GOB Chemistry

A Comprehensive Set of Imperfect Notes

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Periodic Table of Chemical Elements

Z	mass
Symbol	
Name	

	1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIIIB	9 VIIIB	10 VIIIB	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
1	1 1.0079 H Hydrogen																	2 4.0025 He Helium
2	3 6.941 Li Lithium	4 9.0122 Be Beryllium											5 10.811 B Boron	6 12.011 C Carbon	7 14.007 N Nitrogen	8 15.999 O Oxygen	9 18.998 F Flourine	10 20.180 Ne Neon
3	11 22.990 Na Sodium	12 24.305 Mg Magnesium											13 26.982 Al Aluminium	14 28.086 Si Silicon	15 30.974 P Phosphorus	16 32.065 S Sulphur	17 35.453 Cl Chlorine	18 39.948 Ar Argon
4	19 39.098 K Potassium	20 40.078 Ca Calcium	21 44.956 Sc Scandium	22 47.867 Ti Titanium	23 50.942 V Vanadium	24 51.996 Cr Chromium	25 54.938 Mn Manganese	26 55.845 Fe Iron	27 58.933 Co Cobalt	28 58.693 Ni Nickel	29 63.546 Cu Copper	30 65.39 Zn Zinc	31 69.723 Ga Gallium	32 72.64 Ge Germanium	33 74.922 As Arsenic	34 78.96 Se Selenium	35 79.904 Br Bromine	36 83.8 Kr Krypton
5	37 85.468 Rb Rubidium	38 87.62 Sr Strontium	39 88.906 Y Yttrium	40 91.224 Zr Zirconium	41 92.906 Nb Niobium	42 95.94 Mo Molybdenum	43 96 Tc Technetium	44 101.07 Ru Ruthenium	45 102.91 Rh Rhodium	46 106.42 Pd Palladium	47 107.87 Ag Silver	48 112.41 Cd Cadmium	49 114.82 In Indium	50 118.71 Sn Tin	51 121.76 Sb Antimony	52 127.6 Te Tellurium	53 126.9 I Iodine	54 131.29 Xe Xenon
6	55 132.91 Cs Caesium	56 137.33 Ba Barium	57-71 La-Lu Lanthanide	72 178.49 Hf Halfnium	73 180.95 Ta Tantalum	74 183.84 W Tungsten	75 186.21 Re Rhenium	76 190.23 Os Osmium	77 192.22 Ir Iridium	78 195.08 Pt Platinum	79 196.97 Au Gold	80 200.59 Hg Mercury	81 204.38 Tl Thallium	82 207.2 Pb Lead	83 208.98 Bi Bismuth	84 209 Po Polonium	85 210 At Astatine	86 222 Rn Radon
7	87 223 Fr Francium	88 226 Ra Radium	89-103 Ac-Lr Actinide	104 261 Rf Rutherfordium	105 262 Db Dubnium	106 266 Sg Seaborgium	107 264 Bh Bohrium	108 277 Hs Hassium	109 268 Mt Meitnerium	110 281 Ds Darmstadtium	111 280 Rg Roentgenium	112 285 Uub Ununbium	113 284 Uut Ununtrium	114 289 Uuq Ununquadium	115 288 Uup Ununpentium	116 293 Uuh Ununhexium	117 292 Uus Ununseptium	118 294 Uuo Ununoctium

57 138.91 La Lanthanum	58 140.12 Ce Cerium	59 140.91 Pr Praseodymium	60 144.24 Nd Neodymium	61 145 Pm Promethium	62 150.36 Sm Samarium	63 151.96 Eu Europium	64 157.25 Gd Gadolinium	65 158.93 Tb Terbium	66 162.50 Dy Dysprosium	67 164.93 Ho Holmium	68 167.26 Er Erbium	69 168.93 Tm Thulium	70 173.04 Yb Ytterbium	71 174.97 Lu Lutetium
89 227 Ac Actinium	90 232.04 Th Thorium	91 231.04 Pa Protactinium	92 238.03 U Uranium	93 237 Np Neptunium	94 244 Pu Plutonium	95 243 Am Americium	96 247 Cm Curium	97 247 Bk Berkelium	98 251 Cf Californium	99 252 Es Einsteinium	100 257 Fm Fermium	101 258 Md Mendelevium	102 259 No Nobelium	103 262 Lr Lawrencium

- Alkali Metal
- Alkaline Earth Metal
- Metal
- Metalloid
- Non-metal
- Halogen
- Noble Gas
- Lanthanide/Actinide

Contents

To The Reader

First and foremost, I genuinely care about the progress of every one of my students and I want to see you all succeed. This is why I decided to write this manuscript. This set of lecture notes was designed with a focus on the student–with a focus on you. It introduces the basic concepts of college chemistry in a way that a student of any level can hopefully understand. Some of the chapters included in this guide can be challenging. Success is not an accident. Only with hard work, patience, and perseverance, you will be able to achieve what you want. I hope to encourage you not only to successfully pass this class. More importantly, I hope to inspire you to see that you can do this.

College chemistry is not an easy subject. You may experience frustration due to the terminology or the math content. This guide is developed in chapters and sections to break down the very basics of chemistry concepts. One of my main goals is to help you solve chemistry problems. Solving problems–not only chemistry problems but problems of any kind–is an extremely useful skill in life. Chemists approach the solving of problems in a very specific way. They use critical thinking and previous knowledge to find solutions based on the information presented. As you study this set of lectures, I encourage you to read the different sections of a chapter, highlight the main ideas and find keywords that represent new concepts. Numerous examples are presented in the chapters with the full solution. A lot of examples are also presented without the worked solution, just including the answer. Plenty of end of the chapter problems is further included. After you read the content of a chapter I highly encourage you to work on the end of the chapter problems. As with any skill, practice makes perfect.

I used numerous tools in this guide to help you focus on the most relevant content. For example, when the numerical problems get too complex, an *analyze the problem box* is included to help you identify what is given and what is asked in the problem.

This set of lectures resonates with the open textbook movement that is taking over CUNY as well as SUNY. Education is expensive and you as a student often rely on textbooks to learn. These valuable educational resources are often used for a very limited period and tossed or returned when a class has finished. The open textbook movement aims to alleviate the cost of education by relying on resources that are free for both the students and the educators. Still, these sources are imperfect and not as curated as textbooks, and this is the price to pay. I warn you this set of the lecture is indeed imperfect, hence its title. Yet, it is the result of many hours of work–indeed months of work. Your role is key. I encourage you first to be understanding and patient, and then to contribute to the development of this guide. With your input, we can make this guide a better educational resource. Mind that this guide was written by an educator and as such, it sometimes uses terms and a way of thinking that corresponds to the educators'point of view.

This set of lectures does not intend to replace any textbook. Indeed, there are many high-quality textbooks in the literature that I recommend. For College Chemistry:

- Chemical Principles: The Quest for Insight by Peter Atkins et al.
- Chemistry: The Central Science by Theodore E. Brown et al.
- Chemistry by Steven S. Zumdahl et al.
- Chemistry: The Molecular Nature of Matter and Change by Martin Silberberg et al.
- Chemistry by Raymond Chang et al.
- Chemistry: Atoms First by OpenStax

For GOB Chemistry:

- General, Organic, and Biological Chemistry: Structures of Life by Karen C. Timberlake et al.

With the help of the textbooks above you can certainly expand and complement the information presented in this guide.

This guide was fully coded in LaTeXfrom the cover or the periodic table to the molecular orbital diagrams or the solid representations. Chemistry is a microscopic science not accessible to the naked eye. Visuals play a very important role in chemistry education. Visuals–in the form of images or diagrams–help make chemistry more apparent to the viewer. One of the weak points of many open education chemistry guides is the visuals. They tend to be simplistic with low quality. This guide extensively relies on images and diagrams and uses the Tikz software package and other open-source tools to freshly generate diagrams everytime the book is compiled. All other images used here are open-source images.

The work of chemists is certainly challenging, but also exciting and rewarding. Chemists produce everything from plastics and paints to pharmaceuticals, foods, flavors, fragrances, detergents, and cosmetics. Chemistry students are well-prepared for medical, veterinarian, dentistry, optometry, or pharmacy school. I hope you enjoy this guide and more importantly I wish you success in your career.

Daniel Torres
Brooklyn

Ch. 1. Measurements

MEASURING is an important part of our everyday lives, and very probably you took several measurements today. You might now be sipping a cup of coffee, or perhaps you checked the outside temperature on a street thermometer. You might be planning to bake a cake and need to use a scale and a cup to measure the flour and sugar. A cup, a thermometer, or a scale are measuring devices. It is critical to know how to accurately measure properties and, more importantly, how to transform measurements using prefixes and unit conversions. By learning how to measure and perform operations with units, you will gain experience performing basic chemistry calculations.

1.1 Math skills

There are a few math skills that are critical to be able to carry out chemistry problems. Those skills ental operating with numbers, converting numbers into scientific notation, solving basic equations, and interpreting graphs.

`Place values` We can identify the place value in any number for all its digits. The place values are named based on the location concerning the decimal place. For digits to the left of the decimal place, the place values are called the ones, the tens, the hundreds, and the thousands. For example, in the value 3456s, the number 3 is the thousands place, the number 4 is the hundreds place, 5 is the tens place and 6 is the one's place. For digits to the right of the decimal place, the place values are called the tenths, the hundredths, and the thousandths, mind the suffix *ths*. For example, in the value 3.456Kg, the number 3 is the one place, the number 4 is the tenth place, 5 is the hundredth place and 6 is the thousandth place.

`Basic algebra` Positive numbers are larger than zero whereas negative numbers are smaller than zero. Negative numbers are written with a negative (−) sign. The multiplication or division of two positive or two negative numbers always gives a positive result, whereas the multiplication or division of a positive and a negative number always gives a negative result. When we add two positive numbers the results are always positive, whereas when we add two negative numbers the results are always negative. When positive and negative numbers are added, the results come from substracting the smallest number from the largest number while keeping the sign of the larger number. In a calculator, there is a key $\boxed{(-)}$ (sometimes shown as $\boxed{(+/-)}$) used to switch the sign of a number.

Solving equations When solving the equation below for x

$$3x - 4 = 8$$

some basic algebra rules apply:

1. **Step one:** Place all like terms in one side
2. **Step two:** Isolate the variable you want to calculate
3. **Step three:** Check the answer

When placing like terms on the same side, you can eliminate terms by adding, subtracting, multiplying, or dividing. Make sure you apply those rules on both sides of the equation at the same time. For the example above, we will first eliminate the -4 by subtracting 4 in both sides:

$$3x - \cancel{4} + \cancel{4} = 8 + 4$$

That gives,

$$3x = 12$$

Now we will remove the 3 by diving by three on both sides:

$$\frac{\cancel{3}x}{\cancel{3}} = \frac{12}{3}$$

Therefore we have

$$x = \frac{12}{3} = 4$$

Scientific notation Numbers in science can often be very large or very tinny. For example, the mass of the earth is 5900000000000000000000000000 grams, whereas the size of an atom is of the order 0.00000000001cm. Scientific notation is a standard notation extensively used in the scientific community to simplify numbers. Numbers not expressed in scientific notation are referred to as numbers in full notation (e.g. 0.234g). Numbers expressed in scientific notation place the decimal after the first significant digit while being accompanied by a power of ten (e.g. 2.34×10^{-1}g). Numbers expressed in scientific notation have two parts: the coefficient (a number larger than one but smaller than 10) and the power of ten. For example, the number 3.45×10^5g has a coefficient of 3.45 and a power of ten with positive power. The power of ten can contain a positive or a negative exponent. Positive exponents represent large numbers, whereas negative exponents represent numbers smaller than 1. To express a number larger than one (e.g. 345000g) in scientific notation, we just need to leave the first digit followed by the decimal point which was moved to the left 5 places, giving 3.45×10^5g as indicated below

345000g (full notation) $\quad 3.45 \times 10^5$g (scientific notation)

To express a number smaller than one (e.g. 0.000134g) in scientific notation, we just need to leave the first digit followed by the decimal point which was moved in this case to the right 4 places, giving 1.34×10^{-4}g as indicated below

0.000134g (full notation) $\quad 1.34 \times 10^{-4}$g (scientific notation)

You can enter scientific notation numbers in a calculator using a specific key that contains the $\times$ character and the power of ten. For example, the number 1×10^6 should be typed in a calculator as 1$\boxed{\text{EE}}$6 or 1$\boxed{\text{EXP}}$6 or 1$\boxed{\times 10^x}$6, deppending on your calculator. For numbers with a negative power of ten, you should use the $\boxed{(-)}$ key (sometimes shown as $\boxed{(+/-)}$) to indicate the sign. For example,the number 1×10^{-5} should be typed in a calculator as 1$\boxed{\text{EE}}\boxed{(-)}$5 or 1$\boxed{\text{EXP}}\boxed{(-)}$5 or 1$\boxed{\times 10^x}\boxed{(-)}$5, deppending on your calculator. The calculator display often can display a scientific notation number differently, based on the calculator brand. Below are three possible scenarios:

1.5×10^{-4}	1.5E-04
6.7×10^{-5}	6.7 -05
4.6×10^{-2}	4.6₁₀-02

You can also convert full notation numbers into scientific notation with a calculator key named $\boxed{\text{SCI}}$ often accessible through the second function key.

Sample Problem 1

Convert the following numbers from full to scientific notation or vice versa:
(a) 7462.97 (b) 0.000234 (c) 0.012

SOLUTION

The answers are 7.46297×10^3, 2.34×10^{-4}, and 1.2×10^{-2}.

STUDY CHECK

Convert the following numbers from full to scientific notation or vice versa:
(a) 12000 (b) 0.00076 (c) 45783

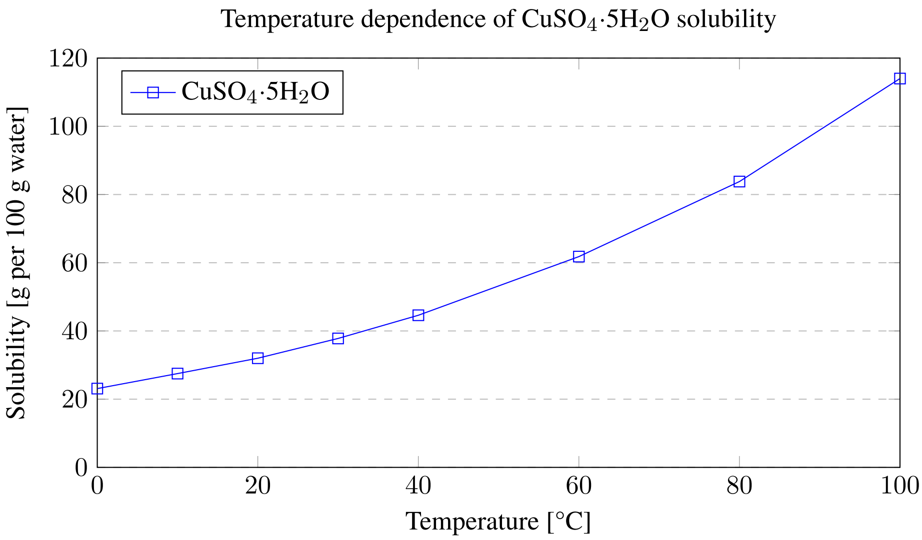


Figure 1.1 A graph plotting the change of solubility of a chemical with temperature.

Interpreting graphs A graph represents data in terms of a relationship between two variables. For example, Figure 1.1 represents the relationship between solubility and temperature. These quantities are plotted along two axes the Y or vertical axis and the X or horizontal axis. The title of the graph indicates what is being represented, and for example, on the graph above we have represented the change of solubility

▼Scales measure mass

© www.wallpaperflare.com

▼Watches are used to measure time

© www.wallpaperflare.com

▼Beakers can carry a liquid volume

© wikipedia

▼Thermometers measure temperature

© PngImg

▼pipets are used in chemistry practice to add an exact volume of liquid

© www.weberscientific.com

of a chemical with temperature. The vertical axis represents solubility given a range between 0 and 120 g/100g of water. The horizontal axis represents temperature with a range between 0 and 100°C. Each point in the graph represents the value of solubility for a given temperature. Based on the graph we can see that solubility slowly increases as temperature increases. We could estimate the value of solubility for any temperature within the range given and for example at a temperature of 50 °C the solubility should be close to 50 g/100g of water.

1.2 Units of Measurements and systems of units

You probably heard the term liter, kilogram, or meter. These are units of measurement. Units can be classified into different *systems of units*. For example, the unit *meter* belongs to a different system than the unit *mile*. In particular, here we will address three main systems: the English System, the Metric System, and the International system. The *Metric System* (MS) is used by scientists throughout the world and is the most common measuring system based on the meter. The English system is mostly used in the US. The *International System of Units* (SI) adopted the metric system in 1960 to provide additional uniformity for units used in the sciences. Table 1.1 summarizes some of the fundamental units for the three systems. This chapter will be mostly based on the SI units. In the following, we will introduce some common units.

`Length` What is your height? The length refers to distance and both the metric and SI unit of length is the meter (m). A smaller unit of length would be the centimeter (cm) that is commonly used in chemistry. The most important units of length are meters, inches, and miles.

`Mass` What is your weight? The mass of an object is a measure of the quantity of material it contains. You may be more familiar with the term weight rather than mass. However, mass and weight are not the same, as weight is a measure of the gravitational pull on an object. It differs depending on your location on the earth–in particular the height of your location. In the metric system, the unit of mass is the gram (g). The SI unit of mass, the kilogram (kg), is used for larger masses such as body weight. A pound, lb, is another unit of mass. The most important units of mass are g, kg, and lb.

`Temperature` How is the weather today? Is it cold or hot? You use a thermometer to measure temperature and for example assess how hot an object is, or how cold it is outside, or perhaps to determine if you have a fever. Temperature tells us how hot or cold an object is. Temperature can be measured in numerous units such as Celsius (°C), Fahrenheit (°F), or kelvins (K).

`Time` How long is your commute to work? It might take you hours to go to work, or maybe minutes. You probably think of time as years, days, minutes, or seconds. Of all these units, the International System of units (SI, abbreviated from the French Systeme international) uses seconds (s) to measure time. Still, time can be measured in s, min, or h and during this chapter, we will learn how to convert units of time.

`Volume` How much milk do you usually buy? Maybe a gallon. Volume is the amount of space that a substance occupies. A liter (L), not a fundamental but a derived SI unit, is commonly used to measure volume. The milliliter (mL) is more convenient for measuring smaller volumes of fluids in hospitals and laboratories. Gallon is still used in everyday life. L, mL, and gallon are units of volume. Units of volume are in general cubic units, so for example one liter is the same as one dm^3. We will cover cubit units further in this chapter.

Chemical laboratory work commonly requires the measurement of volume. There are two main types of glassware used to measure volume in a chemistry lab: graduated tools and volumetric tools. Volumetric pipets, flasks, and burets are the most accurate; the glassware makers calibrate these to a high level of accuracy, usually measured in terms of tolerance, which is the uncertainty in a measurement made with the glassware. Class A volumetric glassware has a lower tolerance than Class B; for class A, the tolerance can be as low as 0.08 ml for a 100 ml flask or pipet. Generally, measurements with class A volumetric glassware can be considered reliable to two places after the decimal point. Graduated cylinders, beakers, and Erlenmeyer flasks have less accuracy than volumetric glassware. Graduated cylinders can generally be considered reliable to within 1 percent. Beakers and Erlenmeyer flasks should not be used to measure volume unless you need only a very crude estimate because their accuracy for volume measurements is so poor. They can hold a much larger volume than any of the other types of glassware, however, which makes them useful for mixing solutions.

Concentration Even though we will devote a whole chapter to solutions and concentration, it felt important to introduce here the unit molarity. In chemistry, the unit molarity (M) refers to the concentration of a solution. That is the larger this number, the larger molarity, and the more concentrated a solution will be. In other words, there will be more substance in the solution.

Sample Problem 2

State the type of measurement indicated in each of the following:

(a) 1ft(foot) (b) 20Kg (c) 3L (d) 300K

SOLUTION

(a) length; (b) mass; (c) volume; (d) temperature;

STUDY CHECK

State the type of measurement indicated in each of the following: (a) 800°F (b) $1m^3$ (c) 3m (d) 67s

Table 1.1 Different unit systems

Measuremts	Metric System	International System (SI)	English System
Length	Meter (m)	Meter (m)	Foot (ft)
Mass	Gram (g)	Kilogram (kg)	Pound (lb)
Time	Second (s)	Second (s)	Second (s)
Temperature	Celsius (°C)	Kelvin (K)	Fahrenheit (°F)
Volume*	Liter (L)	Cubic meter (m^3)	Gallon (gal)
Ammount of substance	Mole (mol)	Mole (mol)	Mole (mol)
Electric current	Ampere (A)	Ampere (A)	Ampere (A)

*Not a fundamental unit

1.3 Significant Figures

Exact numbers result from counting. For example, think about how many eggs are there in your refrigerator, there might be three and this number is exact. Differently, numbers that result from a measurement are called measured values and they are subject to uncertainty– in other words error. For example, if you weigh a single egg on a scale depending on the type of scale you used and the person who carries out the measurement, you will measure 70g or 71g, or maybe 70.8g. The mass of an egg is a measured property and hence some of the digits of the measurement are uncertain. The goal of this section is, given a value, to calculate the number of significant figures of a number (we will refer to significant figures as SF, or SFs). Another goal is to estimate significant figures in the calculation to express the result with the right number of digits and significant figures.

`Measured numbers` *Measured numbers* result from measuring a property such as the weight or length of an object. Those measurements result from using a measuring device such as a scale or a ruler, for example. Imagine we want to measure the length of both objects presented in Figure 1.2. The metric rules presented have a set of marked divisions which determine the number of figures given by the measurement. For example, the ruler on the left has 1cm and 0.1cm divisions, whereas the rule on the right only has 1cm divisions, hence giving fewer figures.

Let us estimate the length of the object on the right. The end of the object on the right is located between 0cm and 1cm, therefore its length is less than 1cm. Still, we can estimate an extra digit by dividing the space between the lines. Still, this last *estimated digit* might differ from person to person. The final measurement would be 0.8cm. However, some people would read the length as 0.7cm whereas others 0.9cm. Let us now estimate the length of the object on the left. The end of the object on the right is located between 3.1cm and 3.2cm, therefore its length is less than 3.2cm. We can estimate an extra digit as well, giving a final measurement of 3.15cm.

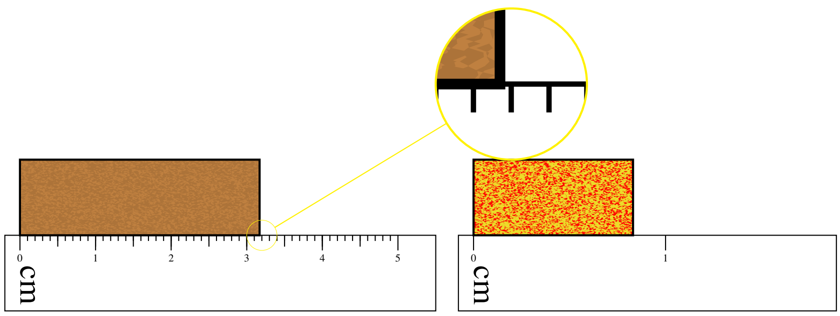


Figure 1.2 Some metric scales with two objects of different lengths. Measurements are (left) 3.15cm (right) and 0.8cm.

`Reading menisci` Reading a liquid meniscus is similar to reading any measuring scale. There are two types of menisci (see Figure 1.3). A concave meniscus, which is what you normally will see, occurs when the molecules of the liquid are attracted to those of the container. This occurs with water and a glass tube. A convex meniscus occurs when the molecules have a stronger attraction to each other than to the container, as with mercury and glass. If the meniscus is concave, read at the lowest level of the curve. If the meniscus is convex, take your measurement at the highest point of the curve. Let us read the menisci from the image below. Readings are 16.0mL (left), 8.5mL (center left), 18.0mL (center right), and 18.5mL (right).

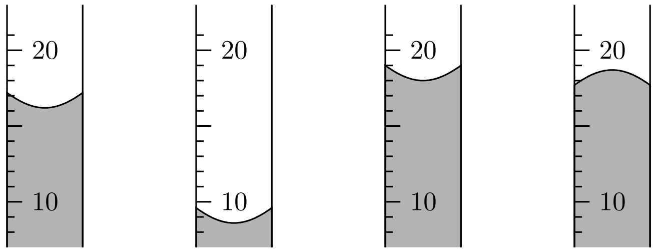


Figure 1.3 Some volumetric measurements in mL are presented in two types of meniscuses. The Left, center-right, and center-left are concave meniscuses, whereas the right image presents a convex meniscus. Readings are 16.0mL (left), 8.5mL (center left), 18.0mL (center right), and 18.5mL (right).

