be used to predict the results of future experiments. The scientific method is applied by combining experimental observation with new ideas until it is possible to predict what will happen with no exceptions. Let us put the scientific method into practice in a clinical situation. Suppose you notice that a certain patient becomes violently ill after taking aspirin. Your hypothesis is that the person is allergic to aspirin. Future experiments seem to support this hypothesis. Every time aspirin is administered in any of its forms, the individual exhibits the same hypersensitivity. Then one day you notice the identical allergic reaction when the same patient drinks a glass of root beer. Because of this new experimental observation you must modify the original hypothesis to state that similar substances present in aspirin tablets and in root beer are causing the allergic reaction. You design a series of experiments to prove that the patient is allergic to aspirin, root beer, or anything containing a chemical known as a *salicylate.* All substances that you test are in agreement with the modified hypothesis and so it stands as theory. All theories survive until new observations no longer can support them. It is probable that 100 years from now, certain theories presented in this text will have to be modified or discarded to explain events that have yet to be observed.

Experiments fall into one of two broad categories. In **qualitative** experiments the presence or absence of some physical quantity is noted. When the physical quantity is measured to see just how much of it there is, the experiment becomes **quantitative.** For example, suppose the experiment is to test for glucose in urine. The observation that the urine sample contains glucose is qualitative; the observation that the urine sample contains 10 mg of glucose is quantitative.

1.3 UNITS AND THE SI SYSTEM

Measured quantities such as 10 mg include a numerical value (10) and a unit (mg). The unit describes the physical quantity that is being measured, which in this case is a unit of mass, the milligram (mg). A practical and useful set of units must be internationally accepted and must include units that are unambiguously defined. Three sets of units in use are the English system, the metric system, and the latter's close relative, the International System of Units, called the **SI system.** The English system, with units such as foot and pound, is rarely if ever mentioned in scientific studies, although it is still used in the United States for other purposes. The metric system, which includes the meter and kilogram units, has been widely adopted, as most countries in the world have "gone metric" or are "going metric." The International System of Units was created in 1969 to clear up any possible confusion about which units should be included in the modern metric system. The **SI system** includes the SI base units, the SI derived units, and the SI prefixes.

SI Units

There are seven **base units** in the SI system: the meter, kilogram, kelvin, second, mole, ampere, and candela. Their names and abbreviations appear in Table 1.1.

TABLE 1.1 SI BASE UNITS

Physical quantity	Name of base unit	Abbreviation
Length	meter	m
Mass	kilogram	kg
Time	second	s
Amount of substance	mole	mol
Temperature	kelvin	K
Electric current	ampere	A
Luminous intensity	candela	cd

TABLE 1.2 SI PREFIXES

Prefix	Abbreviation	Meaning*
pico	p	10^{-12} ($1/10^{12}$ or one-trillionth)
nano	n	10^{-9} ($1/10^{9}$ or one-billionth)
micro	μ	10^{-6} ($1/10^{6}$ or one-millionth)
milli	m	10^{-3} ($1/10^{3}$ or one-thousandth)
centi	c	10^{-2} ($1/10^{2}$ or one-hundredth)
deci	d	10^{-1} (1/10 or one-tenth)
kilo	k	10^{3} (1000 or one thousand times)
mega	M	10^{6} (1,000,000 or one million times)

*See Appendix A for coverage of scientific notation.

Two of them, the ampere and the candela, will not be discussed further, since they will not be encountered again in this text. By some combination of the base units, it is possible to express the unit for any other measured quantity.

The slash sign (/) means "per" or "divided by."

The **SI derived units** are those which are formed by multiplication or division of the SI base units. One example of a derived unit is the unit for speed, which is meters divided by seconds or meters per second, m/s. We will introduce more derived units as they appear in the topics to come.

SI Prefixes

Many SI units are formed from a prefix and a base unit.

The SI prefixes are used to form multiples or fractions of SI units. Many quantities are conveniently expressed in terms of a prefix plus a base unit. For example, the quantity 1000 m may also be written as 1 kilometer (1 km), since the prefix kilo means 1000. The prefixes most often added to base units are listed in Table 1.2. That one of the base units, the kilogram, already has a prefix is a peculiar feature of the SI system, which is explained in the discussion about mass in Section 1.5. Multiples and fractions of the kilogram are formed by attaching an SI prefix not to kilogram (kg) but to gram (g).