Exact numbers Exact numbers are numbers obtained by counting and not by measuring or obtained by a relationship that compared two units in the same measuring system. For example, the number of students in a class is exact as we need to count to get this number. Similarly, the number of grams in a kilogram, a thousand, is exact as the relationship between kilogram and gram is exact. Exact numbers do not have significant figures and do not limit the number of figures in a calculation.

Significant figures of numbers In general, all numbers different than zero are significant and for example, the number 123 has three significant figures. Similarly, the number 45 has two significant figures. Zeros are also significant except when:

Exception 1 *A zero is not significant when placed at the beginning of a decimal number.* For example, the number 0.123 has three significant figures, as the first zero is not significant. Similarly, the number 0.002340 has four significant figures as the first three zeros are not significant but the last zero it is. Mind the rule affects only the zeros at the beginning. A final example:

$$0.032 \ \ \text{(2SF)}$$

Exception 2 *A zero is not significant when used as a placeholder in a number without a decimal point.* For example, the number 1000 has only one significant figure, and the number 3400 has two. Let us consider more examples. The number 120 has two significant figures, as according to the second rule the last zero is not significant. Differently, the number 1203 has four significant figures, as the zero in between two numbers is not affected by either the first or the second rule. A final example,

$$3200 \ \ \text{(2SF)}$$

Exception 3 *A zero in a number expressed in scientific notation is significant.* For example, the zero in 3.0×10^{-2} is significant, and the number has 2SFs. A final example:

$$3.2020 \times 10^{2} \ \ \text{(5SF)}$$

Sample Problem 3

Indicate the number of significant figures in the following numbers: 123, 4567, 1200, 340, 0.001, 0.023 and 0.0405.

SOLUTION

123 has three significant figures, whereas 4567 has four SF. 1200 has only 2SF

as the last two zeros are not significant, and 340 has only 2SF as the last zero is not significant. 0.001 has only one significant figure as the first 3 zeros are not significant and 0.023 has only two SFs. Finally, 0.0405 has threee SFs as the first two zeros are not significant but the zero between 4 and 5 is indeed significant.

STUDY CHECK

Indicate the number of significant figures (SFs) in the following numbers: 4560, 0.123, 1000 and 0.0030.

Significant figures in calculations Two different rules allow you to express the result of calculations with the correct number of figures.

Rule 1 ($+\ -$) *For additions or subtractions, the results has the same number of decimal places as the number with the least decimal places in the calculation.* For example:

$$34.3451 + 34.5 = 68.8\ (+\ -\text{ less decimals})$$

If you add $34.3451 + 34.5$ you will obtain 68.8451, however, as 34.3451 has four decimal places (4DP) and 34.5 has one decimal place (1DP), the result of adding both numbers will have to have only one decimal place, therefore 68.8451 needs to be rounded to 68.8 (1DP). Overall, we have:

$$34.3451\text{ (4DP) } + 34.5\text{ (1DP) } = 68.8\text{ (1DP)}$$

Rule 2 ($\times\ \div$) *For multiplications and divisions, the number of significant figures of the result should be the same as the least number of significant figures involved.* For example, if you carry the following multiplication:

$$4500 \times 342 = 1500000\ (\times\ \div\text{ less SFs})$$

the number 4500 (2SF) has two significant figures, whereas the number 342 (3SF) has three significant figures. If we multiply both numbers the results should contain just two significant figures. The result of multiplying 4500×342 is 1539000 (4SF), however, this number needs to be rounded into two significant figures into 1500000 (2SF). Overall we have:

$$4500\text{ (2SF) } \times 342\text{ (3SF) } = 1500000\text{ (2SF)}$$

Sometimes we will have to add significant zeros in order to present the final result of a calculation with the correct number of digits. For example:

$$8.00\text{ (3SF) } \div 2.00\text{ (3SF) } = 4\text{ (shows in calculator) } = 4.00\text{ (3SF)}$$

Sample Problem 4

Do the following calculation with the correct number of figures.

$$\frac{88.5 - 87.57}{345.13 \times 100}$$

SOLUTION

We will analyze each number indicating the number of SF and Digits (DP): 88.5(3SF, 1DP), 87.57(4SF, 2DP), 345.13(6SF, 2DP) and 100(1SF, 0DP). The result of doing the addition needs to be rounded to one single decimal place: $88.5 - 87.57 = 0.93 \simeq 0.9$. After that we have only multiplications and divisions and hence we will now focus on the number of SFs:

$$\frac{0.9\ (1\text{SF})}{345.13\ (5\text{SF})\ \times 100\ (1\text{SF})}$$

The result of this operation needs to be rounded to one SF:

$$\frac{0.9}{345.13 \times 100} = 2.6077 \times 10^{-5} \simeq 3 \times 10^{5}\ (1\text{SF})$$

STUDY CHECK
Do the following calculation with the correct number of figures: $(24.56 + 2.433) \times 0.013$

Rounding The following rules indicate how to round numbers:
Rule 1 If the digit to be removed is less than 5 then the preceding digit stays the same. For example, 1.123 rounds to 1.12.
Rule 2 If the digit to be removed is more or equal to 5 then the preceding digit is increased by one. For example, 1.126 rounds to 1.13
Rule 3 When rounding to a specific number of significant figures we need to look only to the first number to the right of the last significant figure. For example, 1.126 rounds to two SF as 1.1
Now, let us analyze a few use cases. Imagine we need to round the number 1234cm to two SF. The results would be 1200cm. Similarly, imagine we need to round the number 0.01264cm to two SF. The results would be 0.013cm.

Table 1.2 Different prefixes

Prefix	Symbol	Meaning	Value
exa	E	1000000000000000000	1×10^{18}
peta	P	1000000000000000	1×10^{15}
tera	T	1000000000000	1×10^{12}
giga	G	1000000000	1×10^{9}
mega	M	1000000	1×10^{6}
kilo	k	1000	1×10^{3}
hecto	h	100	1×10^{2}
deca	da	10	1×10^{1}
–	–	1	1×10^{0}
deci	d	0.1	1×10^{-1}
centi	c	0.01	1×10^{-2}
milli	m	0.001	1×10^{-3}
micro	μ	0.00001	1×10^{-6}
nano	n	0.000000001	1×10^{-9}
pico	p	0.000000000001	1×10^{-12}
femto	f	0.000000000000001	1×10^{-15}
atto	a	0.000000000000000001	1×10^{-18}

1.4 Prefixes & Conversion Factors

Let's consider the following measurements: 1 **k**m, 2 **c**m and 3 m that can be read as one kilometer, two centimeters, and three meters. The word kilo (**k**) and centi (**c**) are called prefixed whereas meter (m) is a simple unit. Table 1.2 lists some of the metric prefixes, their symbols, and their decimal values. A kilometer is larger than a meter, whereas a centimeter is smaller than a meter. Prefixes such as kilo or centi are attached to units to make numbers more manageable. For example, the radius of the earth is 6356 km, and this number is easier to handle than 6356000m. At the same time, we can attach any prefix to different units. Hence, we can talk about a centimeter (**c**m) but also about a centisecond (**c**s) or centiliter (**c**L). All these units have the same prefix.

`How to identify prefixes?` Look for example at the measurement 2 cm. Centi (c) is the prefix and means 1×10^{-2} and meter (m) is the unit that refers to length. Another example, is 7 kg read as seven kilograms. Kilo (k) is the prefix and means 1×10^{3}, whereas gram (g) is the unit that refers to mass. The prefix refers to the first letter whereas the unit refers to the last letter.
A unit with a prefix can be bigger or smaller than the plain unit–this is the unit without a prefix–, depending on the prefix. The following prefixes reduce the unit: deci, centi, milli, micro, nano, pico, and femto. For example, a fs (femtosecond) is smaller than a s (second). Differently, the following prefixes increase the unit: Tera, Giga, Mega. For example, a Tb (terabyte) is larger than a b (byte). A byte is a unit used in computer science.

`How to write unit equalities and conversion factors` Unit equalities are simple expressions that relate a unit with a prefix. For example, one centimeter (cm) is 1×10^{-2}m. Hence we can write this as unit equality:

$$\boxed{1cm = 1 \times 10^{-2}m} \text{ unit equality}$$

Let's compare cm and m. The first, cm, is a unit with a prefix, whereas m is simply a unit of length without a prefix. To know how many m are there in a cm we need to write down a conversion factor. Think about prefixes as synonymous with a number. In this way, centi stands for 1×10^{-2}, so

$$\boxed{\frac{1cm}{1 \times 10^{-2}m} \text{ or } \frac{1 \times 10^{-2}m}{1cm}} \text{ conversion factor}$$

The relationship above is also called a *unit factor* as both ratios are equal to one.

Sample Problem 5

Complete each of the following equalities and conversion factors:

(a) 1dm=______m

(b) 1km=______m

(c) $\frac{1nm}{\_\_\_\_\_\_m}$

(d) $\frac{\_\_\_\_\_\_m}{1cm}$

SOLUTION

(a) 1dm=$1 \times 10^{-1}m$; (b) 1km=$1 \times 10^{3}m$; (c) $\frac{1nm}{1 \times 10^{-9}m}$; (d) $\frac{1 \times 10^{-2}m}{1cm}$;

STUDY CHECK

Second is a unit of time. Complete each of the following equalities and conversion factors involving seconds:

(a) 1cs=______s (b) $\frac{\_\_\_\_\_\_s}{1Ts}$ (c) $\frac{\_\_\_\_\_\_s}{1Ms}$

1.5 Using Conversion Factors

Unit equalities in the form of conversion factors are used to convert one unit into another. Sometimes one wants to get rid of a prefix, such as when we transform centimeter (cm) into meter (m). Sometimes, one wants to convert a prefix into another prefix. An example would be converting centimeters (cm) to millimeters (mm). Let's work on some examples.

Removing or adding prefixes Imagine that you need to remove a prefix from a unit, and convert 3 km (we will call this one the original unit) into meters (this is the final unit). First, you would need the conversion factor corresponding to the prefix (centi) from Table 1.2. Then you need to arrange the conversion factor by placing the prefix at the bottom of the fraction. This will cancel out the prefix in the original unit and the bottom part of the conversion factor, hence leaving the final unit on top of the conversion factor. The arrangement would be:

$$3\cancel{km} \times \frac{1 \times 10^3 m}{1\cancel{km}} = 3000m$$

Imagine now that you need to add a prefix into a unit, and convert 4000 m in km. The same would apply for this case, but now you will have to arrange the conversion factor so that the prefix is on the top:

$$4000\cancel{m} \times \frac{1 \text{ km}}{1 \times 10^3 \cancel{m}} = 4km$$

Sample Problem 6

The length of a textbook page is 20cm. Convert 20cm to meters, expressing the result in scientific notation.

SOLUTION

In order to convert 20cm into meters, we need to remove the prefix (centi) leaving the unit (meter) without any prefix. We will use the conversion factor that relates m to cm: $\frac{1 \times 10^{-2} m}{1cm}$ or $\frac{1cm}{1 \times 10^{-2} m}$. We will arrange the conversion factor so that cm cancels giving m and hence we will use $\frac{1 \times 10^{-2} m}{1cm}$:

$$20\cancel{cm} \times \frac{1 \times 10^{-2} m}{1\cancel{cm}} = 2 \times 10^{-1} m$$

The original units and on the bottom of the conversion factor cancel and we get meters, the final unit.

STUDY CHECK

Convert 100m to km, expressing the result in scientific notation.

Switching prefixes To switch a prefix into another prefix, such as transforming 30 millimeters (30 mm) into centimeters (cm), you will need two different conversion factors: the first conversion factor will remove the original unit (mm) introducing an intermediate unit, meters (m), whereas the second conversion factor will remove the intermediate meter and introduce the final unit (cm). You will get the conversion factors from Table 1.2. You will arrange the first conversion factor so that the original unit cancels out with the bottom of the first conversion factor, giving you an intermediate unit. You will arrange the second conversion factor so that the intermediate unit cancels out with the bottom of the second conversion factor giving the final unit. For this example:

$$30\cancel{mm} \times \frac{1 \times 10^{-3}\cancel{m}}{1\cancel{mm}} \times \frac{1cm}{1 \times 10^{-2}\cancel{m}} = 3cm$$

Sample Problem 7

The length of a textbook page is 20cm. How many mm correspond this length, expressing the result in scientific notation.

SOLUTION

We want to convert 20 cm into mm, that is, we are switching prefixed. In order to do this, you need two conversion factors: $\frac{1 \times 10^{-2}m}{1cm}$ and $\frac{1 \times 10^{-3}m}{1mm}$. You will have to arrange the number (20cm) and the two conversion factors in the following form:

$$20\cancel{cm} \times \frac{1 \times 10^{-2}\cancel{m}}{1\cancel{cm}} \times \frac{1mm}{1 \times 10^{-3}\cancel{m}} = 2 \times 10^{2}mm$$

STUDY CHECK

Convert 100mm to km, expressing the result in scientific notation.

1.6 Units of volume and area

How big is your apartment? You might be living in a $750ft^2$ loft in Brooklyn or a larger house Upstate. Often times we encounter cubic or square units such as cubic centimeters (cm^3) or square feet (ft^2). The equivalencies for cubic or square units should take into account the unit power (power of two or power of three). If $1cm = 1 \times 10^{-2}m$, for square units the relation should be squared and $1cm^2 = 1 \times (10^{-2})^2m^2 = 1 \times 10^{-4}m^2$. Another example, for the case of mm and mm^3:

$$\boxed{\frac{1mm}{1 \times 10^{-3}m} \quad \text{and} \quad \frac{1mm^3}{1 \times 10^{-9}m^3}}$$

Let us work on an example in which we want to convert $30m^2$ into m^2:

$$30\cancel{m^2} \times \frac{1cm^2}{1 \times 10^{-4}\cancel{m^2}} = 3 \times 10^{5}cm^2$$

Sample Problem 8

How many m^2 is $20cm^2$, expressing the result in scientific notation.

SOLUTION

In order to convert $20cm^2$ to square meters, we need to remove the centi prefix and that will give us the unit square meter without any prefix. We will use the conversion factor that relates m^2 to cm^2: $\frac{1 \times 10^{-4}m^2}{1cm^2}$ or $\frac{1cm^2}{1 \times 10^{-4}m^2}$.

$$20\cancel{cm^2} \times \frac{1 \times 10^{-4}m^2}{1\cancel{cm^2}} = 2 \times 10^{-3}m^2$$

STUDY CHECK

Convert $100m^3$ to dm^3, expressing the result in scientific notation.

Table 1.3 Table containing some common unit equalities

Unit	Equality
Inches (in)-centimeters (cm)	$2.54^{\dagger}$ cm = 1 in
miles (mi)-meters (m)	1 mi = 1609.34m
minutes (min)-hours (h)	60 min = 1 h
minutes (min)-seconds (s)	60 s = 1 min
pound (lb)-grams (g)	454 g = 1 lb
cubic centimeter (cm^3)-mililiters (mL)	1 mL = 1cm^3
Liter (L)-cubic decimeters (dm^3)	1 L = 1dm^3
drops-mililiters* (mL)	1 mL = 15 drops

* There are several definitions of a drop

† the number is exact

`Liters and milliliters` Units such as L or mL are units of volume. As volume is a three-dimensional property, those units somehow have to be related to the units of length. One liter is the same as one dm^3 and one ml is the same as one cm^3 (See Figure 1.4). In the allied health field, the units mL are also written as cc as in cubic centiliters.

$$\boxed{1L = 1dm^3 \text{ and } 1mL = 1cm^3(cc)}$$

Let us work on an example in which we want to convert $30cm^3$ into L:

$$30\cancel{cm^3} \times \frac{1\cancel{mL}}{1\cancel{cm^3}} \times \frac{1 \times 10^{-3}L}{1\cancel{mL}} = 3 \times 10^{-2}L$$

Sample Problem 9

Convert 30 m^3 into L, expressing the result in scientific notation.

SOLUTION

In order to convert m^3 into L we just need to remember that the L actually refers to dm^3, therefore is connected to meter. We will first convert m^3 into dm^3 and then dm^3 into L.

$$30\cancel{m^3} \times \frac{1\cancel{dm^3}}{1 \times 10^{-3}\cancel{m^3}} \times \frac{1L}{1\cancel{dm^3}} = 3 \times 10^4 L$$

STUDY CHECK

Convert 40L to cm^3, expressing the result in scientific notation.

1.7 Using other equalities

How many hours are 300 minutes, or how many centimeters is 2 inches? Some of the units conversion is not based on a power of ten relationships and do not contain prefixes such as kilo or centi. Table 1.3 lists some of the common equalities that can be easily converted into conversion factors. As an example, the unit equivalency between hours and minutes is $60min = 1h$ and the conversion factor would be $\frac{60min}{1h}$ or $\frac{1h}{60min}$.

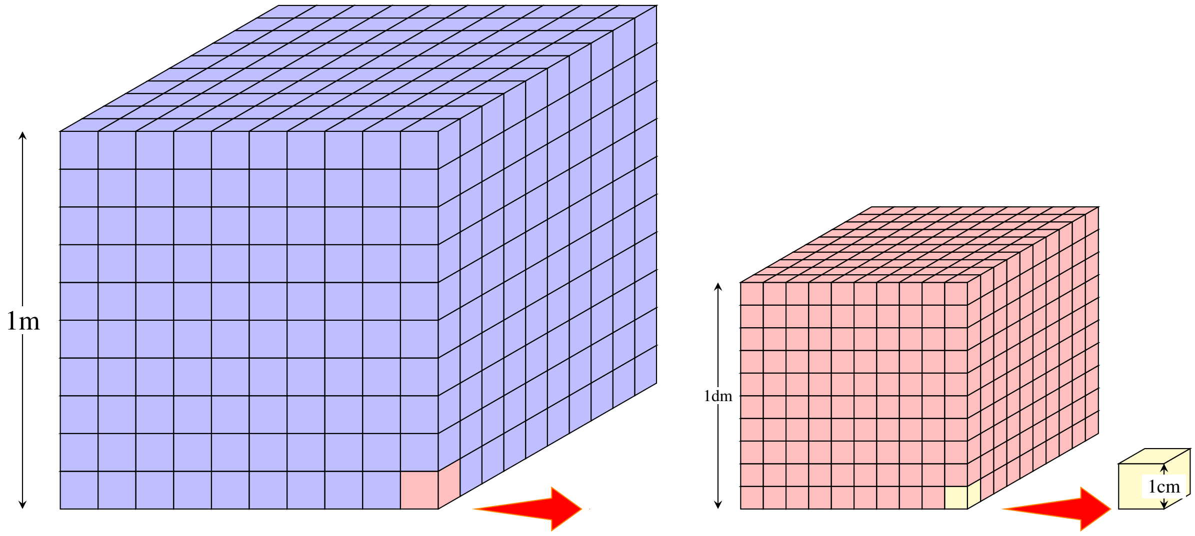


Figure 1.4 The left cube with a side of 1m has a volume of $1m^3$. The central cube with a side of 1dm has a volume of $1dm^3$, that is 1L. The right cube with a side of 1cm has a volume of $1cm^3$, that is 1mL.

Sample Problem 10

Convert 20 in to cm, expressing the result in scientific notation.

SOLUTION

We want to convert 20 inches into centimeters. The relationship between Inch and centimeter is given in Table 1.3. In order to do this, you need the conversion factor: $\frac{1in}{2.54cm}$ or $\frac{2.54cm}{1in}$. You will have to arrange the number (20 in) and the conversion factor in the following form:

$$20\cancel{in} \times \frac{2.54cm}{1\cancel{in}} = 5.080 \times 10^1 cm$$

STUDY CHECK

Convert 200mL to drops, expressing the result in scientific notation.

1.8 Measurements and uncertainty

Most chemistry experiments require the measurement of a property (mass, volume, temperature, color...), and the validity of those experiments will depend on the reliability of each measurement. The reliability of a measurement is usually considered in terms of its *accuracy* and its *precision*. At the same time, when experimenting oftentimes, one needs to repeat measurements. The results need to take into account the average measurement as well as the standard derivation of the series of results.

`Uncertain and certain digits: how to report measured numbers`

Consider the volume as measured in the glassware on the left side of Figure 1.5. Some would say the liquid meniscus occurs at 1.52mL. Mind we need to estimate the last number by interpolating between the 0.1mL marks. As the last digit of the number associated with the volume measurement is estimated, another person could measure the volume as 1.53mL. The table below indicates the measurements of five different people.

Person	1	2	3	4
Volume(mL)	1.52	1.53	1.51	1.57

All measurements have in common the 1.5 part. The digit 5 is called a certain digit. The digit to the right of 5 is called the uncertain digit, as in a measurement it needs to be estimated. When reporting a measurement, we need to report up to the uncertain digit. These numbers on measurement are called *significant figures*. It is important to understand that all measured properties are subject to uncertainty. Uncertainty depends on one hand on the measuring device. The uncertainty on the left meniscus in Figure 1.5 is in the hundreds of the mL, whereas on the right meniscus occurs in the tenths of mL. Uncertainty also depends on the measurement process. As the uncertain digit needs to be reported in every measurement, the uncertainty on the last number on measurement is normally assumed to be ± 1. For example, the right meniscus should be reported as $1.52 \ \pm 0.01$mL

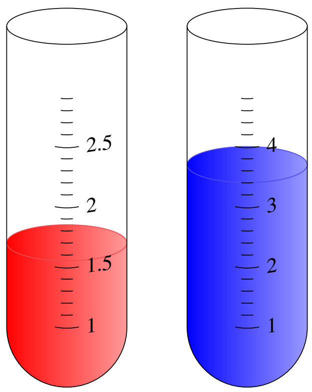


Figure 1.5 Two measuring devices. (Left side) The measuring device gives volumes to the hundredth place. (Right side)The measuring device gives volumes to the tenth place.

`Precision and accuracy` The terms precision and accuracy are two different terms often used to describe the reliability of measurement. Accuracy refers to the degree of agreement between the measured value and the true value, the real value of the property measured. Measurements that closely agree with the true value are accurate. Precision refers to the agreement between several measurements. Two measurements are accurate when they are close, independently of the true value of the property measured. Two different types of errors occur during a measurement. On one hand, random error refers to the fact that measured values can be above or below the true value. On the other hand, systematic error is an error that occurs each time in the same direction. Figure 1.6 displays accurate and inaccurate measurements as well as precise and imprecise measurements.

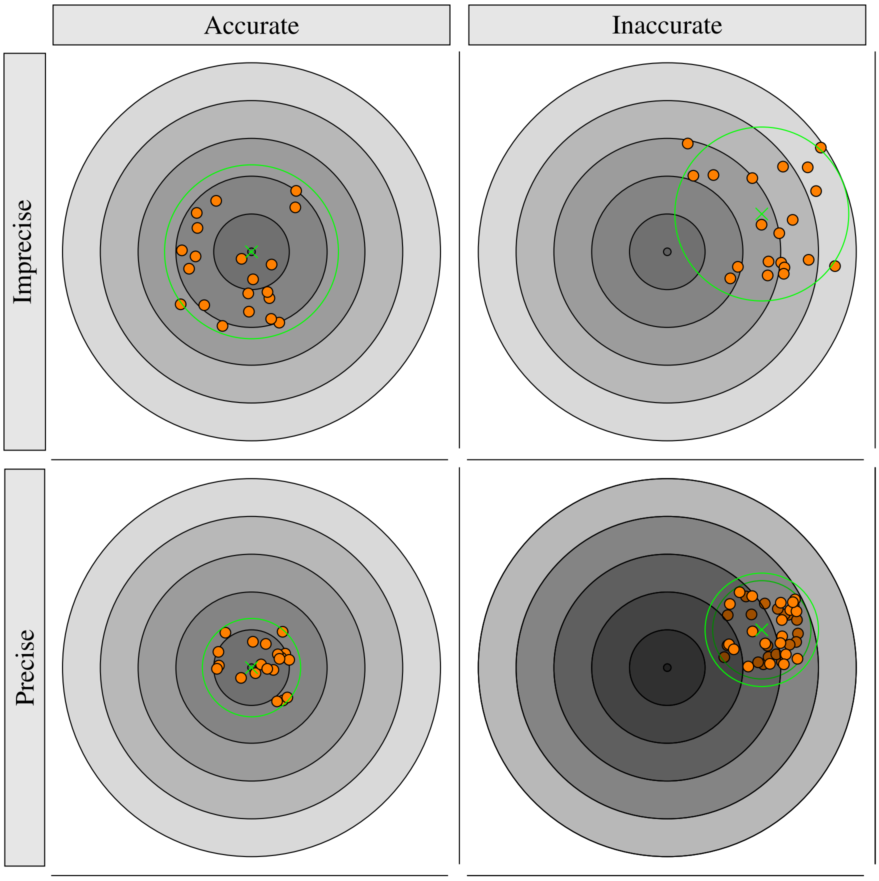


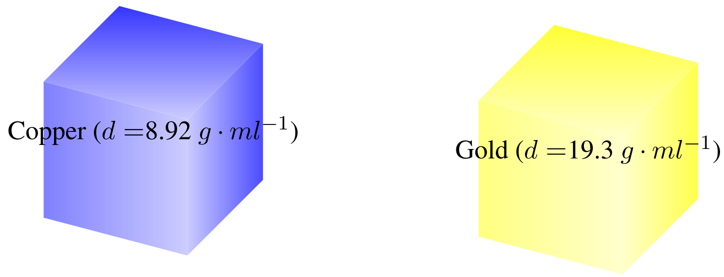
Figure 1.6 (Top Left) An accurate but imprecise measurement (Top Right) An inaccurate and imprecise measurement (Top Left) An accurate and precise measurement (Bottom Right) An inaccurate but precise measurement

1.9 Density

Density refers to the mass of a substance with respect to its volume. This is an unique property for each substance. Table 1.4 reports the density of numerous substances. Indeed, density is often used as an identification tag. The formula for density is

$$\boxed{\text{Density} = \frac{\text{Mass of substance}}{\text{Volume of substance}}} \tag{1.1}$$

For example, the density for copper is 8.92 $g \cdot ml^{-1}$ and for gold is 19.3 $g \cdot ml^{-1}$. By measuring density only, you would be able to differentiate copper than gold. The larger density the more compact is an object and that means the more mass per volume it has. At the same time, for the same volume, the larger density the larger the mass of the metal.



`Density and mixing` A small piece of ice will float on the water. The reason for that is density: the density of ice (0.9g/mL) is smaller than the density of water (1.0g/mL) and hence ice will stay on top of the water. Objects with a density larger than 1 g/mL will sink whereas objects with a density smaller than this value will float. Figure 1.7 showcases how objects with density larger than water will sink whereas objects with a smaller density will float. If you add a drop of vegetable oil to a glass of water, the drop will float. This is because the density of oil is smaller than 1g/mL.

Table 1.4 Density of some common substances at 273.15 K and 100 kPa

Substance	Density (g/mL)	Physical State
Helium	0.2	gas
Hydrogen	0.1	gas
Water	1.0	Liquid
Cooking oil	0.9	Liquid
Mercury	13.5	Liquid
Tetrachloroethene	1.6	Liquid
Gold	19.3	solid
Plastics	1.2	solid
Ice	0.916*	solid

*Ice is given at T <273.15 K

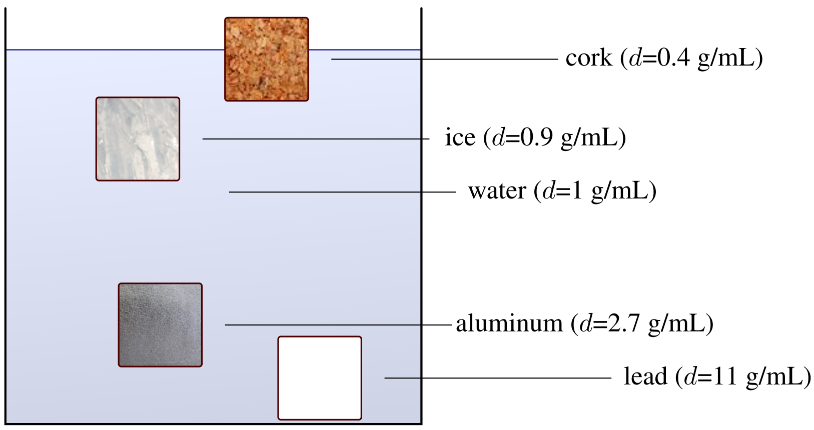


Figure 1.7 Objects with a larger density than water will sink whereas objects with a smaller density will float.

Sample Problem 11

In the figure, we mixed three liquids of density: A (0.5 g/mL), B(2 g/mL) and C(1 g/mL). Identify each liquid.

SOLUTION

The heavier the liquid, that is the larger density, the lower the liquid will arrange in the mixture. From top to bottom we have A, C and B.

STUDY CHECK

Indicate the order that the following invisible liquids will appear in a cylinder when mixed: (a) benzene(0.87 g/mL) and (b) water(1 g/mL)

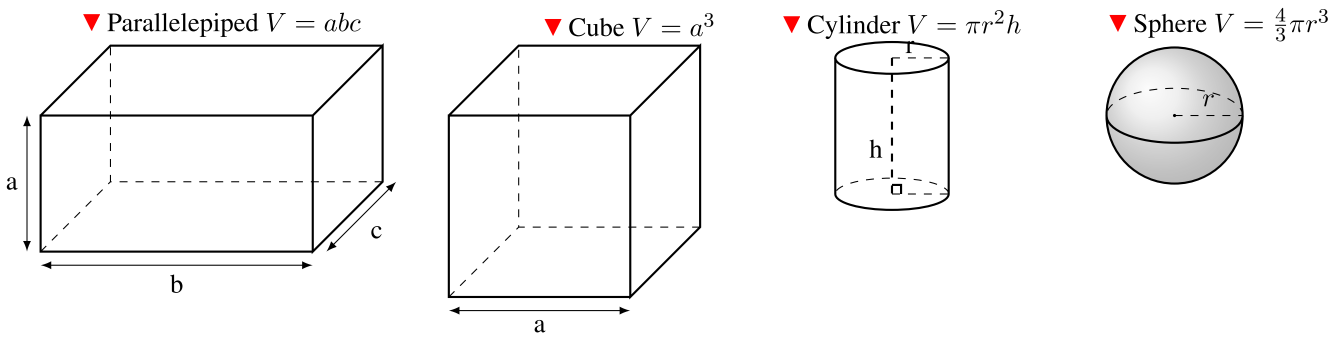
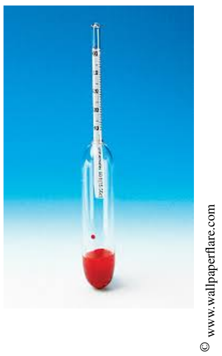


Figure 1.8 Volume of some objects

▼A hydrometer



▼Dipstick used to measure specific gravity

`Density and the volume of objects` Density depends on volume and in particular the larger volume the smaller density. Figure 1.8 displays the formulas to calculate the volume for some common objects, like a sphere or a cube. For example, the radius of a sphere with density d and mass m corresponds to $\sqrt[3]{3m/4d \cdot \pi}$, and the side of a cube with density d and mass m corresponds to $\sqrt[3]{m/d}$. The density of liquids results from measuring the mass of a given volume of the liquid. Differently, density is harder to obtain for solids. For metals, we can calculate density by the immersion method: when a metal is immersed in water, the water rises. This increase in volume corresponds to the volume of the solid. This way, density results from the direct measurement of mass and the measurement of volume by displacement.

The following example demonstrates density calculation with the immersion method.

Sample Problem 12

After adding a 30g object into a cylinder filled of water, the level of water rises from 60mL to 90mL. Calculate the density of the object.

SOLUTION

Density is mass over volume. The mass of the object is 30g and its volume is (90-60)mL that is 30mL. Hence: $d = 30g/30mL = 1g/mL$.

STUDY CHECK

A lead weight used in the belt of a scuba diver has a mass of 226 g. When the

> weight is placed in a graduated cylinder containing 200.0 mL of water, the water level rises to 220.0 mL. What is the density of the lead weight (g/mL)?

`Specific gravity` The specific gravity (ρ) of a substance is the ratio between its density and the density of a reference, normally water. It is simply calculated by dividing the density of the substance and the density of water (1g/mL at room temperature).

$$\boxed{\rho = \frac{\text{density of substance}}{\text{density of water}}} \tag{1.2}$$

A substance with a specific gravity of 1 has a density of 1g/mL. This is a unitless property that can be measured with an instrument called a hydrometer. For example, the specific density of the urine in the body is used to identify diabetes or kidney malfunctioning.

CHAPTER 1

Math skills

1.1 Solve for x in the following algebraic equations: (a) $\frac{3x+1}{2} = 2$ (b) $\frac{2x-1}{3} = 2$ (c) $\frac{2}{3} = \frac{1}{3x}$

1.2 Solve for x in the following algebraic equations: (a) $3x + 1 = 5$ (b) $2x - 1 = 5$ (c) $\frac{3}{2} = \frac{1}{2x}$

1.3 Compute the following calculations involving scientific notation: (a) $\frac{6.5\times10^3}{3\times10^2}$ (b) $\frac{6.1\times10^{-3}}{3\times10^4}$ (c) $\frac{1.3\times10^{-3}}{2.5\times10^{-2}}$

1.4 Compute the following calculations involving scientific notation: (a) $\frac{2.4\times10^{-3}}{(5)(4.6\times10^{-6})}$ (b) $\frac{1}{(3)(4\times10^{-1})}$ (c) $\frac{1}{(4)(1\times10^{-4})}$

Units of Measurements and systems of units

1.5 Indicate the magnitude measured in the following measurements: (a) 2 L (b) 5 cm

1.6 Indicate the magnitude measured in the following measurements: (a) 100 Kg (b) 10h

1.7 Answer the following questions: (a) Indicate the metric base unit for mass (b) Indicate the metric base unit for time (c) Indicate metric base unit for volume

1.8 Answer the following questions: (a) Indicate basic unit of mass in the SI (b) What magnitude measures the amount of space occupied by a substance

Significant Figures

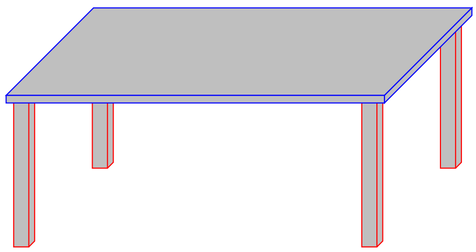
1.9 Indicate if the following statements represent measured or exact numbers: (a) The number of apples in a bag (b) The weight of an apple (c) The number of cm in a inch

1.10 Indicate if the following statements represent measured or exact numbers: (a) The number of students in a classroom (b) The height of a student

1.11 Identify the following numbers as measured or exact: (a) The number of students enrolled in a class is 25 (b) The temperature of a sample is 25$^\circ C$ (c) The mass of an object is 10 Kg (d) 1000grams is 1Kg

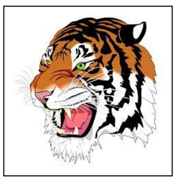
1.12 Identify the following numbers as measured or exact: (a) The length of an object is 4 cm (b) A chair has four legs (c) The density of a solution is 0.6 g/mL (d) One meter equals 100 cm

1.13 Identify the following numbers as measured or exact given the image below:



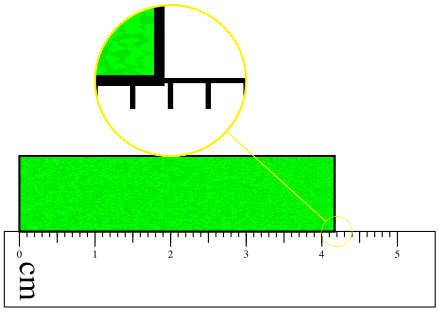
(a) The height of the table (b) The number of table legs (c) The width of the table (d) The number of tables in the image

1.14 Identify the following numbers as measured or exact given the image below:

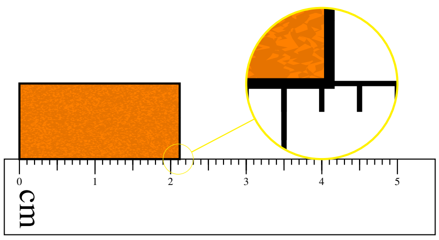


(a) The weight of the tiger (b) The number of ears of the tiger (c) The height of the tiger (d) The number of whiskers

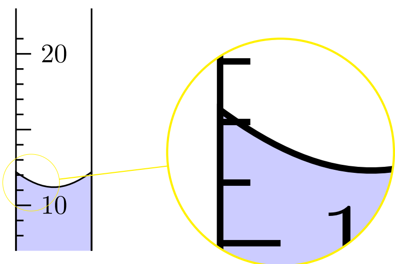
1.15 Give the following measurement with the correct number of digits, while indicating the number of significant figures of the measurement and the estimated digit.



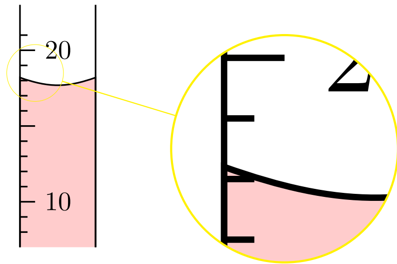
1.16 Give the following measurement with the correct number of digits, while indicating the number of significant figures of the measurement and the estimated digit.



1.17 Give the following measurement in mL with the correct number of digits, while indicating the number of significant figures of the measurement and the estimated digit.



1.18 Give the following measurement in mL with the correct number of digits, while indicating the number of significant figures of the measurement and the estimated digit.



1.19 Indicate the number of SFs. (a) 0.0032 m (b) 5100 m (c) 0.510 m (d) 0.0051 m (e) 500 m (f) 45.030

1.20 Which of the following measurents contains the designated CORRECT number of significant figures? (a) 0.05600 cm (5 SF) (b) 0.03040 cm (3 SF) (c) 456 000 cm (3 SF) (d) 1.304 cm (2 SF) (e) 3.12050 cm (4 SF)

1.21 Carry the following calculations with the correct number of digits or significant figures:

(a) $0.2301 + 0.123$ (c) $88.1 - 87.57$

(b) $0.2301 - 1.12$ (d) $24.56 + 2.4$

1.22 Carry the following calculations with the correct number of digits or significant figures:

(a) 523×5000 (c) 27.0×0.01

(b) $5/0.123$ (d) $345.13/100$

1.23 Round the following numbers to 3SFs: (a) 12849m (b) 5111s (c) 2.4566×10^{-3}Kg (d) 0.051376cm (e) 573456mm (f) 0.0293845μm

1.24 Round the following numbers to 1SFs: (a) 12849m (b) 5111s (c) 2.4566×10^{-3}Kg (d) 0.051376cm (e) 573456mm (f) 0.0293845μm

Prefixes & Conversion Factors

1.25 Fill the gap in the following unit equalities or conversion factors: $1Km =$ ______ m

1.26 Fill the gap in the following unit equalities or conversion factors: $1cm =$ ______ m

1.27 Fill the gap in the following unit equalities or conversion factors: $\dfrac{1nm}{\_\_\_\_\_\_ m}$

1.28 Fill the gap in the following unit equalities or conversion factors: $\dfrac{1fs}{\underline{\qquad} s}$

1.29 Fill the gap in the following unit equalities or conversion factors:

$1Tm = \underline{\qquad}$m

$1dm = \underline{\qquad}$m

$1cg = \underline{\qquad}$g

$1ms = \underline{\qquad}$s

$1qt = \underline{\qquad}$mL

$1L = \underline{\qquad} dm^3$

$1lb = \underline{\qquad}$g

$1in = \underline{\qquad}$cm

1.30 Fill the gap in the following unit equalities or conversion factors:

(a) $\dfrac{1\text{km}}{\underline{\qquad}\text{m}}$

(b) $\dfrac{1\text{cm}}{\underline{\qquad}\text{m}}$

(c) $\dfrac{1\text{ms}}{\underline{\qquad}\text{s}}$

(d) $\dfrac{\underline{\qquad}\text{L}}{10^{-3}\text{L}}$

(e) $\dfrac{\underline{\qquad}\text{L}}{10^{-2}\text{L}}$

(f) $\dfrac{\underline{\qquad}\text{g}}{10^{-1}\text{g}}$

Using Conversion Factors

1.31 Fill the gap in the following unit conversion:

$$70\cancel{cm} \times \frac{\underline{\qquad} m}{1\cancel{cm}} = 0.7m$$

1.32 The following conversion factor is used to convert $100\mu m$ into m. Fill in the gaps:

$$100\cancel{\mu m} \times \frac{\underline{\qquad} m}{1\cancel{\mu m}} = 1 \times 10^{-4}m$$

1.33 The following conversion factor is used to convert $40m$ into nm. Fill in the gaps:

$$40\cancel{m} \times \frac{1nm}{\underline{\qquad}\cancel{m}} = 4 \times 10^{10}nm$$

1.34 The following conversion factor is used to convert $500m$ into μm. Fill in the gaps:

$$500\cancel{m} \times \frac{1\mu m}{\underline{\qquad}\cancel{m}} = 5 \times 10^{8}\mu m$$

1.35 Complete the following unit conversion:

$$100\cancel{Gm} \times \frac{\underline{\qquad} m}{\underline{\qquad}\cancel{Gm}} = \underline{\qquad} m$$

1.36 Complete the following unit conversion:

$$50\cancel{km} \times \frac{\underline{\qquad} m}{\underline{\qquad}\cancel{km}} = \underline{\qquad} m$$

1.37 Complete the following unit conversion:

$$2\cancel{m} \times \frac{\underline{\qquad} cm}{\underline{\qquad}\cancel{m}} = \underline{\qquad} cm$$

1.38 Complete the following unit conversion:

$$0.3\cancel{m} \times \frac{\underline{\qquad} mm}{\underline{\qquad}\cancel{m}} = \underline{\qquad} mm$$

1.39 Complete the following unit conversion:

(a) $0.5\cancel{\mu g} \times \dfrac{\underline{\qquad}\text{g}}{\underline{\qquad}\cancel{\mu g}} = \underline{\qquad}\text{g}$

(b) $125\cancel{\text{L}} \times \dfrac{\underline{\qquad}\text{mL}}{\underline{\qquad}\cancel{\text{L}}} = \underline{\qquad}\text{mL}$

1.40 Complete the following unit conversion:

(a) $100\cancel{\text{nm}} \times \dfrac{\underline{\qquad}\text{m}}{\underline{\qquad}\cancel{\text{nm}}} = \underline{\qquad}\text{m}$

(b) $10\cancel{\text{dm}} \times \dfrac{\underline{\qquad}\text{m}}{\underline{\qquad}\cancel{\text{dm}}} = \underline{\qquad}\text{m}$

1.41 The following conversion factor is used to convert $30cm$ into km. Fill in the gaps:

$$30\cancel{cm} \times \frac{\underline{\qquad}\cancel{m}}{1\cancel{cm}} \times \frac{1km}{\underline{\qquad}\cancel{m}} = 3 \times 10^{-4}km$$

1.42 The following conversion factor is used to convert $50dm$ into cm. Fill in the gaps:

$$50\cancel{dm} \times \frac{\underline{\qquad}\cancel{m}}{1\cancel{dm}} \times \frac{1cm}{\underline{\qquad}\cancel{m}} = 500cm$$

1.43 Fill the gap in the following conversion factors:

$$20\cancel{cm} \times \frac{1 \times 10^{-2}m}{1\cancel{cm}} \times \frac{1mm}{1 \times 10^{-3}\cancel{m}} = \underline{\qquad} mm$$

1.44 Fill the gap in the following conversion factors:

$$5\cancel{mm} \times \frac{1 \times 10^{-3}m}{1\cancel{mm}} \times \frac{1nm}{1 \times 10^{-9}\cancel{m}} = \underline{\qquad} nm$$

1.45 Set up the conversion factor to convert 50cm into inches:

$$50\cancel{cm} \times \frac{\rule{2cm}{0.3cm}}{\rule{2cm}{0.3cm}} = \rule{2cm}{0.3cm}\,in$$

1.46 Fill the gap in the following conversion factors:

$$20\cancel{cm} \times \frac{\rule{2cm}{0.3cm}}{\rule{2cm}{0.3cm}} = 7.87in$$

1.47 Compute the following power of ten calculations:

(a) $(10^2)^2$ (c) $(10^{-6})^2$

(b) $(10^2)^3$ (d) $(10^{-2})^2$

1.48 Compute the following power of ten calculations:

(a) $(10^1)^3$ (c) $(10^{-5})^2$

(b) $(10^3)^4$ (d) $(10^{-4})^5$

1.49 Set up the following conversion factor:

$$400\cancel{cm^2} \times \frac{\rule{2cm}{0.3cm}}{\rule{2cm}{0.3cm}} = \rule{2cm}{0.3cm}\,m^2$$

1.50 Set up the following conversion factor:

$$0.4\cancel{mm^3} \times \frac{\rule{2cm}{0.3cm}}{\rule{2cm}{0.3cm}} = \rule{2cm}{0.3cm}\,m^3$$

1.51 Fill the gap in the following unit equalities or conversion factors:

$1\text{cm}^2 = \rule{2cm}{0.3cm}\,\text{m}^2$ $1\text{cm}^3 = \rule{2cm}{0.3cm}\,\text{m}^3$

$1\text{dm}^3 = \rule{2cm}{0.3cm}\,\text{m}^3$ $1\text{dm}^2 = \rule{2cm}{0.3cm}\,\text{m}^2$

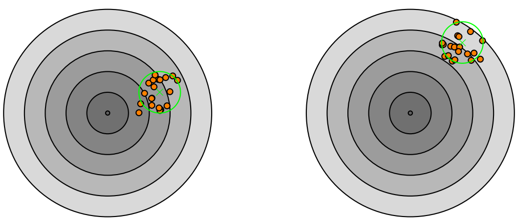
1.52 Fill the gap in the following unit equalities or conversion factors:

$1\text{mm}^2 = \rule{2cm}{0.3cm}\,\text{m}^2$ $1\mu\text{m}^2 = \rule{2cm}{0.3cm}\,\text{m}^2$

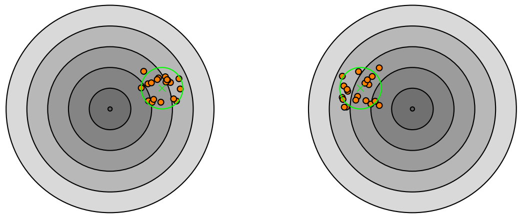
$1\text{nm}^2 = \rule{2cm}{0.3cm}\,\text{m}^2$ $1k\text{m}^2 = \rule{2cm}{0.3cm}\,\text{m}^2$

Measurements and Uncertainty

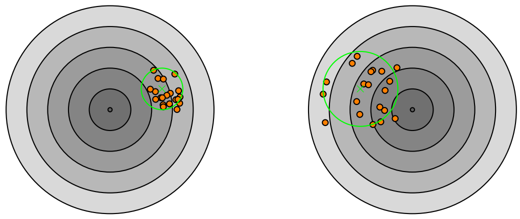
1.53 Compare the precision and accuracy of the following measurements:



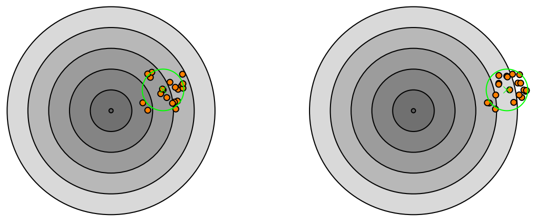
1.54 Compare the precision and accuracy of the following measurements:



1.55 Compare the precision and accuracy of the following measurements:



1.56 Compare the precision and accuracy of the following measurements:



Density

1.57 Determine the density (g/mL) of a 0.01 L sample of a salt solution that has a mass of 50 g.

1.58 Determine the density (g/mL) of a 0.05 L sample of a salt solution that has a mass of 10 g.

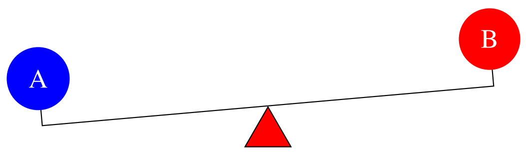
1.59 Which one of the following substances will float in gasoline, which has a density of 0.66 g/mL? Assume no mixing: (a) table salt (2.16 g/mL) (b) balsa wood (0.16 g/mL) (c) sugar (1.59 g/mL) (d) aluminum (2.70 g/mL) (e) mercury (13.6 g/mL)

1.60 You have a large water tank used as a cooler in a party and you have a bunch of cans: a coke can, a diet coke can, a water can and a schweppes can. You add all unopened cans on the tank. Describe the final vertical distribution of cans in the tank. Which can will stay on top and which will sink in more?

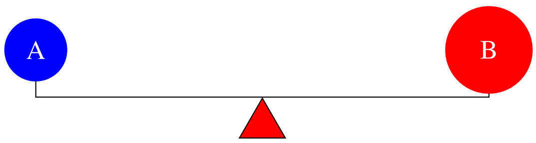
1.61 A nugget of gold with a mass of 521 g is added to 50.0 mL of water. The water level rises to a volume of 77.0 mL. What is the measured density of the gold?

1.62 A graduated cylinder contains 28.0 mL of water. What is the new water level after 35.6 g of silver metal is submerged in the water if the density of silver is 10g/mL?

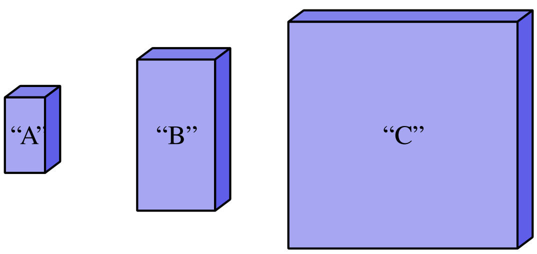
1.63 Which of the circles is more dense: A or B.