SAMPLE EXERCISE 1.1

Using Tables 1.1 and 1.2 express the following quantities in terms of a prefix plus a base unit:

(a) 0.01 meter

(b) One-millionth of a second

Solution:

(a) $0.01 = 1/100$

(b) One-millionth $= 1/10^6 = 10^{-6}$

Prefix: centi (c)
Base unit: meter (m)
Complete unit: centimeter (cm)

Prefix: micro (μ)
Base unit: second (s)
Complete unit: microsecond (μs)

EXERCISE 1.1

The mean cell hemoglobin (MCH) refers to the average mass of hemoglobin per red blood cell. The normal MCH is 27 to 31 pg. Would the value 3×10^{-11} g fall within this range?

Definitions of SI Units

To measure a physical quantity we must compare it with some other well-defined physical quantity. For example, suppose we wished to measure the length of a room. One way would be to walk across the room, carefully placing one foot in front of the other. The length of the room could then be measured in terms of some number of "person-feet." But whose feet? Person-feet are very poorly defined and could be acceptable only in the crudest measurements. What makes the SI base units excellent references for measuring quantities is that they are clearly and unambiguously defined. In the next two sections we will take a careful look at the definitions for two of the SI *standard* units, the meter and the kilogram. Standard units such as these must be unchanging and reproducible.

1.4 LENGTH

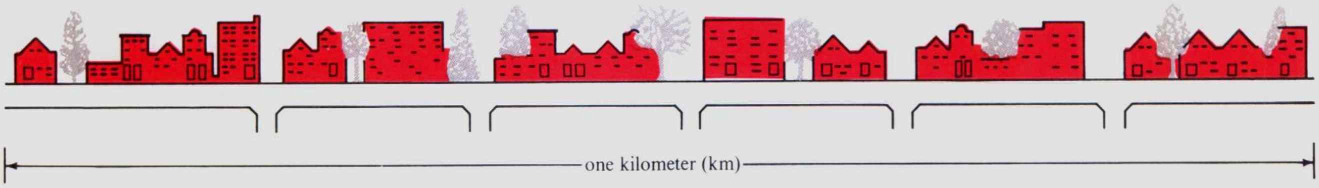
The meter is the base unit of length.

Until the year 1960 the **meter** was defined as the distance between two scratch marks engraved on a metal bar made of the metals platinum and iridium, which are especially resistant to corrosion. This bar is still carefully housed in Sèvres, France at the International Bureau of Weights and Measures. Meter sticks were made by comparison with this standard or with copies of it.

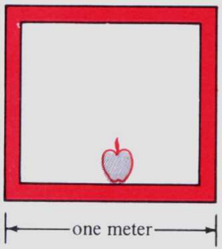
Wavelengths of light are discussed in section 2.11.

A better method for defining the meter divides the meter distance into a number of very small divisions equal to the wavelength of the red light which comes from glowing krypton gas, which glows in the same way that the neon in neon lights does. The number of red wavelengths which spans the distance between the scratches on the standard meter bar is 1,650,763.73, a number which can be accurately reproduced anywhere in the world. This new definition did not change the length of the meter but considerably improved the accuracy with which it could be measured. Another advantage is that objects, even those made of relatively unreactive metals, can change with time, but a number of wavelengths always stays the same.

Some important SI length units and the distances to which they correspond can be illustrated by viewing an ordinary street scene in ever-increasing detail. This is shown in Figure 1.1. Useful relationships involving both SI and non-SI length units are given in Table 1.3.

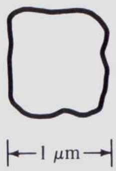


(*a*) One **kilometer (km)** corresponds to about six city blocks.

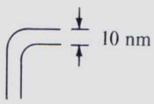


(*b*) One window of one house is about one **meter (m)** wide.

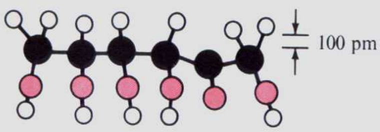
(*c*) On the windowsill is an apple which is about 1 **decimeter (dm)** wide.



(*d*) Taking a microscopic look into the apple reveals a single plant cell which is about one **micrometer (μm)** on a side.



(*e*) Surrounding the plant cell is a cell wall about 10 **nanometers (nm)** thick.



(*f*) Inside the cell is a sugar molecule, fructose, which includes 24 atoms which are roughly 100 **picometers** (pm) thick.

FIGURE 1.1
Length units.