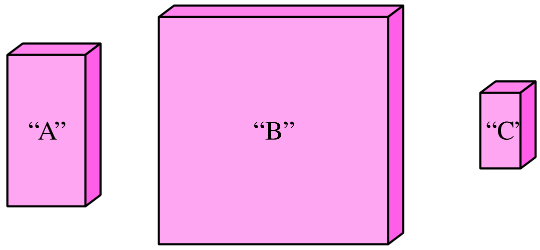
1.64 Which of the circles is more dense: A or B.



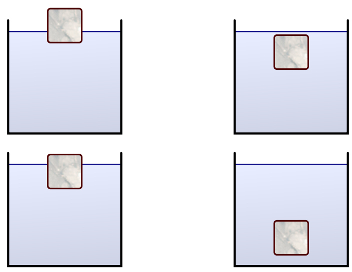
1.65 The objects below have all the same mass but represent different metals cobalt, gold, and palladium. Given that the density of cobalt is 8g/mL, whereas the density of gold is 19g/mL and the density of palladium is 12g/mL, identify each object as cobalt, gold, and palladium.



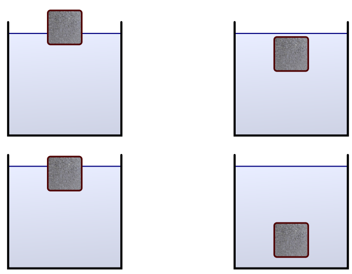
1.66 The objects below have all the same mass but represent different metals iron, nickel, and tungsten. Given that the density of iron is 7g/mL, whereas the density of nickel is 9g/mL and the density of tungsten is 20g/mL, identify each object as iron, nickel, or tungsten.



1.67 The diagrams below represent a solid immersed in a liquid, with some of the solids floating and others sinking. Match the diagrams with the description: (a) The density of the solid is 1.1 times the density of the liquid (b) The density of the solid is the same the density of the liquid (c) The density of the solid is 0.8 times the density of the liquid (d) The density of the solid is 0.5 times the density of the liquid



1.68 The diagrams below represent a solid immersed in a liquid of density 0.9g/mL, with some of the solids floating and others sinking. Match the diagrams with the description: (a) The density of the solid is 1.2g/mL (b) The density of the solid is 0.9g/mL (c) The density of the solid is 0.7g/mL (d) The density of the solid is 0.4g/mL



1.69 Answer the following questions involving specific gravity: (a) A sample has a density of 0.9g/mL. How much is its specific gravity? (b) A sample has a specific gravity of 1.20. How much is its density? (c) True or false: an orange juice sample should have a specific gravity larger than one.

1.70 Answer the following questions involving specific gravity: (a) True or false: an olive oil sample should have a specific gravity larger than one. (b) A sample has a mass of 5mg and a specific gravity of 0.87. Calculate its volume in mL.

1.71 Answer the following questions involving dosages: (a) An IV solution needs to be delivered at a rate of 120mL/h. How long it takes to deliver 200mL? (b) Infant Tylenol is given from a 160mg/5mL suspension based on the infant's weight. A 12lb infant requires 2.5mL. How many milligrams of Tylenol are contained in the dosage given?

1.72 Answer the following questions involving dosages: (a) A patient requires 1g of medication given every three hours. The medication in stock was found in tablets of 200mg. How many tablets do you need in three days? (b) A medication needs to be given based on the patient's body weight as 2mg/Kg. If a patient weighs 70Kg and the mediation stock is 100mg/mL, how many mL are needed?

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Ch. 2. Energy and Matter

ENERGY involves many aspects of our everyday life. Chemical reactions happening in our bodies consume or release energy as we walk, study, and even breathe. We also use energy at home to warm food or turn on the lights, drive our cars, and go to work. The energy needed for our body to function comes from food. If we do not eat for a while, we run low on energy. Similarly, the burning of fossil fuels such as oil, propane, or gasoline provides enough energy to maintain our homes. On the other hand, how do we measure the energy released or consumed in a chemical process? This chapter will answer this and other questions as it covers different aspects of the interplay between chemistry and energy. You will learn about temperature, heat, and how an energy exchange implies changes in temperature.

2.1 Matter

Matter–the material of the universe–represents anything with mass that occupies space. It has different levels of organization and complexity. We can classify matter in terms of composition. Some substances are made of a single component whereas others contain multiple components. At the same time, some substances are made of many components while they appear to be made of a single component. Figure 2.1 displays the classification of matter in terms of mixture nature and composition.

`Solids, liquids & gases` *Solid* has a well-defined shape and volume. Think about an ice cube, for example, that is made of water in the form of ice. In an ice cube, attractive forces keep the shape of the cube constant. The water molecules in the ice are arranged in rigid patterns and they can only vibrate in the solid. A *liquid*, on the other hand, has a well-defined volume. However, liquids do not have a constant shape, as their shape will depend on the shape of the container. Think about water which is a liquid. You can find many different bottle shapes. In all of them, the molecules of water will arrange to occupy the shape of the container. The volume of the liquid–the amount of space the liquid occupies–will be constant but not the shape. In a liquid, the particles move randomly but are still attached. A *gas* does not have a well-defined shape or volume. In a gas, the particles are randomly distributed and barely interact with each other, as they move at high speeds, taking the shape and volume of their container. Figure 2.2 displays the microscopic structure of the three different states of matter of water. The process of boiling and freezing represent *physical changes* and during physical changes, these substances involved do not change their composition. When water boils, both steam and water are made of the same component, H_2O. When a substance undergoes a *chemical process*, it will change its composition. For example, when burning a piece of paper, paper made of carbon, hydrogen, and hydrogen becomes ash made mostly of carbon. As such, during a chemical process, a substance (e.g. paper) becomes g. carbon).

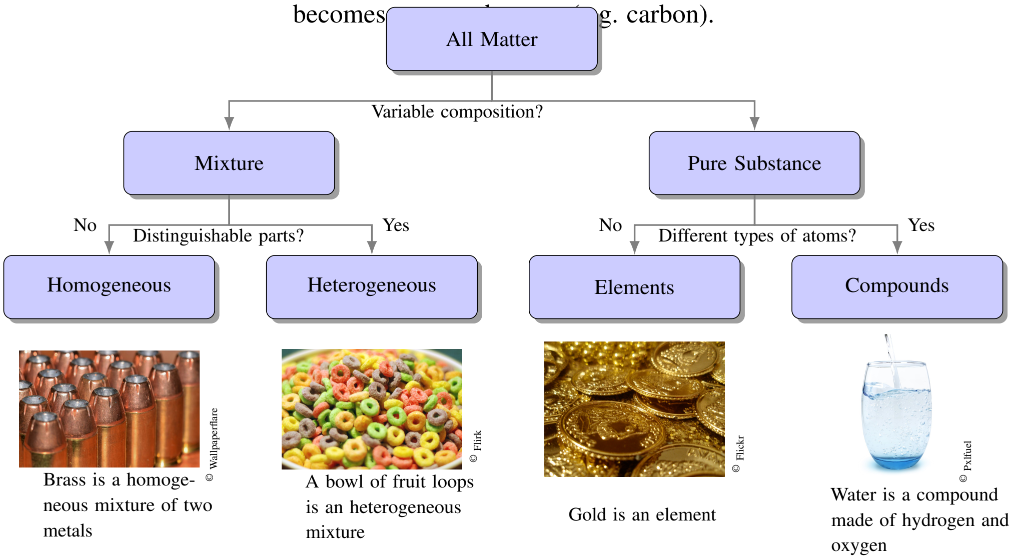


Figure 2.1 Classification of the matter

`Pure Substances and mixtures` On one hand, *pure substances* have definite composition, being only made of a single component. For example, water and gold are pure substances. However, there are two different types of pure substances: elements and compounds. *Elements* is composed of only one type of atom. Examples are silver, iron, and aluminum. They all contain one type of substance, and for example, iron is only made of iron atoms. *Compounds* are combinations of different elements. For example, water, H_2O is a compound made of a combination of hydrogen and oxygen atoms. On the other hand, *mixtures* are physical combinations of different pure substances. Mixtures have variable compositions. For example, the air is a mixture of oxygen and nitrogen. Wood, soda pop, or soil are all mixtures. Mixtures can be homogeneous or heterogeneous. In a *homogeneous mixture*–also known as solutions–the composition is uniform throughout the sample. An example of a homogeneous mixture is salty water, a solution of salt and water. *Heterogeneous mixtures* are mixtures in which the components are not uniformly distributed throughout the sample. An example would be a chocolate chip cookie in which you can differentiate the dough and the chocolate.

Sample Problem 13

Classify as element, compound, homogeneous mixture, heterogeneous mixture:
(a) An iron nail (b) Milk (c) Sugar (d) miso soup

SOLUTION

(a) An iron nail is an element as it is only made of iron, a single material; (b) Milk is a homogeneous mixture as it is made of water, fat, protein even though you only see a single substance; (c) Sugar is a compound made of carbon and other constituents; (d) miso soup is a mixture of water, fat and other chemicals and therefore is a mixture. As you can differentiate its constituents we call this heterogeneous mixture.

STUDY CHECK

Classify as element, compound, homogeneous mixture, heterogeneous mixture:

(a) muscle milk (b) water (c) a gold ring (d) rice & beans

▼The molecular hexagonal structure of ice, solid water.

▼The molecular structure of liquid water.

▼The molecular structure of steam, gaseous water.

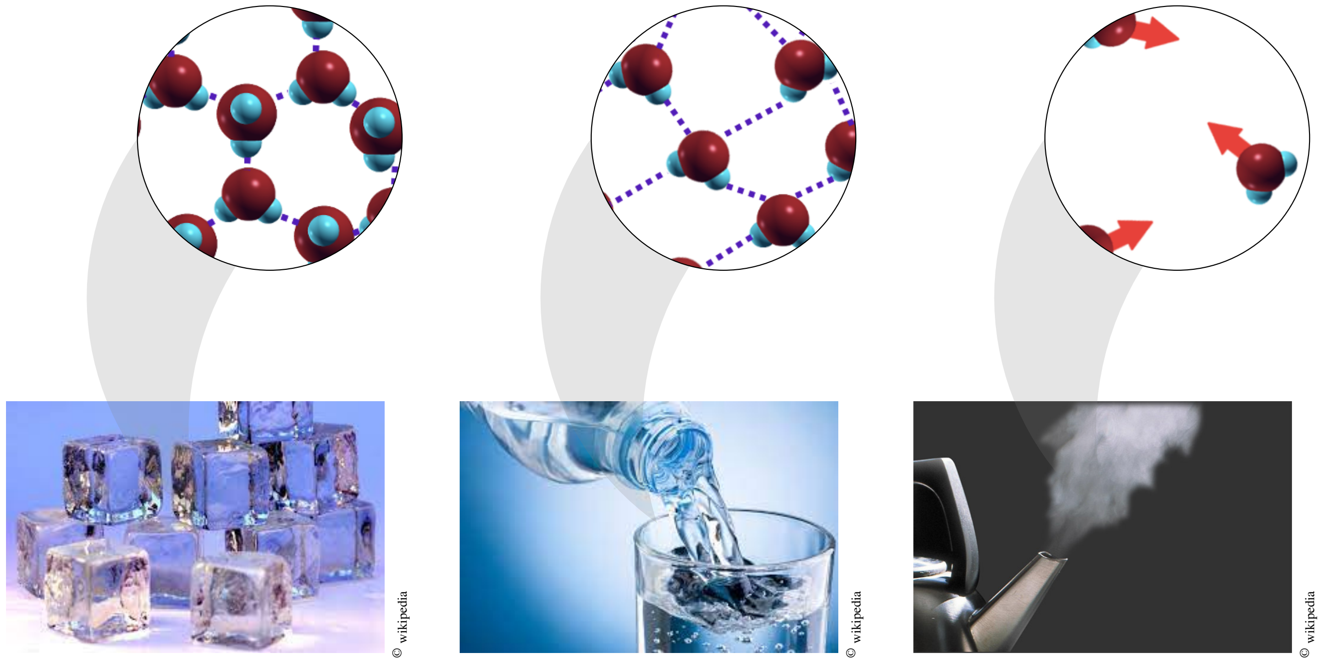


Figure 2.2 Three different states of matter, solid, liquid, and gas. Solid molecules are locked into fixed positions, while liquid molecules are together but still can move. Gas molecules are apart and move freely.

2.2 Energy

When you are running, walking, dancing, or thinking, you are using energy to do work. Energy is defined as the ability to do work. Suppose you are climbing a steep hill. Perhaps you become too tired to go on; you do not have sufficient energy to do any more work. Now suppose you sit down and have lunch. In a while, you will have obtained energy from the food, and you will be able to do more work and complete the climb. Similarly, chemical energy is the energy stored in the structure of chemicals and it depends on the arrangement of molecules and the nature of these molecules.

`Potential & Kinetic Energy:` `heat` Energy can be classified as potential energy or kinetic energy. Kinetic energy is the energy of motion and any fast-moving object has kinetic energy. Think about a bullet coming out of a gun; as the bullet moves very fast it contains kinetic energy that can be released when it collides with a target. Potential energy is energy stored in objects located at a certain height. A boulder resting on top of a mountain has potential energy because of its location. If the boulder rolls down the mountain, the potential energy becomes kinetic energy. Water stored in a

▼Flowers convert sunlight into chemical energy

© www.wallpaperflare.com

▼a bullet has kinetic energy

© www.wallpaperflare.com

▼water on a dam has potential energy

© wikipedia

▼thermal energy refers to heat

© wikipedia

reservoir has potential energy. When the water goes over the dam, the potential energy is converted to kinetic energy. The potential energy resulting from the interaction of charged particles is called electrostatic energy. Heat refers to thermal energy, which is associated with the random motion of particles in a substance and therefore is related to kinetic energy. A frozen pizza feels cold because the particles in the pizza are moving slowly. As the pizza receives heat, the motion of the particles increase. Eventually, the particles have enough energy to make the pizza hot and ready to be eaten. When a substance receives heat it gets warmer and it raises its temperature, whereas if it loses heat it gets cooler and its temperature decreases.

`Mechanical energy: work` The sum of potential and kinetic energy is called mechanical energy. Mechanical energy refers to the ability to do work. Examples of work are a car engine moving or a balloon expanding its volume.

`The law of conservation energy` In this chapter, we will analyze energy changes associated with chemical reactions. To do this, we need to define the system and its surroundings. The system will be the chemical reaction often happening in a beaker, whereas the surroundings would be the area surrounding the beaker. The system plus its surroundings are called the universe. The beaker may lose energy, and in that case, energy will flow from the system to the surroundings. Similarly, the system may gain energy, flowing from the surroundings to the beaker. In a closed system, the energy is conserved and when one type of energy disappears, a different type of energy will appear. As an example, if you drop an object from the top of a building, originally the object had potential energy that converts into kinetic energy as the object gains speed. This is called the law of conservation of energy.

`Energy units` Two different units of energy are often employed: calories (cal) and joules (J). Joule is the SI unit of energy equal to $\mathrm{kg \cdot m^2/s^2}$. One can transform calories into joules and joules into calories using the following conversion factor:

$$\boxed{1\mathrm{cal} = 4.184\mathrm{J}} \quad \text{or} \quad \boxed{\frac{1\mathrm{cal}}{4.184\mathrm{J}}} \tag{2.1}$$

As a note, often you will read on food labels the caloric content of certain foods. In these labels, they use the unit Calorie, with capitalized C, which is not the same as a calorie. One Calorie represents a kilocalorie and contains 1000 calories.

Sample Problem 14

Convert the following energy values:

(a) 50000 cal to Kcal (b) 48001 J to cal

SOLUTION

We will use the conversion factor for kilo and the relationship between calorie and joule:

(a) $50000\mathrm{cal} \times \frac{1\mathrm{kcal}}{1000\mathrm{cal}} = 50\mathrm{kcal}$; (b) $48000\mathrm{J} \times \frac{1\mathrm{cal}}{4.184\mathrm{J}} = 11472\mathrm{cal}$.

STUDY CHECK

Convert the following energy units: (a) 200 cal to Kcal (b) 7000 J to cal

`Thermodynamics` Thermochemistry is a subject of a broader field called thermodynamics, which studies the interconversion of energy (heat and other types) and mass. Thermodynamics study systems, like chemical reactions. The term system, refers to the part of the universe being study. Systems can be classified as: open, closed and isolated systems. An open system can exchange mass and energy with its surroundings, whereas a closed system can only exchange mass and not energy. Isolated systems cannot exchange neither mass nor energy with its surroundings. The state of a system is characterized by the values of its volume, pressure, temperature, energy and composition, so that if a system receives heat it will change its state. Energy, volume, pressure and temperature are called state functions–or state properties–, as these properties are determined by the state of the system, independently of the path used to reach that state. In another words, these properties are path-independent. For example, in a building, the floor location of a person would be a state function, as it would not matter the path the person took to reach that state. In contrast the amount of effort to make it to a specific floor will not be a state function, as it changes depending on the path used.

2.3 Temperature

Temperature indicates how hot or cold a substance is compared to another substance. Heat always flows from a substance with a higher temperature to a substance with a lower temperature until the temperatures of both are the same. When you drink hot coffee or touch a hot pan, heat flows to your mouth or hand, which is at a lower temperature. When you touch an ice cube, it feels cold because heat flows from your hand to the colder ice cube. Three units of temperature often employed are celsius (°C, T_C), Fahrenheit (°F, T_F) or Kelvin (K, T_K). If you need to convert temperature units from Fahrenheit to celsius or from celsius to Fahrenheit you need to use the formulas below:

$$\boxed{T_F = 1.8T_C + 32} \qquad \boxed{T_F = 1.8T_K - 459.4} \qquad \boxed{T_K = T_C + 273} \tag{2.2}$$

Sample Problem 15

Convert 25 °C to °F.

SOLUTION

1. **Step one:** list of the given variables.

Analyze the Problem	Given	Asking
	$T_c = 25°C$	T_F

2. **Step two:** use the formula $T_F = 1.8T_C + 32$ to convert from °C to °F.
3. **Step three:** solve for $T_F = 1.8 \times 25 + 32 = 77°F$.

STUDY CHECK

Convert 200°C to K.

2.4 From energy to temperature

Heat transforms in a temperature change. Some substances like metals can increase their temperature very quickly with a small amount of heat received, whereas others need a larger amount of heat to rise their temperature. Think about why you use oil to deep fry food. Why not use water? First of all, oil can raise its temperature very quickly and on top of that it does not boil easily.

`Heat capacity` The heat capacity c of a material is defined as:

$$c = \frac{\text{heat adsorbed}}{\text{temperature increase}} \tag{2.3}$$

This is a characteristic property of each material that indicates the energy required to rise its temperature and can be expressed in cal/°C or J/°C units. As this property depends on the amount of matter, oftentimes the heat capacity is expressed per mass as the specific heat capacity also known as *specific heat* (c_e) or mole unit as the *molar heat capacity* c_m. For example, the specific heat of water is 1cal/g°C that is the same as 4.184J/g°C. That means that we need to give 1 calorie to warm up one gram of water $1°C$. Similarly, the specific heat of aluminum, a metal, is 0.2cal/g°C or 0.89J/g°C; that means the energy needed to raise the temperature of an aluminum gram is 0.2 calories of 0.89 J. Mind the difference between these two values: we need to give 1 cal to increase the temperature of a gram of water in $1°C$, whereas we need to give 0.2 cal to increase the temperature of a gram of aluminum in $1°C$. Why are these two numbers so different? The answer is that water and aluminum are different materials. Normally metals warp up very easily, that is, they need less heat to increase their temperature, whereas liquids need more heat to increase their temperature. That is why pans and cooking pots tend to be metallic. Table 2.1 lists specific heats of common substances. Mind the specific heat if water is a well know value that you need to be familiar with:

$$\boxed{c_e^{\mathrm{H_2O}} = 4.184\mathrm{J/g°C} \quad or \quad c_e^{\mathrm{H_2O}} = 1\mathrm{cal/g°C}} \tag{2.4}$$

Table 2.1 Values of specific heat for different materials

Material	Specific heat (J/g°C)	Material	Specific heat (J/g°C)
$H_2O_{(l)}$	4.184	$Fe_{(s)}$	0.444
ethyl alcohol$_{(l)}$	2.460	$Au_{(s)}$	0.129
vegetable oil$_{(l)}$	1.790	$Cu_{(s)}$	0.385
$NH_{3(l)}$	4.700	$H_2O_{(s)}$	2.010
Dry Air$_{(g)}$	1.0035	$CO_{2(g)}$	0.839

`Heat` When a material receives heat, that heat normally becomes temperature as the temperature of the material increases. For example, if you warm milk in a microwave, the milk's temperature increases from room temperature (25°C) to a higher temperature. How to estimate the temperature increase given the heat received? Or how to estimate the heat needed to increase the temperature of an object? We can use the following formula:

$$\boxed{Q = m \cdot c_e \cdot (T_f - T_i)} \tag{2.5}$$

where:

Q is the amount of heat received, either in cal or J.

m is the mass of material in grams

c_e is the specific heat of the material (in cal/g°C or J/g°C)

$T_f - T_i = \Delta T$, is the temperature change from the initial to the final temperature

A system can receive or give away heat and this is indicated by the sign of Q. The sign convention for heat is:

$Q > 0$ the system receives heat *and* $Q < 0$ the system gives away heat

Sample Problem 16

How many calories are absorbed by a 45.2g piece of aluminum ($c_e = 0.214\frac{cal}{g°C}$) if its temperature rises from 25°C to 50°C.

SOLUTION

1 **Step one:** list of the given variables.

Analyze the Problem	Given	Asking
	$c_e = 0.214\frac{cal}{g°C}$ $m = 45.2g$ $T_{initial} = 25°C$ $T_{final} = 50°C$	Q

2 **Step two:** use the formula $Q = m \cdot c_e \cdot (T_{final} - T_{initial})$ to transform the temperature increase into heat absorbed. Mind this formula depends on the mass involved and the specific heat of the material, in this case, aluminum.

3 **Step three:** solve $Q = 45.2 \cdot 0.214 \cdot (50 - 25) = 241.82cal$.

STUDY CHECK

How many calories are absorbed by 100g of Gold ($c_e = 0.0308\frac{cal}{g°C}$) if its temperature rises from 25°C to 100°C.

In the previous example you needed to convert temperature into heat. In the next example, the heat is given and you need to calculate the final temperature of an object after it receives a certain amount of heat.

Sample Problem 17

A 50g piece of aluminum ($c_e = 0.214\frac{cal}{g \cdot °C}$) initially at 25°C absorbs 100cal. Calculate the final temperature of the aluminum piece.

SOLUTION

1 **Step one:** list of the given variables.

Analyze the Problem	Given	Asking
	$c_e = 0.214\frac{cal}{g \cdot °C}$ $m = 50g$ $T_{initial} = 25°C$ $Q = 100cal$	T_{final}

2 **Step two:** use the formula $Q = m \cdot c_e \cdot (T_{final} - T_{initial})$ that converts temperature increase to heat absorbed.

3 **Step three:** solve $100 = 50 \cdot 0.214 \cdot (T_{final} - 25)$ for T_{final}:

$$\begin{aligned}
100 &= 50 \cdot 0.214 \cdot (T_{final} - 25) && \text{divide by 50 in both sides}\\
\frac{100}{50} &= 0.214 \cdot (T_{final} - 25) && \text{divide by 0.214 in both sides}\\
\frac{100}{50 \cdot 0.214} &= (T_{final} - 25) && \\
9.34 &= (T_{final} - 25) && \text{add 25 in both sides}\\
9.34 + 25 &= T_{final} && \\
34.34 &= T_{final} &&
\end{aligned}$$

The final temperature of the aluminum piece is $34.34°C$.

STUDY CHECK

A 200g piece of iron ($c_e = 0.1 \frac{cal}{g \cdot °C}$) initially at 15°C absorbs 1000cal. Calculate the final temperature of the metal piece in °C.

2.5 Energy value of food

Table 2.2 Energy value of food

Food Type	Energy value ($\frac{kcal}{g}$)
Carbohydrates	4
Fat	9
Protein	4

`Calories in food` How much food do you eat? How many calories do you ingest a day? When you are watching your food intake, the Calories you are counting are kilocalories (1000cal, Kcal, or Cal). In the field of nutrition, it is common to use the Calorie, Cal (with an uppercase C) to indicate 1000 cal or 1 kcal.

$$\boxed{1Cal = 1000cal} \quad \text{or} \quad \boxed{\frac{1Cal}{1000cal}} \quad \text{or} \quad \boxed{\frac{1000cal}{1Cal}} \tag{2.6}$$

`Energy values` Do you ever eat pasta? Think about how does your body feel after you eat pasta? Normally, whenever you eat pasta in a few hours you need to eat again more food. Differently, whenever you eat meat, that is enough to keep you going for a longer time. Similarly, eating a salad for lunch brings you less energy than a pizza slice. This is because each type of food–each ingredient–contains different energy. We refer to this as the energy value of food ϵ. Table 2.2 lists energy values for common ingredients. To compute the energy (E) provided by a certain mass of food (m) we need to multiply the mass times the energy value (ϵ):

$$E = m \cdot \epsilon \tag{2.7}$$

For example, the energy value of fat ϵ_{fat} is $9\frac{kcal}{g}$, which means that if you eat three grams of fat that will bring you a given amount of energy E_{fat}:

$$E_{fat} = 3\cancel{g} \times 9\frac{kcal}{\cancel{g}} = 18kcal$$

Normally, when you eat a food plate, you ingest energy from the different types of ingredients of that plate: fat, carbs, or protein.

Sample Problem 18

A Big Mac from McDonalds contains 28g of fat ($9kcal/g$), 46g of carbs ($4kcal/g$) and 25 g of protein ($4kcal/g$), where the caloric values are indicated in parenthesis. What is the total energy content of a Big Mac?

SOLUTION

1 **Step one:** list of the given information and the unknown.

Analyze the Problem	Given	Asking
	$m_{fat} = 28g$ $\epsilon_{fat} = 9\frac{kcal}{g}$ $m_{carb} = 46g$ $\epsilon_{carb} = 4\frac{kcal}{g}$ $m_{prot} = 25g$ $\epsilon_{prot} = 4\frac{kcal}{g}$	Energy Content

2 **Step two:** use the formula $E_{fat} = m_{fat} \cdot \epsilon_{fat}$ to calculate the energy coming from fat. And do the same for carbs and protein.

3 **Step three:** Compute the energy coming from each ingredient and add all the values:

$$E_{fat} = m_{fat} \cdot \epsilon_{fat} = 28\cancel{g} \times 9\frac{kcal}{\cancel{g}} = 252kcal \qquad \text{energy from fat}$$

$$E_{carb} = m_{carb} \cdot \epsilon_{carb} = 46\cancel{g} \times 4\frac{kcal}{\cancel{g}} = 184kcal \qquad \text{energy from carbs}$$

$$E_{prot} = m_{prot} \cdot \epsilon_{prot} = 25\cancel{g} \times 4\frac{kcal}{\cancel{g}} = 100kcal \qquad \text{energy from protein}$$

The total energy content of a Big Mac is: $E_{fat} + E_{carb} + E_{prot} = 532kcal$

STUDY CHECK

A pepperoni pizza slice contains 11g of fat ($9kcal/g$), 36g of carbs ($4kcal/g$) and 14 g of protein ($4kcal/g$), where the caloric values are indicated in parenthesis. What is the total energy content of a pizza slice?

CHAPTER 2

Matter

2.1 Classify the following objects as an element, compound, and homogeneous mixture, a heterogeneous mixture or none of the others: (a) an energy drink (b) helium (a gas) (c) sulfur (a solid) (d) milk (e) milkshake (f) gelato (g) air (h) granite (i) uranium (a solid)

2.2 Classify the following objects as an element, compound, and homogeneous mixture, a heterogeneous mixture or none of the others: (a) a copper wire (b) a chocolate cookie (c) a chocolate-chip cookie (d) vinegar (e) ice (f) baking soda (g) aluminum foil (h) vitamin C

2.3 Which of the following is a property of a solid?(a) It takes the shape of the container. (b) It fills the volume of the container. (c) The particles move at a rapid rate. (d) The interactions between its particles are very weak. (e) The particles have fixed positions and are very close together.

2.4 Which of the following is a property of a liquid? (a) It takes the shape of the container. (b) it has no volume. (c) The particles move at a rapid rate. (d) The interactions between its particles are unesxistent. (e) The particles have fixed positions and are very close together.

2.5 Which of the following is a property of a gas? (a) It has no shape. (b) It fills the volume of the container. (c) The particles move slowly. (d) The interactions between its particles are strong. (e) The particles have fixed positions and are very close together.

2.6 Which of the following is a property of a gas and a liquid? (a) It flows. (b) It takes the shape of the container. (c) It has no shape. (d) It fills the volume of the container. (e) The particles move slowly. (f) The interactions between its particles are strong. (g) The particles move at a rapid rate. (h) The interactions between its particles are unesxistent. (i) The particles have fixed positions and are very close together.

Energy and temperature

2.7 Answer the following questions: (a) What is the name of the energy associated with the motion of particles in a substance? (b) What is the name of the energy stored in heigh?

2.8 Discuss the changes in potential and kinetic energy in the following scenarios: (a) When water falls down a waterfall (b) When a person throughs away an object high up in the air

2.9 Indicate whether the following statement describes kinetic or potential energy: (a) Water on top of a waterfall (b) Hitting a wall with a hammer

2.10 Indicate whether the following statement describes kinetic or potential energy: (a) a car moving in the road (b) Water on the bottom of a waterfall

2.11 Carry the following conversions: (a) 100°C to K (b) 200°F to K (c) 500K to °F

2.12 Carry the following conversions: (a) 20°C to °F (b) 300K to °C (c) 41°F to °C

2.13 Carry the following conversions: (a) 100 Cal into kcal (b) 10000 cal into J (c) 4565J into Cal

2.14 Carry the following conversions: (a) 650J into calories (b) 50 kcal into Cal (c) 3.25 kcal into joules

From energy to temperature

2.15 The same amount of heat is provided to a sample of two different metals, metal A ($c_e(A) = 0.3\frac{cal}{g°C}$) and metal B ($c_e(B) = 0.4\frac{cal}{g°C}$). Both samples have the same mass and are at the same temperature. Which metal A or B would reach a higher temperature?

2.16 The same amount of heat is provided to a sample of two different metals, metal A ($c_e(A) = 0.5\frac{cal}{g°C}$) and metal B ($c_e(B) = 0.4\frac{cal}{g°C}$). Both samples have the same mass and are at the same temperature. Which metal A or B would reach a higher temperature?

2.17 Two samples, A and B, have the same mass and are at the same temperature. If they are equally heated the final temperature of A is three times the one for B.

Compare the specific heats of the samples.

2.18 Two samples, A and B, have the same mass and are at the same temperature. If they are equally heated the final temperature of B is two times the one for A. Compare the specific heats of the samples.

2.19 A 50g piece of aluminum ($c_e = 0.214\frac{cal}{g°C}$) initially at 25°C absorbs 100cal. Calculate the final temperature of the aluminum piece.

2.20 A 200g piece of iron ($c_e = 0.1\frac{cal}{g°C}$) initially at 15°C absorbs 1000cal. Calculate the final temperature of the metal piece.

2.21 How many calories are required to raise the temperature of a 35 g sample of iron from 25°C to 35°C? Iron has a specific heat of $0.108\frac{cal}{g°C}$.

2.22 What is the final temperature of a 35 g sample of iron initially at 25°C after receiving 50cal? Iron has a specific heat of $0.108\frac{cal}{g°C}$.

2.23 What is the initial temperature of a 50 g sample of aluminum that after receiving 50cal reaches a temperature of 50°C? Al has a specific heat of $0.2\frac{cal}{g°C}$.

2.24 What is the specific heat of a metal if a 100 g sample at 25°C warms up until 50°C after receiving 100cal?

Calorimetry

2.25 A potato contains 20 g of carbohydrate. If carbohydrate has a caloric value of 4 $kcal/g$, how many kcal are obtained from the carbohydrate in the potato?

2.26 A diet has a total caloric intake of 1400 kcal. The diet consists of 50.% carbohydrate, 35% protein, and 15% fat. The number of kcal of protein in the diet is

2.27 A serving of fish contains 50 g of protein and 4 g of fat. If protein has a caloric value of 4.0 $kcal/g$ and fat has 9 $kcal/g$, how many kcal are in the serving?

2.28 One large egg contains 6 g of protein and 6 g of fat. If protein has a caloric value of 4.0 $kcal/g$ and fat has 9 $kcal/g$, how many kcal are in the egg?

Ch. 3. Atoms

MATTER is everywhere around you, from the water you drink to the air you inhale. The matter is made of elements and elements are made of atoms. Even the atoms of an element can be different, having a distinct number of protons and neutrons. This chapter covers the principles of atomic structure. You will learn what makes an atom and will be able to quantify the particles that make atoms. Perhaps more importantly, you will also learn about the periodic table of elements and the different types of chemical formulas.

3.1 The periodic table

The periodic table (see Figure 3.1) is a chart containing all known elements arranged in increasing number of electrons per atom in a way that elements with similar chemical and physical properties are located together. The periodic table contains all existing elements–some of them are synthetic others are natural–that form the matter arranged in columns and rows. Every element has a different name accompanied by a symbol that represents its name. The tabular arrangement of elements in the form of rows and columns allows further classification of the elements according to their properties. This section will cover the different features of the periodic table.

Elements and Symbols Elements cannot be broken down into simpler substances. For example, aluminum is an element only made of aluminum atoms and if you analyze the composition of a piece of this metal you would only find aluminum atoms. Chemical symbols are one- or two-letter abbreviations that represent the names of the elements. Only the first letter is capitalized and if a second letter exists in the element's name, the second letter should be lowercase. For example, the chemical symbol for aluminum is Al, written as capital A and lowercase l.

Sample Problem 19

Give the symbol or name the following elements: C, Oxygen, N, Phosphorus, Au, Iron, Na and Iodine.

SOLUTION

The name of the element with symbol C is carbon, whereas the chemical symbol of oxygen is O. Similarly, N stands for nitrogen, whereas the chemical symbol of Phosphorus is P. The chemical symbol of Au is Gold. The chemical symbol of Iron is Fe and the chemical symbol of Iodine is I.

STUDY CHECK

Give the symbol or name the following elements: Ni.

`Periods and groups` The periodic table (see Figure 3.1) contains all elements arranged in rows and columns. The horizontal rows are called *periods* and the vertical columns are called *groups or families*. For example, the first period contains hydrogen (H) and helium (He), whereas the second group contains Beryllium (Be), Magnesium (Mg), Calcium (Ca), Strontium (Sr), Barium (Ba) and Radium (Ra). There are seven periods (periods 1-7) and 18 groups. Some of the groups are labeled with an A (e.g. group 8A) whereas others are labeled with a B (e.g. group 8B). Group numbers can be found written with roman numbers and a letter (A or B) or with a more modern group numbering of 1-18 going across the periodic table. For example, group 2 (Mg-Ra) can also be called IIA, and group 13 (B-Ti) is also known as IIIA.

`Properties in the periodic table` The physical and chemical properties of some elements of the table (see Figure 3.1) are similar, and these similarities led to the organization of the periodic table. Elements in the same group share properties and for example, oxygen and sulfur have similar properties: both are reactive elements. Differently, the properties across periods change going from metals to nonmetals. For example, the properties of Li and Ne are very different, and lithium is a reactive metal whereas neon is a nonreactive gas.

`Metals, Nonmetals, and Metalloids` Overall, the elements of the periodic table (see Figure 3.1) can be classified as metals, nonmetals, and metalloids. Metals are those elements on the left of the table and nonmetals are the elements on the right of the table. The elements between metals and nonmetals are called metalloids and include only B, Si, Ge, As, Sb, Te, Po, and At. Metals are shiny solids and usually melt at higher temperatures. Some examples of metals are Gold (Au) or Iron (Fe). Nonmetals are often poor conductors of heat and electricity with low melting points. They also tend to be matt (non-shinny), malleable, or ductile. Some examples of nonmetals are Carbon (C) or Nitrogen (N). Metalloids are elements that share some properties with metals and others with nonmetals. For example, they are better conductors of heat and electricity than nonmetals, but not as good conductors as metals. Metalloids are semiconductors because they can act as both conductors and insulators under certain conditions. An example of metalloids is Silicon (Si) which should not be confused with silicone, a chemical employed in prosthetics.

Sample Problem 20

Answer the following questions: (a) Give the group and period of the following elements, and give the name: Ca, Ir, and C. (b) Classify as alkali metal, alkali earth metal, transition metal, halogen or noble gas, and give the name: Mg, Li, Co, He, F. (c) Classify as metal, nonmetal or metalloid, and give the name: Ba, N, Si.

SOLUTION

(a)The period and group of Ca (Calcium) is 2 (2A) and 4, respectively. The period and group of Ir (Iridium) is 9 (8B) and 6, respectively. The period and group of C (Carbon) is 14 (IVA) and 2, respectively. (b) Mg (Magnesium) is an alkali earth metal, whereas Li (Lithium) is a alkali metal. Co (Cobalt) is a transition metal. He (Helium) is a noble gas. F (Fluorine) is an halogen. (c) Ba (Barium) is a metal. N (Nitrogen) is a nonmetal. Si (Sillicon) is a metalloid.

STUDY CHECK

Answer the following questions: (a) Give the group and period of the following elements, and give the name: Cl. (b) Classify as alkali metal, alkali earth metal,

> transition metal, halogen or noble gas, and give the name: Ne. (c) Classify as metal, nonmetal or metalloid, and give the name: W.

`Classification of elements in terms of groups` Some of the groups in the periodic table (see Figure 3.1) have specific names such as alkali metals, alkaline earth metals, transition metals, chalcogens, halogens, or noble gases. Alkali metals are the group 1A elements: lithium (Li), sodium (Na), potassium (K), rubidium (Rb), cesium (Cs), and francium (Fr). Alkali elements are soft and shiny metals, and they are also good conductors of heat and electricity, with low melting points. Alkali earth metals are group 2A (2) elements: beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra). Transition metals are the elements from groups 3 to 12 and they are located in the middle of the table. Chalcogens are group 6A (16) elements: oxygen (O), sulfur (S), selenium (Se), tellurium (Te), and polonium (Po). Halogens are group 7A (17) elements: fluorine (F), chlorine (Cl), bromine (Br), iodine (I), and astatine (At). Halogens are very reactive elements. Finally, noble gases are group 8A (18) elements: helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), and radon (Rn). They are inert and rarely combine with other elements in the periodic table, like a noble family: have you ever met a royal?

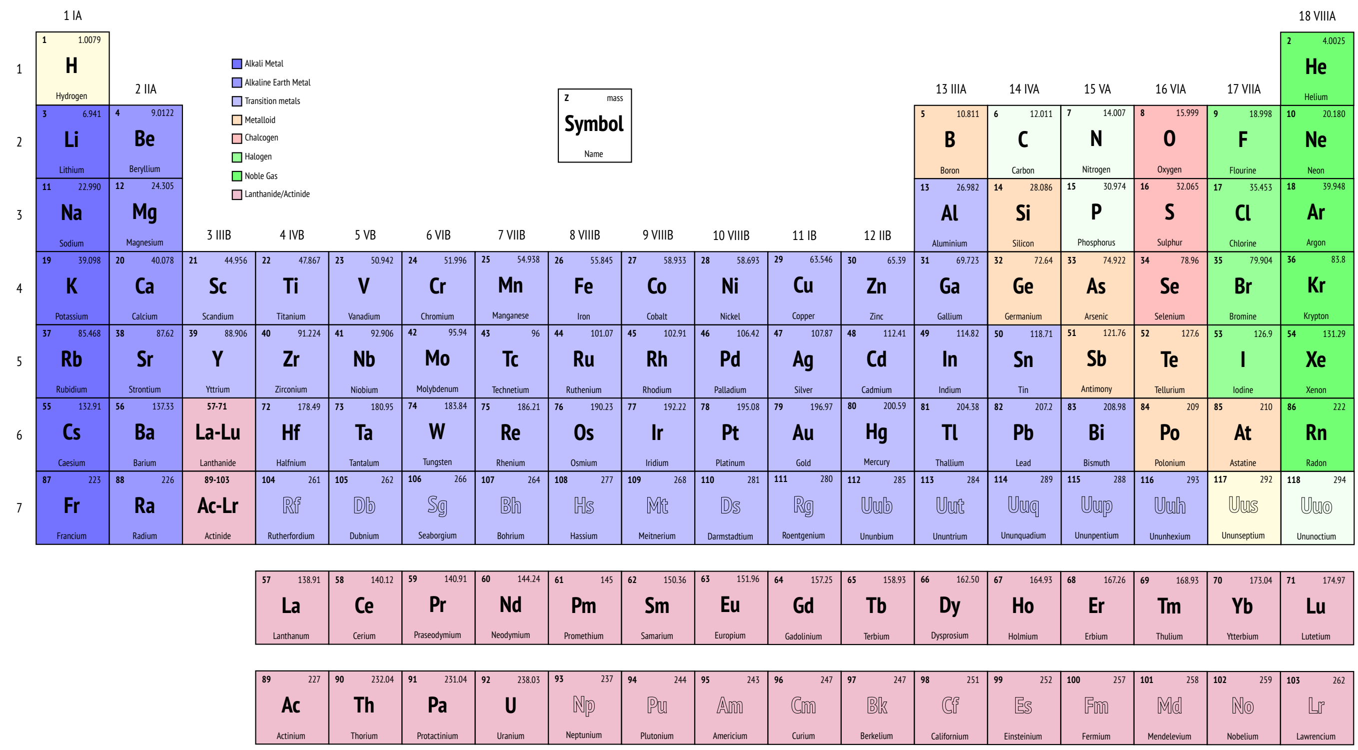
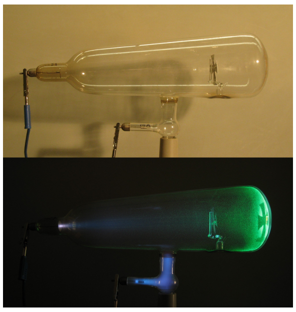


Figure 3.1 The periodic table of the elements

`How to classify Hydrogen` At first sight, hydrogen (H) may seem to be put in the wrong spot on the periodic table (see Figure 3.1). Although it is located at the top of Group 1A (1), it is not an alkali metal, as it has very different properties. Thus hydrogen does not belong to the alkali metals, being nonmetal.

3.2 Early experiments of the atom

▼A cathode tube



▼Millikan's apparatus

▼A plum pudding, with the electrons represented by the raisins

Scientists wondered about the nature of the atom and its structure for years. In a series of experiments carried out in the late nineteenth century, scientists such as J.J. Thomson, Henri Becquerel, and Ernest Rutherford gained insight into the nature and structure of the atom. These remarkable scientists and these creative experiments helped shape the view of the atom that we have nowadays.

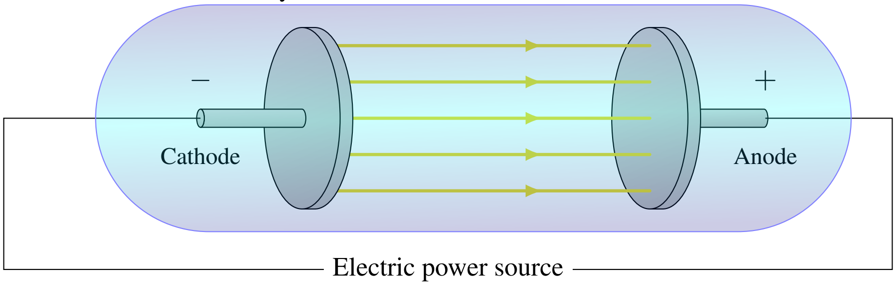


Figure 3.2 Cathode-ray tube made of two electrodes and a partially evacuated gas tube. The electrons, generated on the negatively charged electrode names cathode, excite the gas in the tube generating glow.

`Charge to mass ratio of an electron` The English researcher J.J. Thomson investigated electric discharges in partially evacuated tubes–tubes in which the air has been partially removed (See Figure 3.2); these tubes, made of a positive and negative electrode, are the base of old-fashion, bulky televisions. Thomson found that rays emanated from the negative electrode when applying high-voltage to these tubes. These rays were named cathodic rays as the negative electrode of the tube is called the cathode. Thomson also found that cathodic rays were repelled from the negative pole of an electric field. Hence, these rays were postulated to be a stream of negatively charged particles, now known as electrons. By studying the deflection of these rays by an electric field, Thomson was able to calculate the charge-to-mass ratio of an electron:

$$\boxed{\frac{e}{m_e} = -1.76 \times 10^8 C/g} \tag{3.1}$$

where:

e is the charge of an electron

m_e is the mass of an electron

Overall, the biggest of Thomson's discoveries was that all atoms are made of negatively charged particles. As atoms are charge-neutral, they are also made of positively charged particles. These observations led to a new atomic model, the *plum pudding model*, that envisioned atoms as a diffuse cloud of positive charges with negative electrons embedded in it. The name plum pudding comes from an English dessert that contains a raising spread. A different scientist, Robert Millikan, revealed the magnitude of the electric charge of the electron. Millikan used an apparatus that dispersed charged oil droplets falling under the influence of an electric field. Given the applied voltage and the droplet mass, Millikan was able to calculate the droplet charge. He found that the oil drop charge was always a whole number times the electron charge,

$$\boxed{e = 1.60 \times 10^{-19}\mathrm{C}} \tag{3.2}$$

where a Coulomb is a unit of charge. With the value of the charge-to-mass ratio of an electron and the electron charge, Millikan was also able to calculate the mass of an electron,

$$\boxed{m_e = 9.11 \times 10^{-31}\mathrm{kg.}} \tag{3.3}$$

`The atom nucleus` Ernest Rutherford carried out further experiments to validate the plum pudding model of the atom. He exposed a thin sheet of metal foil to α particles known to be massive and positively charged particles. According to the plum pudding model, the bulky α particles should have crashed through the thin foil and traversed through without being deflected. However, the results did not corroborate his expectations. Indeed, some particles traversed the film, whereas others were slightly deflected and some were strongly deflected at large angles. These observations did not corroborate the plum pudding model. However, they contributed to the creation of the modern atomic mode in which a large number of positive charges were concentrated at a point–called the nucleus– instead of being spread whereas the electrons move around the nucleus at large distances from it. The figure below represents the experiment carried out by Rutherford in which alpha particles were scattered on a think field made of gold.

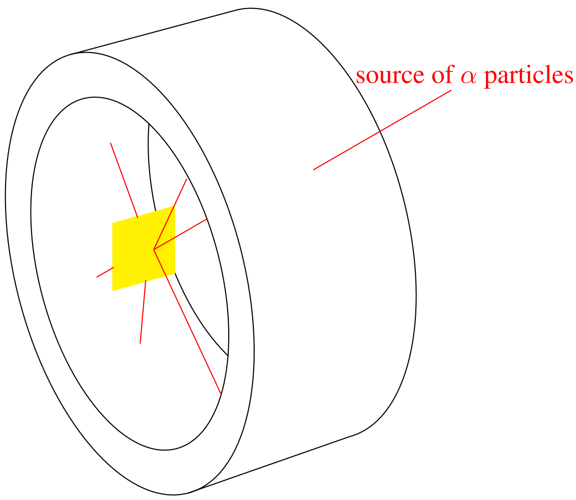


Figure 3.3 Rutherford's scattering: the elastic scattering of charged particles by the Coulomb interaction after a particle beam passed through a thin gold foil obstruction.

`Radiation` In the last part of the nineteenth century, scientists came to the discovery that some materials were able to produce high-energy radiation. Among the scientist working in this field, Henri Becquerel discovered that pitchblende, a mineral containing uranium, was able to produce an image on a photographic plate in the absence of light. In the early twentieth century, three different types of radiation were discovered: alpha radiation, beta radiation, and gamma radiation. Further studies revealed that gamma radiation was made of gamma particles, high-energy radiation, whereas beta radiation was made of high-energy electrons. Alpha particles were found to be positively charged, with a charge twice the charge of an electron and a mass 7300 times that of the electron. These particles were indeed Helium's nucleus, resulting from removing electrons from atoms of Helium.

3.3 The atom

Atoms are the smallest piece of an element that retains their characteristics. They are the building blocks of matter. This section covers the structure of the atom. You will learn how to calculate the number of subatomic particles that made an atom and how to differentiate atoms of an element–all atoms of an element are not equal.

`Atomic Structure` An atom is an electrically neutral, spherical entity made of a nucleus surrounded by negatively charged electrons. Atoms contain three atomic particles: the proton, neutron, and electron. Protons have a positive charge (+), whereas electrons carry a negative charge (–). Both electrons and protons have the same charge in magnitude but with opposite signs. Neutrons on the other hand are neutral, and they have no electrical charge. Protons and neutrons are located in the core of the atom, which is called the nucleus, and account for the mass of the atom. The only exception is the hydrogen atom, the smallest element, with just one proton in the nucleus. Electrons are delocalized in the exterior part of the atoms. They are not necessarily located in a specific spot and their existence spreads in the area next to the nucleus. Electrons move rapidly and are spread and held by nuclear attraction. Atoms are neutral without a charge as the number of electrons and protons are the same. Some atoms have a positive charge, resulting in removing electrons, and we call these cations. Others–called anions–can have a negative charge as a result of accepting a negatively charged electron. The mass of a proton or neutron is 2000 times larger than the mass of an electron and the atom's diameter is more than 10000 times the diameter of its nucleus. The nucleus is very dense being 99% of an atom's mass while occupying a small volume.

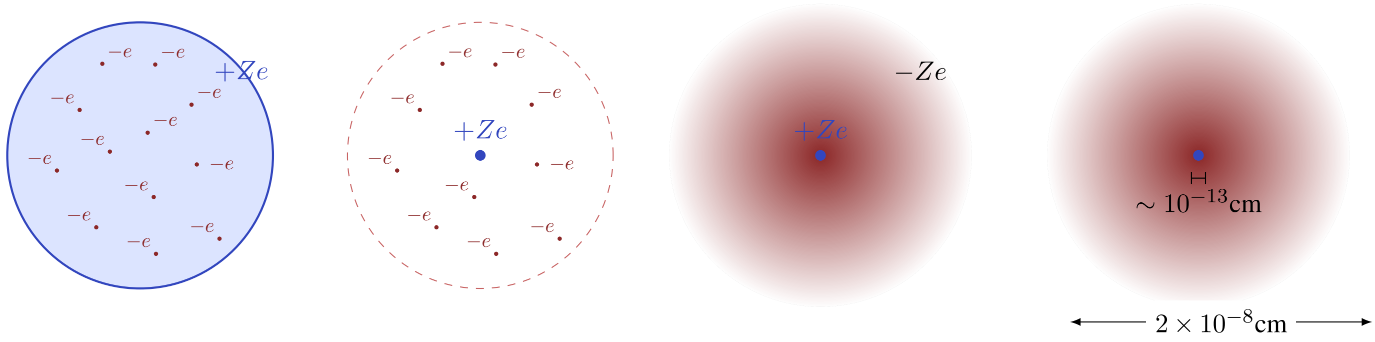
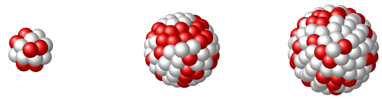


Figure 3.4 Three models of the atom. From left to right: the plum puding model, the updated plum pudding model according to Rutherford's observations, (two right images) the modern atomic model. Z is the atomic number of the atom.

Elements are made of atoms, and each atom of an element is characterized by an atomic number (Z) and a mass number (A). The atomic number (Z) of an element indicates the number of protons in an atom. This number can be easily located in the periodic table (see Figure 3.1). All atoms of an element have the same atomic number, whereas the atomic number of different elements differ. For example, Carbon has an atomic number of Z=6, whereas Oxygen has an atomic number of Z=8. The mass number (A) of an element indicates the combined number of protons and neutrons. Mass numbers can not be found in the periodic table. More importantly, different atoms of the same element can have different mass numbers. For example, a Carbon atom made of 6 neutrons and 6 protons has a mass number of A=12. Both A and Z for an atom X are indicated in the following form called isotope notation:

$$\texttt{Atomic and mass number } {}^{A}_{Z}\mathrm{X}$$

As an example, the notation ${}^{24}_{12}$Mg means that the atomic number of Mg is Z=12 and the mass number is A=24. Using the isotope notation, one can quickly identify the number of protons, neutrons, and electrons in an atom. As the atomic number is always indicated on the bottom part (e.g. Mg has 12 electrons). At the same time, the number of electrons and protons in a neutral atom is the same–neutral means an atom without a charge. The number of neutrons of an isotope can be computed by subtracting the atomic number from the mass number. Below you can find three different atoms, an atom of Carbon with 12 protons and neutrons, a larger atom of Calcium with 44 protons and neutrons, and an even larger atom of Gold with 197 protons and neutrons.



$^{12}_{6}C$ $^{44}_{20}Ca$ $^{197}_{79}Au$

Sample Problem 21

Calculate the number of protons, neutrons and electrons of the following atoms:

(a) $^{27}_{12}Mg$ (b) $^{22}_{10}Ne$ (c) $^{20}_{10}Ne$

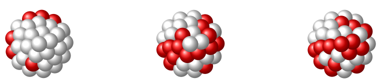
SOLUTION

(a) $^{27}_{12}Mg$ has 12 electrons (Z=12) and 12 protons as well (the number of electrons and protons are the same if the atom is neutral), and 15 neutrons, as 27-12=15. (b) $^{22}_{10}Ne$ has 10 electrons and 10 protons, and 12 neutrons. (c) $^{20}_{10}Ne$ has10 electrons and 10 protons, and 10 neutrons as well.

STUDY CHECK

Calculate the number of protons, neutrons and electrons of the following atoms:
(a) $^{32}_{16}S$ (b) $^{34}_{16}S$ (c) $^{36}_{16}S$

Isotopes All atoms of an element have the same atomic number but may differ in terms of mass number. Isotopes are atoms of the same element with different numbers of neutrons and therefore with different mass numbers but with the same atomic number. For example: $^{24}_{12}Mg$, $^{25}_{12}Mg$ and $^{26}_{12}Mg$ are three isotopes of Mg. $^{27}_{12}Mg$ is heavier than $^{24}_{12}Mg$ as it contains more neutrons and protons in the nucleus. Most elements occur in nature in a particular isotopic composition, and each of the isotopes has a specific proportional abundance. For example, the abundance of $^{24}_{12}Mg$ is 79%, and the abundance of $^{25}_{12}Mg$ and $^{26}_{12}Mg$ is 10% and 11%, respectively. This means, $^{24}_{12}Mg$ is more abundant than for example $^{26}_{12}Mg$.



$^{24}_{12}Mg$ $^{25}_{12}Mg$ $^{26}_{12}Mg$

Another example of isotopes can be found in Carbon, with two naturally occurring isotopes. In the case of charged atoms, we have the cations have fewer electrons than their corresponding atom, whereas anions have more electrons, both based on their charge. The mass of an atom is measured relative to the mass of an atomic standard, the Carbon-12 atom, whose mass is defined as 12 atomic units of mass, amu. For example, the mass of ^{1}H is 1.008 amus. The term atomic unit of mass has been renamed to dalton (Da). Therefore, the mass of ^{1}H is 1.008 amus or 1.008 Da. The atomic mass is a relative unit of mass equivalent to 1.66054×10^{24}g.

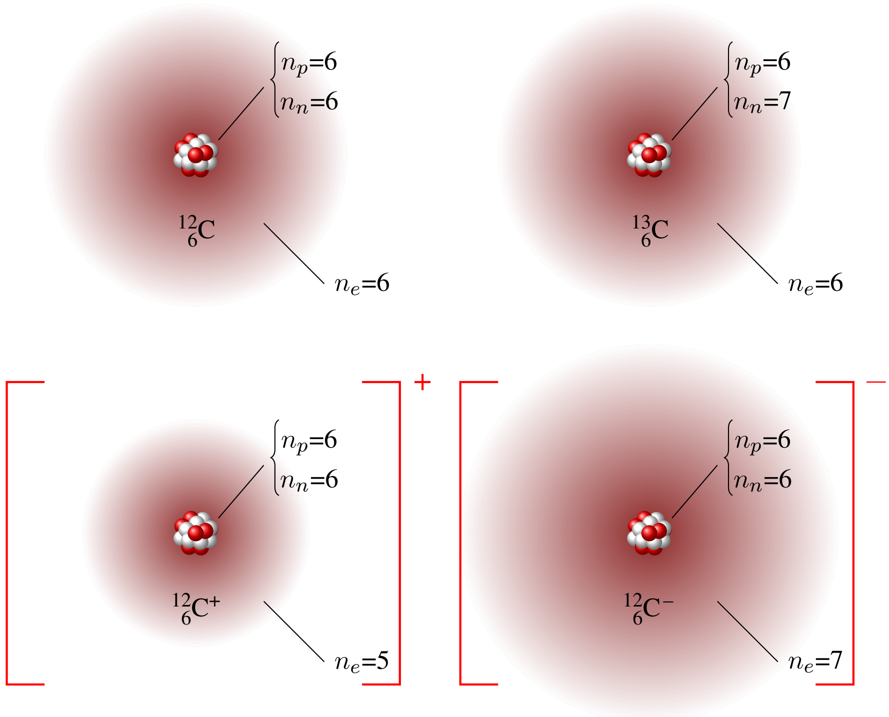


Figure 3.5 Representations of four different atoms, two neutral atoms on top and two ions on the bottom.

`Average atomic mass` As atoms are made of numerous isotopes–this means different atoms of the same element but with a different number of neutrons and hence different weights. The average atomic mass (also called atomic weight) represents the mass of the atoms of an element and results from all existing isotopes taking into account their abundance. It is the average of the masses of the naturally occurring isotope weighted according to their abundance expressed in atomic mass units or daltons. We can think of % *relative abundance*, and for example, the % relative abundance of ^{1}H is 99%. But we can also think of *fractional abundance*, that in the case of ^{1}H would be 0.99. For an element with n isotopes each with different masses (A_1, A_2, ..., A_n) and different fractional abundances (f_1, f_2, .., f_n), the atomic mass is given by

$$\text{Atomic mass} = \sum_{i=1}^{n} A_i \cdot f_i = A_1 \cdot f_1 + A_2 \cdot f_2 + \cdots + A_n \cdot f_n$$

Note that when adding the fractional abundances of all isotopes, one should obtain a value of one:

$$\sum_{i=1}^{n} f_i = f_1 + f_2 + \cdots + f_n = 1$$

Atomic masses can be simply found in any periodic table (see Figure 3.1) for each element. For example, the atomic mass of oxygen (O) is 15.999 amu and the atomic mass of nitrogen (N) is 14.007 amu. The atomic mass found in the periodic table is an average that results from including the mass of the different isotopes and their abundance. Table 3.1 lists the relative abundance of a series of common isotopes.

Sample Problem 22

Naturally occurring copper (Cu) consists of 69.17% ^{63}Cu and 30.83% ^{65}Cu. The mass of ^{63}Cu is 62.939598 amu, and the mass of ^{65}Cu is 64.927793 amu.

What is the atomic mass of copper?

SOLUTION

The weighted average is the sum of the mass of each isotope times its fractional abundance. We have that the isotope ^{63}Cu has a mass of 62.939598 amu and an abundance of 69.17%, that is the same as 0.6917. At the same time, the isotope ^{65}Cu has a mass of 64.927793 amu and an abundance of 0.3083. After adding both contributions, we have:

$$62.939598\,amu \times \frac{69.17}{100} + 64.927793\,amu \times \frac{30.83}{100} = 63.55\,amu$$

We can use the table below to obtain the final result:

Isotope	m_i (amu)	f_i	$m_i \times f_i$
^{63}Cu	62.939598	0.6917	43.53
^{65}Cu	64.927793	0.3083	20.02
Average atomic mass (amu)			$= \sum m_i \times f_i$ =63.55

STUDY CHECK

Lithium is made up of two isotopes, Li-7 (7.016003 amu) and Li-6 (6.015121 amu). Calculate the percent abundance of each isotope knowing that lithium's atomic weight is 6.94 amu.

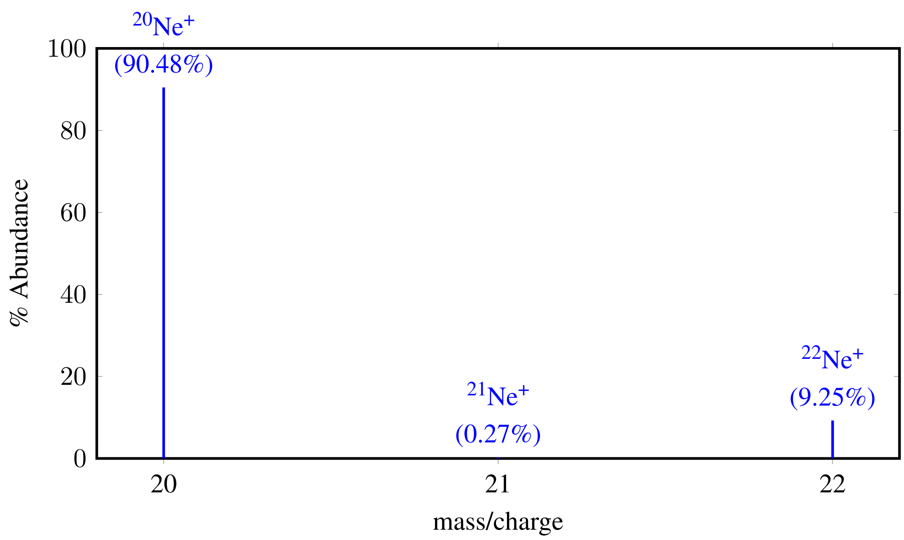


Figure 3.6 Mass spectra of Neon with three peaks corresponding to three different isotopes with different relative abundance.

Mass spectrometry Mass spectroscopy is a technique used to determine the isotopic composition of an element. With this technique, we can measure the relative mass and abundance of small atomic particles. In this technique, high-energy electrons collide with atoms or molecules to form ionized atoms. For example, if Ne would be analyzed, high energy electrons would produce Ne^+ characterized by its mass charge ratio, m/e. Different isotopes would have different m/e ratios. The positively charged particles produced after the impact would be attracted toward a series of negatively charged plates with slits in them. Some of these particles would pass through the slits into an evacuated tube exposed to the effect of a magnetic field. As the particles enter the evacuated region their paths are bent so that the lightest particles (low m/e) are more deflected than the heaviest particles (high m/e). Finally, the particles would stick to a detector recording their relative position and abundance. The spectrometer also provided the mass ratio of an isotope with respect to the mass standard, ^{12}C.

Table 3.1 Isotope abundance of some elements

Element	Isotope	% Abundance	Element	Isotope	% Abundance
Hydrogen	^{1}H	99.9885%	Silicon	^{28}Si	92.2297%
	^{2}H	0.0115%		^{29}Si	4.6832%
Helium	^{3}He	0.000137%		^{30}Si	3.0872%
	^{4}He	99.999863%	Sulfur	^{32}S	94.93%
Lithium	^{6}Li	7.59%		^{33}S	0.76%
	^{7}Li	92.41%		^{34}S	4.29%
Boron	^{10}B	19.9%		^{36}S	0.02%
	^{11}B	80.1%	Chlorine	^{35}Cl	75.78%
Carbon	^{12}C	98.93%		^{37}Cl	24.22%
	^{13}C	1.07%	Argon	^{36}Ar	0.3365%
Nitrogen	^{14}N	99.632%		^{38}Ar	0.0632%
	^{15}N	0.368%		^{40}Ar	99.6003%
Oxygen	^{16}N	99.757%	Potassium	^{39}K	93.2581%
	^{17}O	0.038%		^{40}K	0.0117%
	^{16}O	0.205%		^{41}K	6.7302%

3.4 The nature of light

Light–also called electromagnetic radiation–is a form of energy. When talking about light we normally refer to visible radiation. However, there are many different types of radiation. Think about the light coming from a bulb, the radiation that warms up your food in a microwave, or even when you warm up a pizza in the oven. This section will cover the properties of light.

▼Standing waves of a guitar

▼Traveling waves

▼Light is a wave

`Light as a wave` Light behaves as a wave. Waves are characterized by their frequency, wavelength, and amplitude (see figure 3.7). The wavelength of a wave (λ, lambda) is the distance between identical points on successive waves (or successive peaks). The frequency of a wave (ν, nu) is the number of waves that pass through a particular point in one second. The amplitude (A) of a wave is the vertical distance from the zero to the top of the peak, or from the zero to the bottom of the peak. The amplitude of a wave is related to the intensity of the radiation. The speed of light through the vacuum is 3×10^8m/s. However, the speed of light depends on the medium and light tends to slow down when traveling in a medium different than a vacuum.

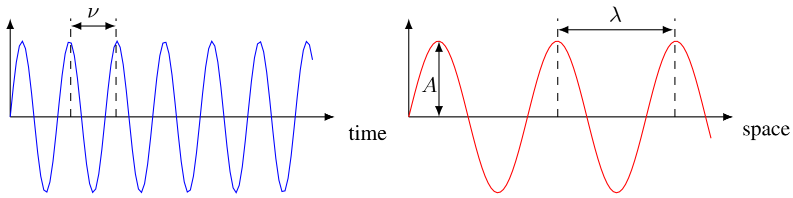


Figure 3.7 Properties of waves. Waves are characterized by its frequency, wavelength and amplitude and in the vacuum they travel at the speed of light.

The electromagnetic spectrum of light Visible light consists of electromagnetic waves, which have an electric field and magnetic field component (see Figure 3.8). These two components share the same wavelength, frequency, and speed but are perpendicular to each other.

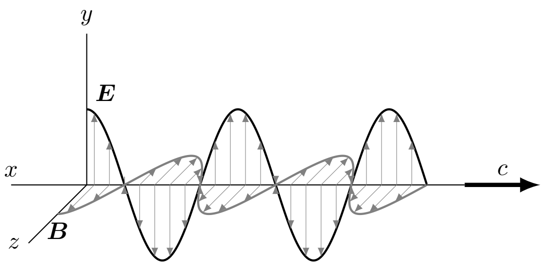


Figure 3.8 The electric (E) and magnetic (B) field components of an electromagnetic wave. Both fields are perpendicular.

The double-slit experiment: light diffraction The double-slit experiment was intended to demonstrate the wave nature of light. When a source of light passes through a narrow opening called a slit a bright spot is generated on the other side of the slit (see Figure 3.9). When a source of light passes through two slits surprising results arise. One would expect to see two bright spots, one per slit. However, what you really would see would be a series of bright spots and dark spots, resulting from the interference of light. As light is a wave it can interfere and light plus light does not always gives more light, and can sometimes genera darkness. The bright spots result from the constructive interference of the light waves whereas the dark spots result from the destructive interference. Overall, waves propagate energy and the results of the propagation can be more light or less light depending on how these waves interfere.

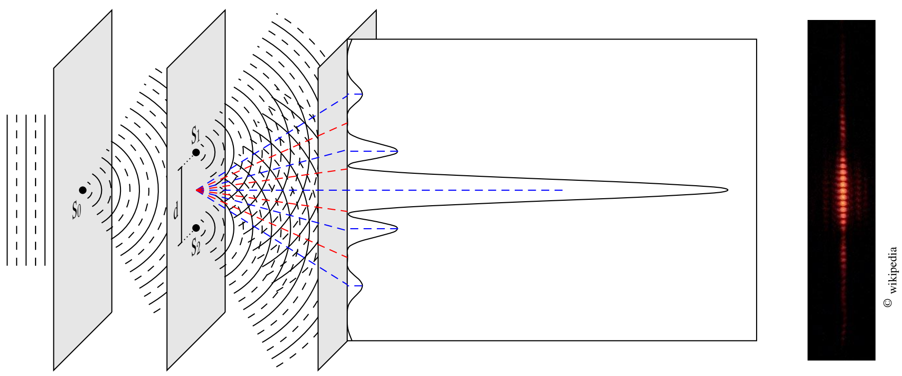


Figure 3.9 The double-slit experiment demonstrating the wave nature of light. Using one slit leads to a single bright spot. Using two slits leads to a set of patterns of light and darkness resulting from the interference of light. Red lines represent destructive interference whereas blue lines represent constructive interferences.

`Types and color of radiation` Depending on its frequency–or on its wavelength–radiation can be classified as gamma rays, x-rays, ultraviolet (UV), visible, infrared (IR), microwaves, or radio waves (see Table **??** and Figure 3.10). For example, radiation with a wavelength of 10^{-2} nm belongs to gamma rays radiation, whereas radiation with a wavelength of 10^4 nm belongs to the Infrared. Gamma rays are the most energetic type of radiation, whereas radio waves are the less energetic waves. At the same time, radio waves have the largest wavelength. Light does not always have color. Only a small range of wavelengths belong to visible radiation and the visible spectrum corresponds to the set of visible frequencies. This means you will not be able to see, for example, IR radiation or gamma rays. The color of the radiation is also dependent on the wavelength–of the frequency as both are related–and for example $\lambda = 450$ nm will be blue light. Ultraviolet radiation is the most energetic visible radiation whereas infrared waves are the less energetic waves of the visible spectrum.

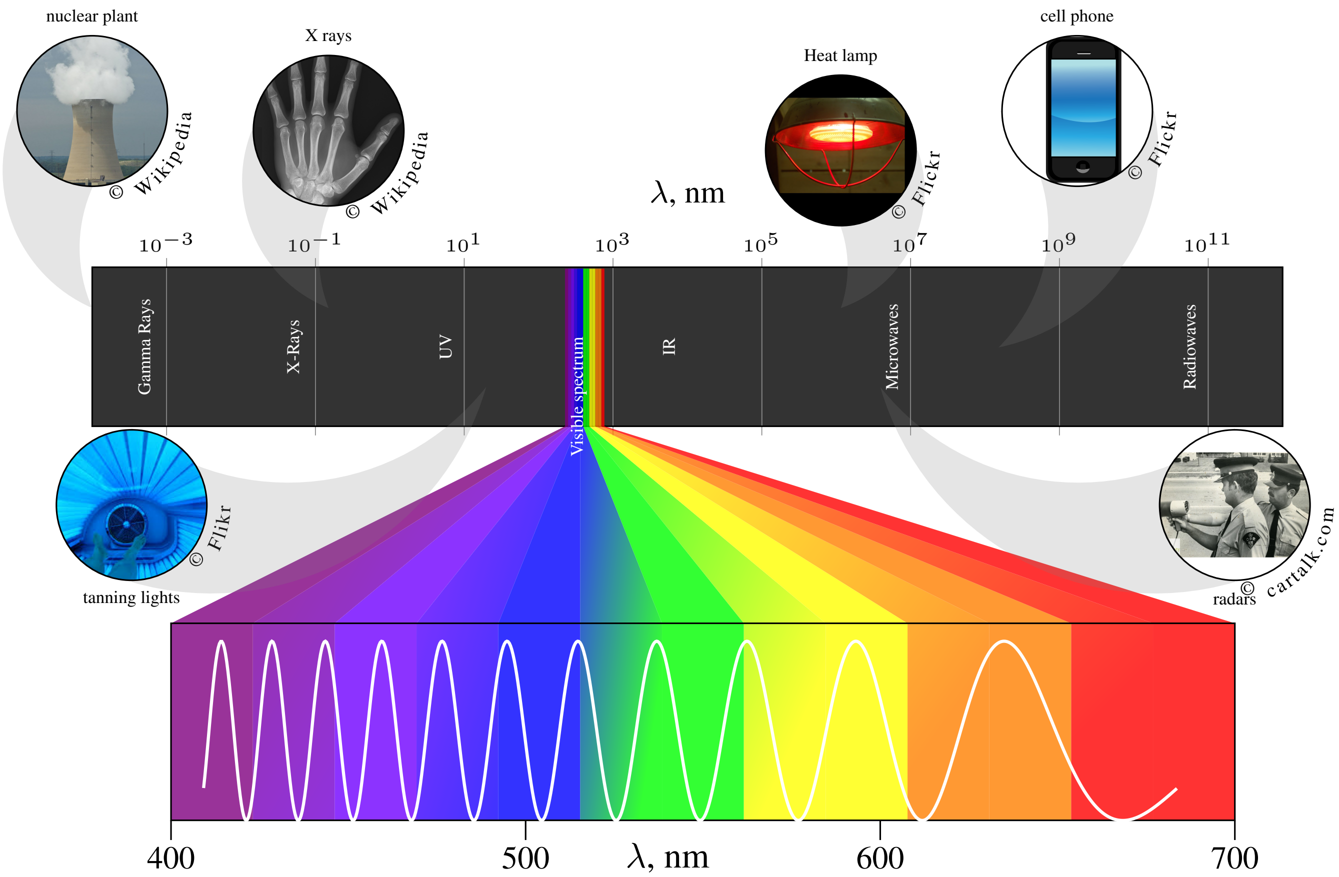


Figure 3.10 Spectrum of the electromagnetic radiation, from gamma rays (shortest wavelength) to radio waves (longest wavelength). The visible part of the spectrum ranges from 400 nm (violet) to 700 nm(red).

Sample Problem 23

Indicate: (a) the color of a radiation with $\lambda = 650nm$; (b) the type of a radiation with $\lambda = 10^5$ nm.

SOLUTION

We can answer the first questions by inspecting the figure above we can see that $\lambda = 650nm$ corresponds to red radiation. To answer the second question we will also use the figure above, where we can see that $\lambda = 10^5$ nm belongs to the infrared.

STUDY CHECK

Indicate: (a) the color of a radiation with $\nu = 400$ nm; (b) the type of a radiation with $\nu = 10^{10}$ Hz.

Answer: (a) violet; (b) Microwaves.

Table 3.3 Types and color of radiation

Type of radiation	ν (Hz)	Color of radiation	λ (nm)
Gamma	$>3 \times 10^{19}$	Violet	380-450
X-rays	$3 \times 10^{19} - 3 \times 10^{16}$	Blue	450-485
UV	$3 \times 10^{16} - 8 \times 10^{14}$	Cyan	485-500
UV-visible	$4 \times 10^{14} - 8 \times 10^{14}$	Green	500-565
IR	$4 \times 10^{14} - 4 \times 10^{11}$	Yellow	565-590
MicroW	$3 \times 10^{11} - 3 \times 10^{8}$	Orange	590-625
RadioW	$3 \times 10^{8} - 3 \times 10^{3}$	Red	625-740

3.5 The atomic line spectra

This section will explain the atomic spectra of atoms and in particular, hydrogen. Atoms emit light, but not any type of light as they emit specific frequencies of radiation. The atomic spectrum is a representation of the different wavelengths of the radiation emitted–or absorbed–by an atom. This section will gain insight into the reasons for the emission of specific frequencies of light and will introduce the Bohr model that justifies the lines in the atomic spectrum of hydrogen.

`Spectrum of atoms` The absorption spectrum of an atom is a representation of the different frequencies at which an atom absorbs or emits radiation. Each atom has a distinctive emission spectrum. Some of these lines correspond to the visible spectrum, that is, they can be seen. Others correspond to higher or lower parts of the spectrum. This section will cover the spectrum of hydrogen.

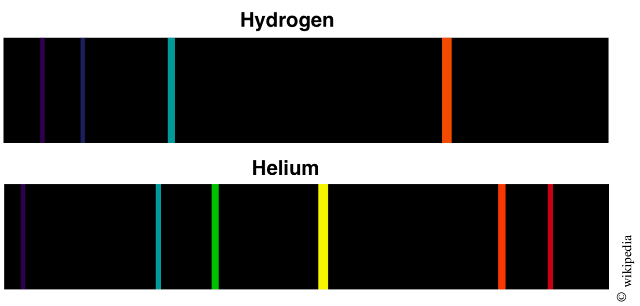


Figure 3.11 Atomic spectrum of two atoms showing the lines corresponding to the visible.

`Atomic line spectrum of hydrogen` Newton showed that sunlight (white light) is composed of various components of different colors. A similar type of radiation called emission spectra can be produced by heating a substance. Think for example of a hot piece of metal. Both the sun and a heated piece of metal have in common the fact that their spectrum is continuous and contains all wavelengths of visible light. We can achieve a similar effect by applying a high-voltage electrical discharge to a gas. The atomic line spectrum of a gas is a set of lines on a black (or sometimes white) background (see Figure 3.11). These lines correspond to radiation emitted (or absorbed) by atoms. For the case of hydrogen (see Figure 3.12), some of these lines correspond to the visible spectrum, that is, have color–these are called the Balmer series. Other lines correspond to other parts of the spectra of radiation. This spectrum is historically important and was used to understand the structure of the electrons in the atom. In contrast to the sunlight spectrum, the atomic spectrum of a gas is not continuous but quantized.

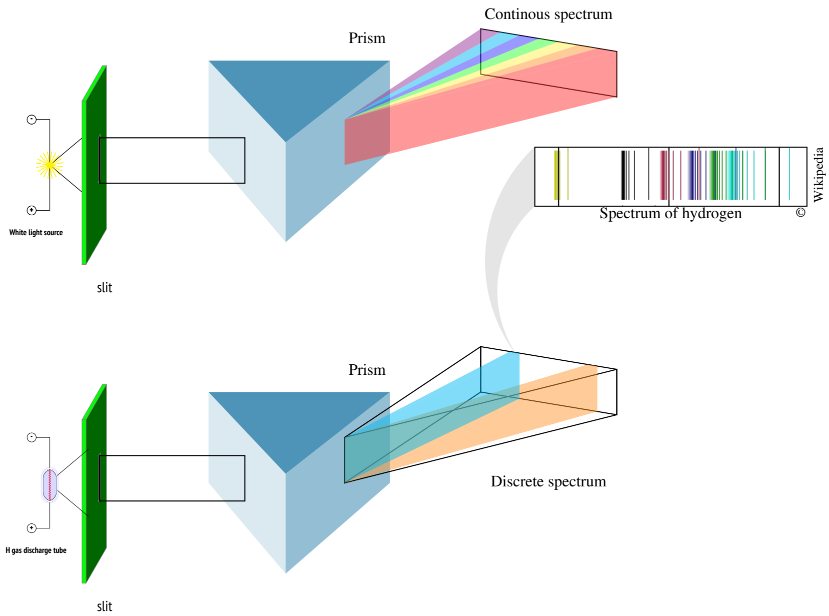


Figure 3.12 The spectrum of hydrogen: white light contains radiation of all colors of the visible spectrum whereas light coming from hydrogen contains a quantized series of lines.

3.6 Atomic orbitals

The precise location of electrons in the atoms is unattainable. Differently, the electron location is described in terms of probability. An orbital is a three-dimensional area with the highest probability of finding an electron. Orbitals have different shapes and energy, as electrons in those have different stability. Some orbitals have a spherical shape, whereas others are lobular.

s `orbitals` There is a single s subshell in every shell and each s subshell contains only a single orbital. For example, in the first energy level, there is a 1s orbital, whereas, in the second energy level, there is a 2s orbital. All s orbitals have an overall spherical shape with increasing size depending on the principal quantum number n (see Figure 3.13). For example, the $1s$ orbital is smaller than the $2s$, and so on.

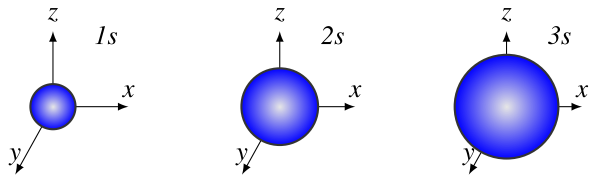


Figure 3.13 Surface plots of three s orbitals.

p `orbitals` There are three different p orbitals and each subshell with n larger than 2 has a p orbital. These three p orbitals are labeled p_x, p_y, and p_z, where the subindexes represent the direction of the axis along which each orbital is oriented. All three p orbitals have the same size, shape, and energy. The larger the principal quantum number the larger the size of the orbital, and for example, a $3p_x$ is larger than a $2p_x$. The boundary surface diagrams of p orbitals expose their shape in the form of two lobes, and positive and one with a negative sign (see Figure 3.15).

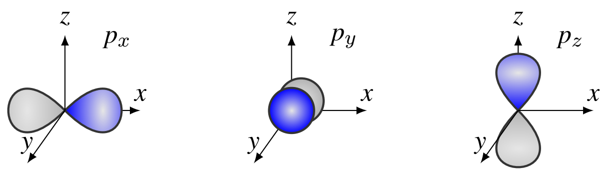


Figure 3.14 Surface plots of the three p orbitals.

d `and` f `orbitals` There are five d orbitals labeled as d_{xy}, d_{xz}, d_{yz}, d_{z^2} and $d_{x^2-y^2}$. The subindexes in the labels are related to the orientation of the orbital in space. None of the d_{xy}, d_{xz}, d_{yz} orbitals cross any of the axis, but d_{z^2} and $d_{x^2-y^2}$ do. In particular, the lobes of the d_{xy} orbitals are located in the xy plane, whereas $d_{x^2-y^2}$ cross the x and y axis and the d_{z^2} crosses the z axis. All five d orbitals have the same energy. These orbitals increase in size based on the principal quantum number and the $3d$ orbitals are smaller than $4d$. There are no d orbitals in the first or second shell. Regarding the f orbitals, there are seven orbitals with very complex shapes (see Figure 3.15).

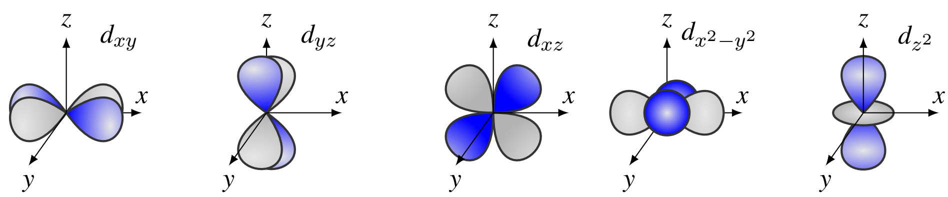


Figure 3.15 Surface plots of the three d orbitals.

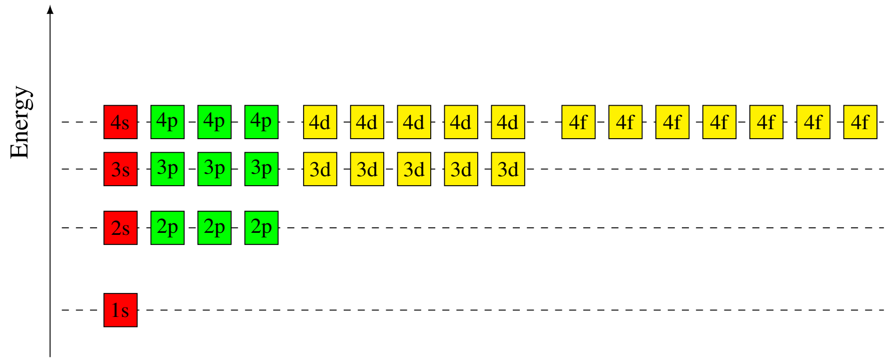


Figure 3.16 Energy of the different hydrogen orbitals. s orbitals have the lowest energy and hence are the most stable. Orbitals in the same shell ($2p$, $2s$) have the same energy. The larger the principal quantum number the more positive the energy, the less stable the orbital.

3.7 Electronic configuration of an atom

Atoms have in general many electrons. These electrons are arranged in the atom in a very specific way creating what we know as an electronic structure. You want to think about the electronic configuration of an atom as a code that tells you the orbital location of each electron in the atom. There are two ways to present electronic configurations. One is called full electronic configuration (for example $1s^2 2s^1$) and the other one is called condensed electronic configuration (for example $[He]2s^1$). The full configuration displays all orbitals in an atom, whereas the abbreviated configuration only displays the valence electrons–these electrons are less-tied to the nucleus–and the noble gas core.

`Electron energy levels and sublevels` The electrons in an atom are arranged in different energy levels. Some levels have lower energy and the electron in these levels are close to the nucleus being also strongly attached to it, whereas other levels have higher energy and the electrons in these levels are less attached to the nucleus. Still, each electron in an energy level has the same energy. The energy levels are labeled with a number n that equals a single number such as 1, 2, 3, and so on. The first energy level is $n = 1$ and never $n = 0$–think of this as an apartment in a building, the first floor is flour one. For example, all electrons in level one $n = 1$ have the same energy. There is a limit to the number of electrons in an energy level and we call this occupancy. Only a few electrons can occupy the lower energy levels, while more electrons can be accommodated in higher energy levels. Level one can only fit two electrons, whereas level two can fit a total number of eight electrons. The maximum number of electrons allowed in any energy level is calculated using the formula

$$2n^2 \tag{3.4}$$

in which n is the energy level. You can see by using this formula that for example, the third level can accommodate 18 electrons. Each energy level consists of one or more sublevels, which contain electrons with identical energy. The number of sublevels in each level corresponds to n. For example, in the first energy level ($n = 1$) we have only one sublevel, whereas in the third energy level ($n = 3$) we have three sublevels.

Sample Problem 24

How many electrons can you fit in the energy level n=3.

SOLUTION

We will use the formula $2n^2$ that gives the number of electrons that fit in a energy level n. As $n = 3$, we can fit 18 electrons in this level. Remember, the larger the energy level the more electrons we can fit.

STUDY CHECK

At a given energy level you can fit 162 electrons. Identify the energy level.

▶Answer: $n = 9$.

`Orbital Filling: the aufbau principle` Atoms in general contain numerous orbitals and each orbital should be filled with electrons. In every orbital, you can fill only a maximum number of electrons. For example, in a s orbital you can place a maximum of two electrons. That is why you will find s^1 orbitals and s^2, with the latest being filled with electrons. In a p orbital you can place a maximum of six electrons and in a d orbital a maximum of ten. Finally, in a f orbital you can place fourteen or fewer electrons. For example, the orbital notation p^2 is correct as in p orbitals you can place six or fewer electrons. In this case, this orbital still has space to accept more electrons. Differently, the notation d^{12} is incorrect, as in d orbitals you can fit ten or fewer electrons and never twelve. The Aufbau (build up) principle states that the electronic configuration of an atom can be obtained by adding one by one all electrons in the element. In order to fill the orbitals you should follow Figure 3.17. You start from the top of the table and follow the arrows that indicate the orbitals ordering. For example, the first orbital to be filled will be $1s$. After that, you should fill in $2s$ and $2p$. After that you should fill $3s$, $3p$, $4s$, $3d$, and $4p$. There is a maximum number of electrons that can occupy each orbital. An s orbital holds a maximum of 2 electrons. A p orbital takes up to 6 electrons, a d orbital can hold up to 10 electrons, and an f orbital holds a maximum of 14 electrons. An orbital can be completely filled with electrons, partially filled, or empty. For example, a $3s^1$ is half-filled with one electron and $2p^6$ is completely filled. Another example, a $3d$ orbital is empty and can accommodate a maximum of 10 electrons.

`Full electron Configuration` The full electron configuration of an atom is obtained by placing the total number of electrons of the atom in different orbitals with increasing energy. For example, the electron configuration for helium is written as

$$\text{He } 1s^2$$

and the one for Li is

$$\text{Li } 1s^2 2s^1$$

First, how do we know the total number of electrons in an atom? That is the same as the atomic number and is indicated in the periodic table. Look for the element and the atomic number is on the top left side of the element. For example, the atomic number of hydrogen is one, and the number of electrons in He is two. Similarly, nitrogen has seven electrons. Second, how do we know which orbitals need to be filled? Figure 3.18 shows the orbital order. You need to start from the top of the image, from orbital $1s$, and proceed to the next arrow, starting from the end of the arrow. This way, after $1s$ goes $2s$ and then $2p$, $3s$, $3p$, $4s$, and $3d$. Mind that every s orbital can only fit two

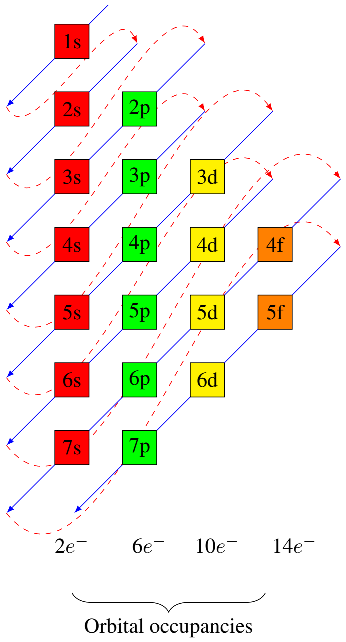


Figure 3.18 Orbital filling order.

electrons, and p orbitals can fit six electrons, and so on. The following example will help you construct the electron configuration for a given atom.

Sample Problem 25

Obtain the full electronic configuration of C.

SOLUTION

The atomic number of C is Z=6 and that means C has 6 electrons. The orbital order from Figure **??** is: $1s$,$2s$, $2p$, $3s$, etc. Each s orbital can fit two electrons, whereas the occupancy of the p orbitals is six electrons. Hence the electronic configuration of C is: $1s^22s^22p^2$. The s orbitals are all filled, whereas the p orbital is only occupied with two electrons.

STUDY CHECK

Obtain the full electronic configuration of Ni.

▸Answer: $1s^22s^22p^63s^23p^64s^23d^8$.

Abbreviated Electron Configuration and orbital diagrams Let us compare the electronic configuration of Ni and Ar. We have:

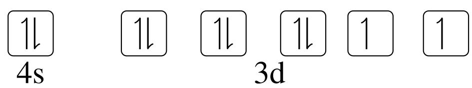
$$\text{Ni } 1s^22s^22p^63s^23p^64s^23d^8$$
$$\text{Ar } 1s^22s^22p^63s^23p^6$$

We call this configuration the *full electronic configuration*. If you look carefully, you will see that both configurations resemble, and in particular, the configuration of Ni is the same as the configuration of Ar with two extra orbitals added. We say Ni has the core of Ar and 10 electrons on its valence. We can represent this fact by indicating Ar with brackets:

$$\text{Ni } 1s^22s^22p^63s^23p^64s^23d^8 \qquad \text{(full electron configuration)}$$
$$\text{Ni } [Ar]4s^23d^8 \qquad \text{(abbreviated electron configuration)}$$

The orbital diagrams are boxed diagrams indicating the valence electrons such as:

Ni [Ar] $4s^23d^8$



We call this last configuration the *abbreviated electronic configuration*. You can figure out faster the abbreviated electronic configuration by looking for the noble gas on the table on the row above the element, and the period (row on the table) of the element. Ni is located in period number four and the noble gas above this period is Ar. At the same time, Ar has 18 electrons. That will give you the core $[Ar]$ with 18 electrons, and the remaining 10 electrons (Ni has 28 electrons) starting with the orbital $4s$, according to period four. The electrons in the noble gas core are called *core electrons*, whereas the rest of the electrons are known as *valence electrons*. For the case of Ni, we have that it has 18 core electrons in the Ar core and 10 electrons in the valence. Let us work on another example: Cd. It has 48 electrons and is located in group 5. The novel gas in the group above is Kr with 36 electrons. The core will be Kr–the novel gas on the period above–and we start right away filling $5s$ electrons–Cd belongs to period five. In the valence electrons, we will place 12 electrons. Hence, the abbreviated configuration will be $[Kr]5s^24d^{10}$.

Sample Problem 26

Obtain the full and abbreviated electronic configuration of Silver (Ag, Z=47) using the fact that the 4d and 5s orbitals switch order, being the d orbital filled first.

SOLUTION

The atomic number of Ag is Z=47 and that means Ag has 47 electrons. The orbital order is: $1s$, $2s$, $2p$, $3s$, etc. Each s orbital can fit two electrons, whereas the occupancy of the p orbitals is six electrons, and d orbitals can fit 10 electrons. We will use the fact that the 4d orbital fills before the 5s. Hence the full electronic configuration of Ag is: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^1$. As Kr is the noble gas on top with 36 electrons, the abbreviated electronic configuration of Ag is: $[Kr]4d^{10}5s^1$. Ag has 36 core electrons and 11 valence electrons.

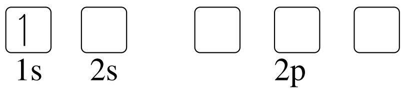
STUDY CHECK

Obtain the abbreviated electronic configuration of Cobalt (Co, Z=27).

Answer: $[Ar]4s^2 3d^7$.

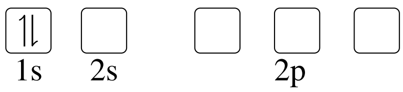
Hund's rule Let us build up the electron configuration of a series of atoms. Starting with hydrogen, with one electron, we have that only the $1s$ orbitals will be filled:

H $1s^1 2s 2p$



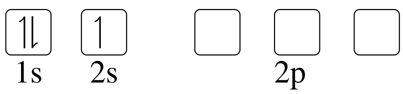
For the case of helium, with two electrons, we have that based on Pauli's principle both electrons have different quantum numbers and to differentiate this we will represent the pair of electrons with arrows in opposite directions. We say both electrons are paired:

He $1s^2 2s 2p$

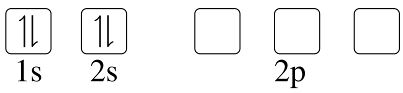


Now, in the case of Lithium and Beryllium, we have a similar situation. Lithium has a single unpaired electron and Beryllium has a set of paired electrons in the $2s$ orbital:

Li $1s^2 2s^1 2p$

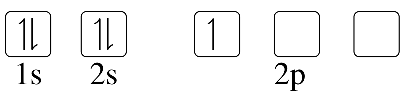


Be $1s^2 2s^2 2p$



The next element, Boron, has one electron in a p orbital. As all p orbitals are degenerate–they have the same energy–we can place that single electron in any of the p orbitals. Normally, we use the one on the left:

B $1s^2 2s^2 2p^2$



For the case of Carbon, as we need to place two electrons in the p orbitals, Hund's rule states that we need to place the electron maximizing the number of unpaired electrons. In other words, the second electron goes into a separate p orbital, just like below:

$$\begin{array}{lccccc} \text{C } 1s^22s^22p^2 & \boxed{\uparrow\downarrow} & \boxed{\uparrow\downarrow} & \boxed{\uparrow} & \boxed{\uparrow} & \boxed{} \\ & \text{1s} & \text{2s} & & \text{2p} & \end{array}$$

Similarly, in the case of Nitrogen, we have:

$$\begin{array}{lccccc} \text{N } 1s^22s^22p^3 & \boxed{\uparrow\downarrow} & \boxed{\uparrow\downarrow} & \boxed{\uparrow} & \boxed{\uparrow} & \boxed{\uparrow} \\ & \text{1s} & \text{2s} & & \text{2p} & \end{array}$$

Now, for the case of oxygen, as we cannot place that extra electron in a separate p orbital we have to start pairing electrons:

$$\begin{array}{lccccc} \text{O } 1s^22s^22p^4 & \boxed{\uparrow\downarrow} & \boxed{\uparrow\downarrow} & \boxed{\uparrow\downarrow} & \boxed{\uparrow} & \boxed{\uparrow} \\ & \text{1s} & \text{2s} & & \text{2p} & \end{array}$$

Finally, in the case of Neon, we have the whole second shell filled with electrons:

$$\begin{array}{lccccc} \text{Ne } 1s^22s^22p^6 & \boxed{\uparrow\downarrow} & \boxed{\uparrow\downarrow} & \boxed{\uparrow\downarrow} & \boxed{\uparrow\downarrow} & \boxed{\uparrow\downarrow} \\ & \text{1s} & \text{2s} & & \text{2p} & \end{array}$$

Sample Problem 27

Obtain the orbital diagram for oxygen given the electron configuration: $[He]2s^22p^4$

SOLUTION

The orbital diagram should include an s and a p sublevel. The s sublevel contains a single orbital represented by a box, whereas the p sublevel contain three different orbitals, and three boxes. The s orbital is fully occupied with two electrons, whereas the p sublevel contain only four electrons. The first three p electron will occupy separate boxes while the fourth electron will occupy the first box with opposite spin.

$$\begin{array}{lcccc} \text{O } [He] & \boxed{\uparrow\downarrow} & \boxed{\uparrow\downarrow} & \boxed{\uparrow} & \boxed{\uparrow} \\ & \text{2s} & & \text{2p} & \end{array}$$

STUDY CHECK

Obtain the orbital diagram for Li given the electron configuration: $[He]2s^1$.

▶Answer: $\boxed{\uparrow}$

3.8 Periodic Trends

Some atomic properties are periodic, which means that they follow certain trends in the periodic table. These properties increase or decrease across a period and then the trend is repeated in each successive period or group. We will consider here five atomic properties: the effective charge, atomic (or ionic) size, the ionization energy, the metallic character, and the electronegativity. Figure 3.22 summarizes the different periodic trends across the periodic table.

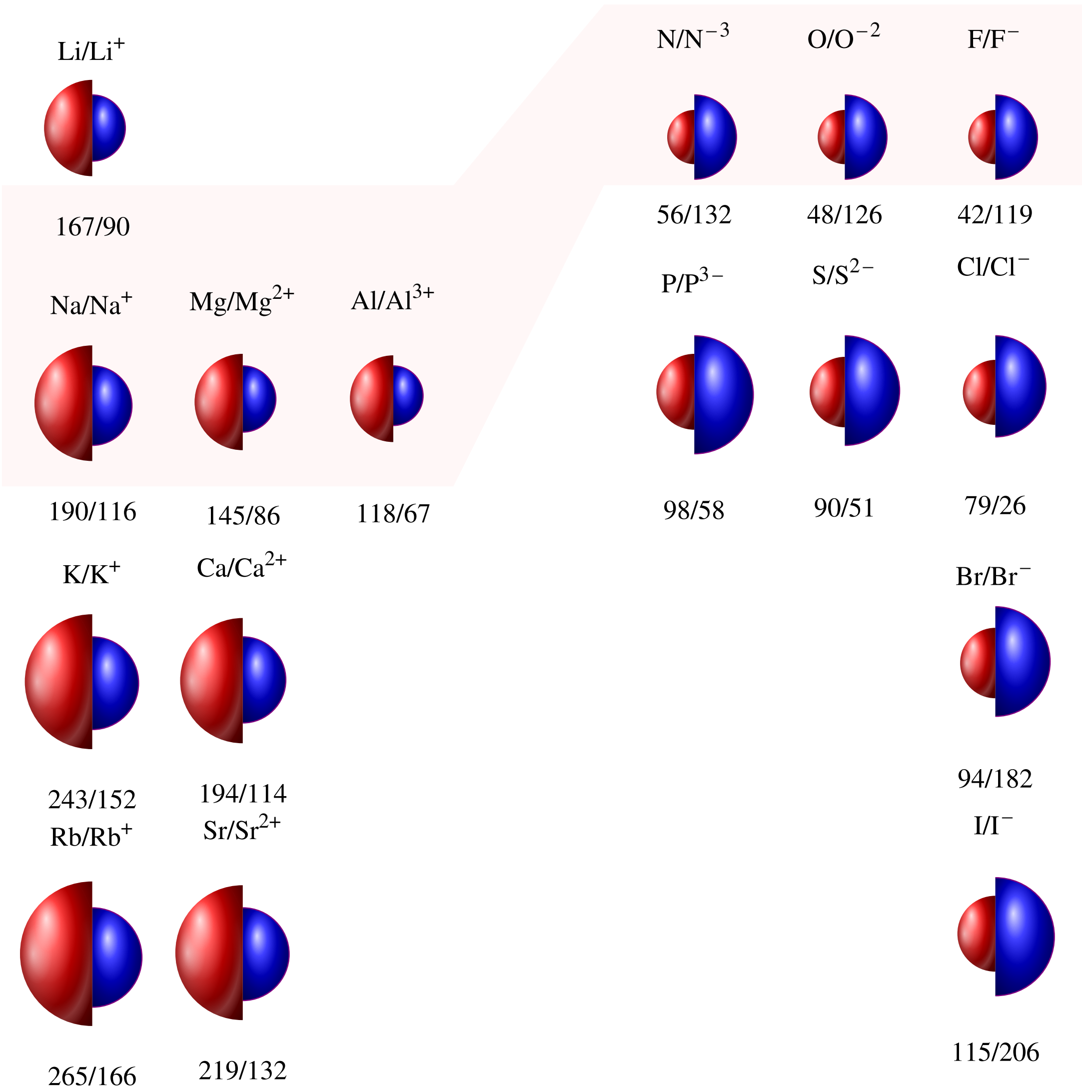


Figure 3.19 Ionic and atomic radius in pm across the periodic table

`Effective charge` The nuclear charge (Z) of an atom is the number of protons in the nucleus. In atoms with more than one electron, each electron is exposed to a nuclear effective charge (Z_{eff}) as the electrons partially compensate for the nuclear charge creating an effective nuclear charge. This way, the effective nuclear charge is the actual magnitude of the positive charge experienced by an electron. In atoms with many electrons, an electron is partially shielded from the positive charge of the nucleus. Among the different types of electrons in an atom (valence and core), core electrons are more effective at shielding, and as such Z_{eff} increases from left to right across the periodic table. As the number of core electrons remains the same and the atomic number increases from left to right across the table, Z_{eff} follows this trend too. The effective nuclear charge is given by

$$Z_{eff} = Z - \sigma$$

where σ is the shielding constant.

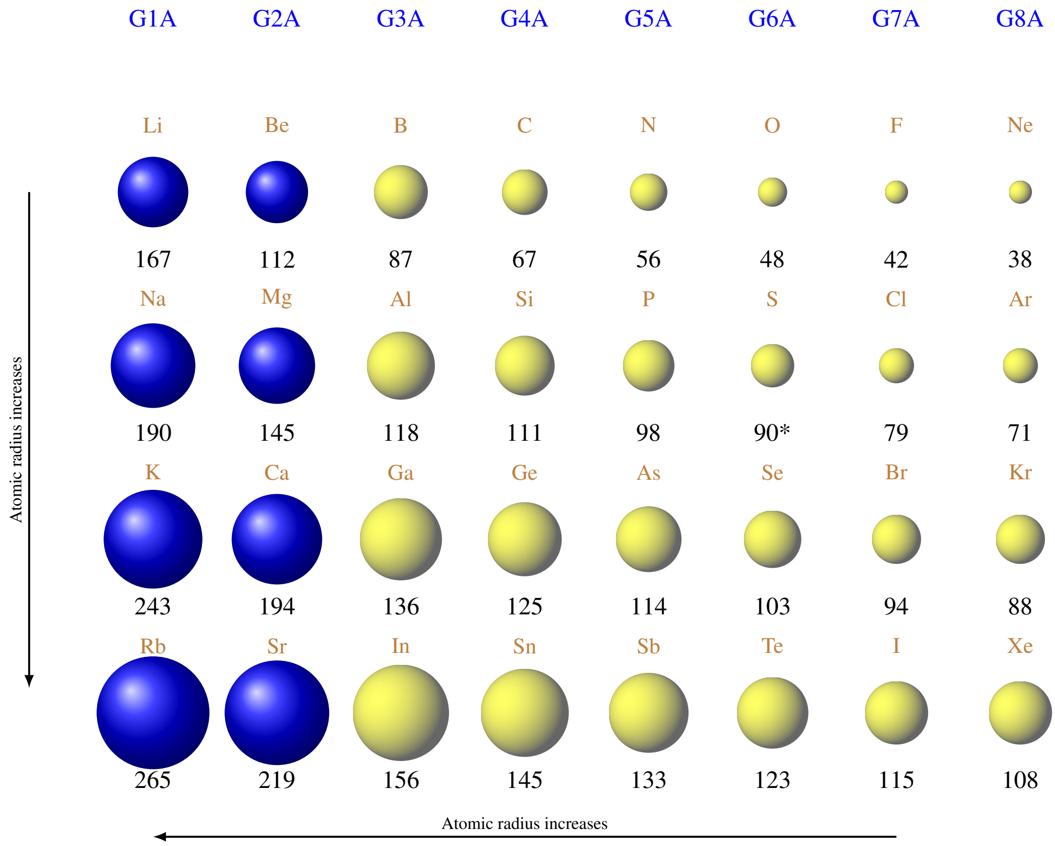


Figure 3.20 Change of atomic radius (calculated values given in pm) across the periodic table. The value with * was estimated

`Atomic radii` The size of atoms can not be specified exactly as the electron density of atoms stands throughout space. Still, we can arbitrarily define the atomic radii of atoms which are obtained by measuring distances between atoms in compounds. For covalent molecules, for example, O_2, we can measure the distance between the atoms in the molecule which is close to 150 pm, hence the atomic radius of oxygen–called the covalent atomic radius–should be close to 75pm. For metallic elements, we can measure the metallic radii using the lattice spacing between atoms. In the periodic table, the atomic radii decrease from left to right of a period and increased from top to bottom of a group (see Figure 3.19 and 3.20)

Table 3.3 Ionization energy in the periodic table in kJ/mol

Element	Z	EI_1	EI_2	EI_3	EI_4	EI_5	EI_6	EI_7	EI_8	EI_9	EI_10
Li	3	520.2	7298.1	11815.0							
Be	4	899.5	1757.1	14848.7	21006.6				*		
B	5	800.6	2427.1	3659.7	25025.8	32826.7					
C	6	1086.5	2352.6	4620.5	6222.7	37831	47277.0				
N	7	1402.3	2856	4578.1	7475.0	9444.9	53266.6	64360			
O	8	1313.9	3388.3	5300.5	7469.2	10989.5	13326.5	71330	84078.0		
F	9	1681.0	3374.2	6050.4	8407.7	11022.7	15164.1	17868	92038.1	106434.3	
Ne	10	2080.7	3952.3	6122	9371	12177	15238	19999.0	23069.5	115379.5	131432

* Shaded cells represent the removal of core electrons

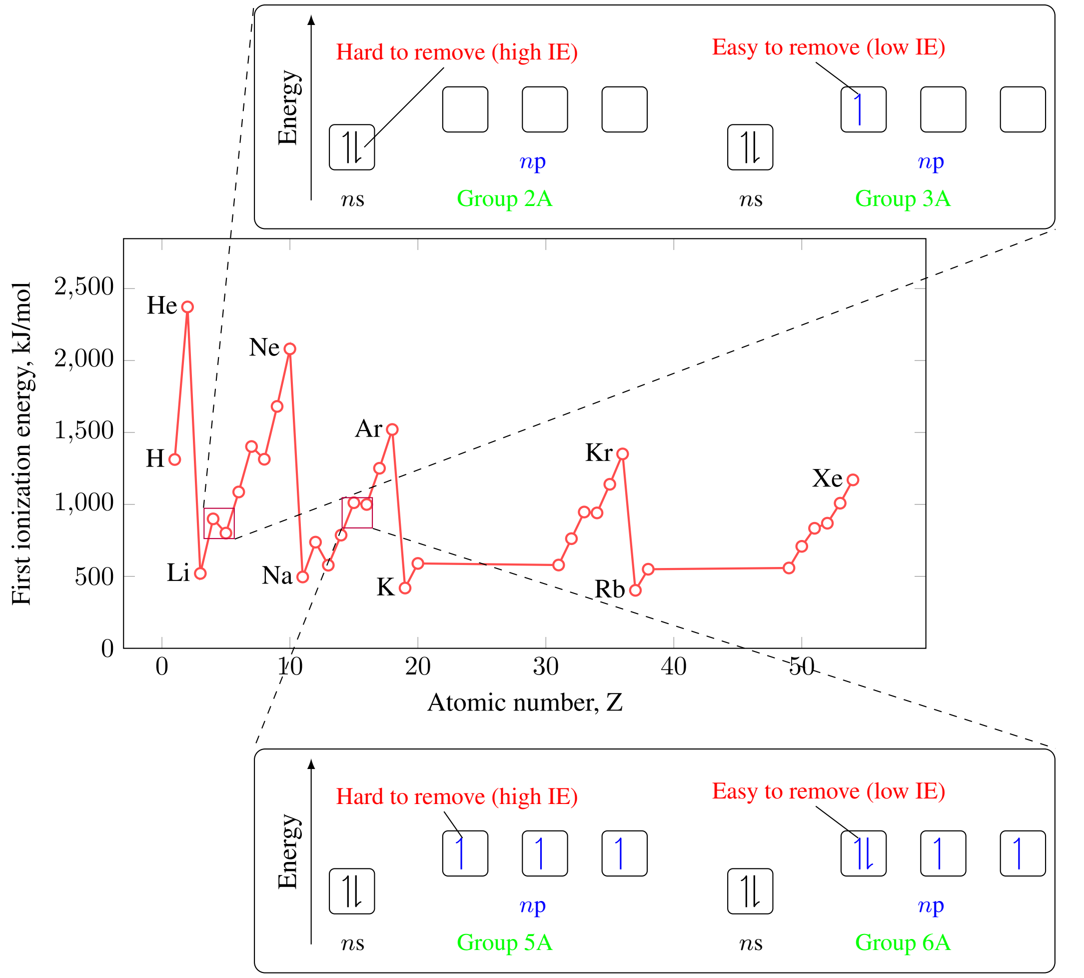


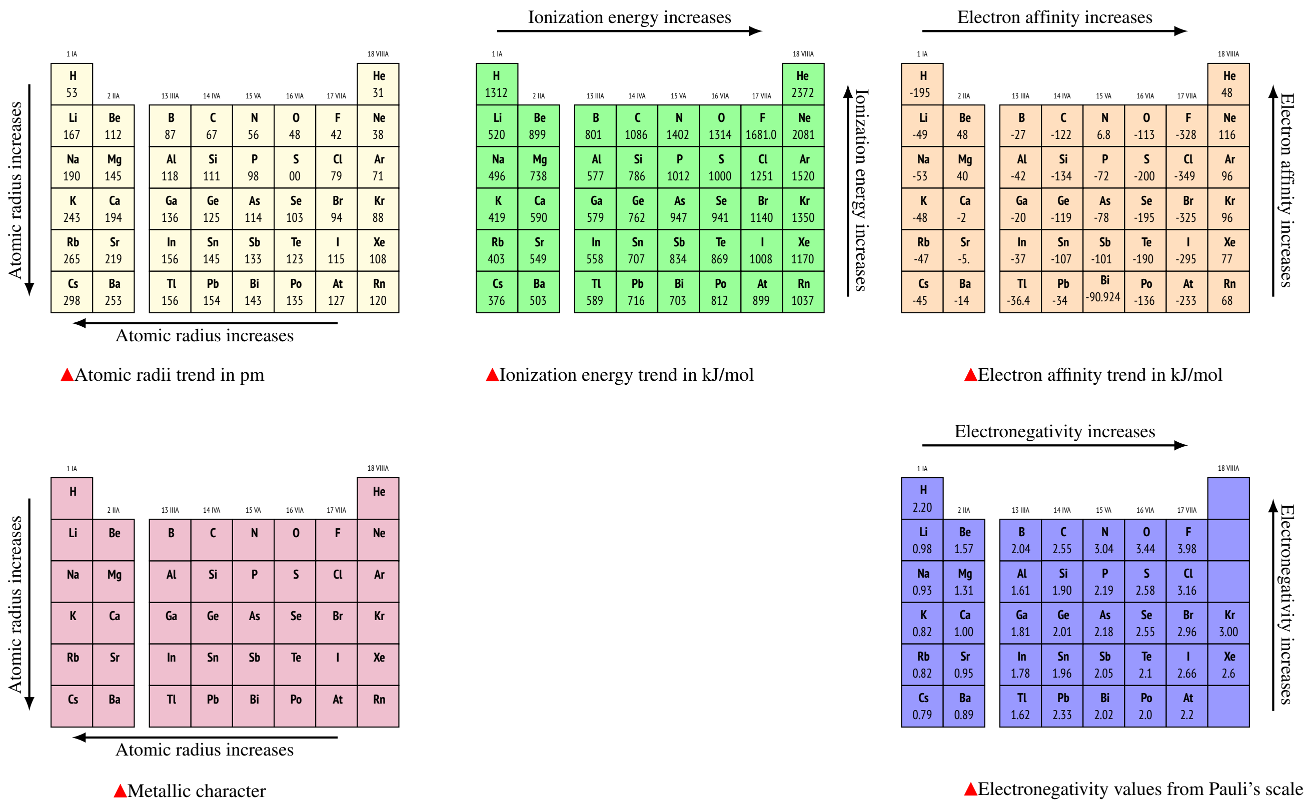
Figure 3.21 Ionization energy graphed with respect to the atomic number

`Ionization Energy` The *ionization energy* (IE) is the energy needed to remove an electron from an atom in a gas state and its fundamental (ground) electronic state. Atoms can have numerous ionization energies (see Table 3.3). The first ionization energy (IE_1) is the energy needed to remove the highest-energy electron on an atom, whereas the second ionization energy (IE_2) is the energy needed to remove the second highest-energy electron on the atom. For the case of boron,

$$B_{(g)} \longrightarrow B^+_{(g)} + e^- \qquad IE_1 = 800.6 kJ/mol$$

$$
\begin{aligned}
B^{+}_{(g)} &\longrightarrow B^{2+}_{(g)} + e^{-} \qquad IE_2 = 2427.1kJ/mol \\
B^{2+}_{(g)} &\longrightarrow B^{3+}_{(g)} + e^{-} \qquad IE_3 = 3659.7kJ/mol \\
B^{3+}_{(g)} &\longrightarrow B^{4+}_{(g)} + e^{-} \qquad IE_4 = 25025.8kJ/mol
\end{aligned}
$$

Given the electron configuration of Boron ([He]$2s^2 2p^1$), the first electron comes from a p orbital, whereas the second comes from a s orbital, which lays closer to the nucleus. The fourth electron comes from the core and the fourth ionization energy is considerably higher than the rest. Overall, we have that ionizing an atom from p orbitals is less costly than ionizing from s orbitals, and ionizing an atom from core orbitals is more costly than ionizing from valence orbitals. The ionization energy increases (this means more energy is needed to remove an electron) going up a group and when going across a period from left to right. In general, the ionization energy is low for metals and high for nonmetals. For example, if we compare the IE for H and He, as He is further right in the same period, it will have a larger IE and more energy will be needed to remove an electron ($IE_H < IE_{He}$). For the case of Li and H, the EI will be higher for H as it is further up in the same group ($IE_H > IE_{Li}$). There are discontinuities in the IE when moving across a period. Table **??** gives the IE for the second period. The first discontinuity happens between groups 2A and 3A as it is easier to remove p electrons than s electrons. The second discontinuity occurs between groups 5A and 6A as it is easier to remove an electron from a double occupies p orbital than from a single-occupied one.



▲Atomic radii trend in pm

▲Ionization energy trend in kJ/mol

▲Electron affinity trend in kJ/mol

▲Metallic character

▲Electronegativity values from Pauli's scale

Figure 3.22 Periodic properties across the periodic table

`Electron affinity` The electron affinity (EA) represents the tendency of an atom to receive one or more electrons hence becoming negatively charged. Specifically, is the energy involved in the process of receiving an electron. As an example, the electron affinity of hydrogen is written as:

$$H_{(g)} + e^- \longrightarrow H^-_{(g)} \qquad EA_1 = -195kJ/mol$$

Here, we consider electron affinities as negative values so that the smaller (the more negative) this value the larger the tendency of the atom to accept electrons. In other words, it would be easier to add electrons to an atom with -10kJ/mol EA, than to an atom with 10kJ/mol EA. In the periodic table, EA decreases (becomes more negative) across a period from left to right. For example, EA for K is -48kJ/mol and for Br is –325kJ/mol. Hence, elements from the right of the table are more willing to gain electrons than elements from the left. On one hand, there are discontinuities in the electron affinity across a period. It is easier to add an electron to the elements in Group 1A than to elements in Group 2A, as an electron added to Group 2 elements would be placed in p orbitals and these have higher–more positive–energies. Also, it is easier to add an electron to the elements in Group 4A than to elements in Group 5A, as an electron added to Group 4 elements would be placed in p orbitals giving three unpaired electrons, whereas electrons on Group 5 are all paired and a new electron would have to add to an already occupied orbital. On the other hand, there are fewer EA variations when going across a group. Finally, when adding more than one electron, the second EA tends to be more positive–less favored–than the first electron addition.

Electronegativity, EN Atoms combine by sharing or giving away electron density. The *electronegativity* of an atom in a molecule is its ability to attract electron density in a bond. Imagine two people holding a pillow with their hands, with the pillow representing a pair of electrons. If one of the people is more electronegative it will pull the pillow closer to them–it will attract more electrons. Electronegativity is loosely related to other periodic properties such as ionization energy and electron affinity. Very electronegative atoms such as F has also a large (negative) electron affinity (easily accepts electrons) and ionization energy (do not ionize easily). There are different electronegativity scales, and the most well-known scale was developed by Linus Pauling. Electronegativity (EN) increases (the atom is more electronegative) when going up a group (column) of the table and when going across a period (row), from left to right of the table. Highly electronegative atoms are located on the top right part of the table, whereas highly electropositive atoms are located on the bottom left part of the table. EN is a relative concept, that is, the electronegativity of an atom can only be measured about another atom. When comparing the EN of I and F, we found that F is more electronegative as it is further up in the same group ($EN_F > EN_I$).

Metallic character All elements in the periodic table have somehow a certain *metallic character* as they all can lose electrons as metals do–that is why metals are good conductors. The elements in the left part of the table are metals with a strong metallic character. Still, the elements on the right side of the table also have a certain metallic character (MC). The metallic character increases (the atom is more metallic) when going down a group (column) of the table and when going across a period (row) from right to left. For example, comparing K and Ca, we have that K is more metallic than Ca, as it is located further to the left in the same column ($MC_K > MC_{Ca}$). In the case of F and Cl, Cl is more metallic as it is located further down a group ($MC_{Cl} > MC_F$).

Sample Problem 28

Compare the following property for the given couple of elements: (a) Atomic radius of N and F. (b) Ionization energy of Na and Cs. (c) Electronegativity of F and I. (d) Metallic character of N and F.

SOLUTION

(a) You go from N to F by moving from left to right on a row of the periodic table, hence the radius of N is larger than the radius of F. This is because the atomic radius decreases from left to right on a row. (b) You go from Na to Cs by moving from up to down on a column of the periodic table, hence the ionization energy of Na is larger than for Cs. This is because the ionization energy decreases from top to bottom on a period. (c) You go from F to I by moving from down to up on a column of the periodic table, hence the electronegativity of F is larger than for I. This is because the electronegativity energy increased from bottom to top on a period. (d) You go from N to F by moving from left to right on a row of the periodic table, hence the metallic character of N is larger than for F. This is because the metallic character increased from right to left on a period.

STUDY CHECK

Compare the following property for the given couple of elements: (a) Atomic radius of F and I. (b) Electronegativity of Cs and Na.

Answer: (a) radius I $>$ F; (b) Electronegativity Na $>$ Cs.

`Ionic radius` Atoms gaining or losing electrons become ions. When an atom loses an electron it reduces its size. In particular, when an atom loses all valence electrons its size reduces considerably. When an atom gains an electron it gains size too as this new electron needs to be placed on an occupied orbital further away from the nucleus. Some ions have a common electronic configuration. For example, Na^+ and F^- both have the electron configuration: $[He]2s^22p^6$. Ions with the same electron configuration are called isoelectronic. For a series of isoelectronic ions, the larger the atomic number the smaller the ionic size due to the increase in nuclear attraction (see Figure **??**).

CHAPTER 3

The periodic table

3.1 Select from below the atomic symbol for the element Gold is: (a) Go (b) Au (c) G (d) Ca (e) Ol

3.2 Select from below the atomic symbol for the element Calcium is: (a) Go (b) Au (c) G (d) Ca (e) Ol

3.3 The atomic symbol for aluminum is: (a) Al (b) Am (c) A (d) Sn (e) Ag

3.4 The atomic symbol for iron is: (a) Ir (b) Fs (c) Fe (d) In (e) Ir

3.5 Ca is the symbol for: (a) Carbon (b) Calcium (c) Cobalt (d) Copper (e) Cadmium

3.6 Write down the symbol for the following elements: (a) Magnesium (b) Manganese (c) Mercury (d) Molybdenum (e) Neodymium (f) Neon (g) Neptunium (h) Nickel (i) Osmium (j) Palladium

3.7 Write down the symbol for the following elements: (a) Phosphorus (b) Platinum (c) Plutonium (d) Polonium (e) Potassium (f) Radium (g) Radon (h) Rhenium (i) Rhodium (j) Rubidium (k) Ruthenium

3.8 Write down the element names for the following chemical symbols: (a) Sc (b) Se (c) Si (d) Ag (e) Na (f) Sr (g) S (h) Te

3.9 Write down the element names for the following chemical symbols: (a) Tl (b) Th (c) Sn (d) Ti (e) W (f) U (g) V (h) Xe (i) Zn (j) Zr

3.10 Which of the following elements is a metal? (a) Nitrogen (b) Lithium (c) Calcium (d) Iron (e) Iodine

3.11 Which of the following elements is a nonmetal? (a) Nitrogen (b) Lithium (c) Calcium (d) Iron (e) Iodine

3.12 Which of the following elements is a alkali metal? (a) Nitrogen (b) Lithium (c) Calcium (d) Iron (e) Ruthenium

3.13 Which of the following elements is a alkaline earth metal? (a) Nitrogen (b) Lithium (c) Calcium (d) Iron (e) Ruthenium

3.14 Which of the following elements is a halogen? (a) Nitrogen (b) Lithium (c) Calcium (d) Iron (e) Iodine

3.15 Which of the following elements is a chalcogen? (a) Oxygen (b) Lithium (c) Calcium (d) Iron (e) Iodine

3.16 What is the symbol of the element in Period 4 and Group 2? (a) Be (b) Mg (c) Ca (d) C (e) Si

3.17 What is the symbol of the element in Period 2 and Group 2? (a) Be (b) Mg (c) Ca (d) C (e) Si

3.18 Identify the group number described by: (a) Starts with Zn (b) Ends with Ra (c) Contains the elements O, S, and Se (d) Is located to the right of Mn (e) Is located to the left of Ti

3.19 Identify the group number described by: (a) Starts with B (b) Ends with Rn (c) Contains the elements Mn, Tc, and Re (d) Is located to the right of H (e) Is located to the left of F

3.20 Identify the elements located at: (a) group 7A and period 3 (b) An halogen at period 5 (c) An alkaline earth at period 7 (d) An alkali at period 3

3.21 Identify the elements located at: (a) group 2A and period 4 (b) An halogen at period 3 (c) An alkaline earth at period 5 (d) An alkali at period 6

3.22 Identify the elements below as metal, nonmetal or metalloid: (a) B (b) Sb (c) Fe (d) Zn (e) H

3.23 Identify the elements below as metal, nonmetal or metalloid: (a) A shinny element (b) C (c) A electrically conducting element (d) A dull element (e) A semiconductor

Early experiments of the atom

3.24 From the following scientist, J.J. Thomson, Robert Millikan, Henri Becquerel, and Ernest Rutherford, indicate who: (a) Discovered the element Uranium in a mineral (b) Scattered atoms on helium on a gold thin layer (c) Unsuccesfully validated the plum pudding model

3.25 From the following scientist, J.J. Thomson, Robert Millikan, Henri Becquerel, and Ernest Rutherford, indicate who: (a) Worked with cathodic rays (b) Calculate the charge-to-mass ratio of the electron(c) Worked with oil drops to calculate the electric charge of the electron

3.26 Based on Rutherford's experiment, answer the following: (a) What did Rutherford expected after aiming particles to the gold foil? (b) What did Rutherford found after aiming particles to the gold foil?

3.27 Rutherford's scattering experiment was based on: (a) The scattering of alpha particles from gold foil (b) The scattering of beta particles from gold foil (c) The scattering of alpha particles from zinc foil (d) The scattering of gamma particles from iron foil

The Atom

3.28 In an atom, the nucleus contains: (a) an equal number of protons and electrons. (b) all the protons and neutrons (c) all the protons and electrons (d) only neutrons (e) only protons

3.29 In an ion, the nucleus contains: (a) an equal number of protons and electrons. (b) all the protons and neutrons (c) all the protons and electrons (d) only neutrons (e) only protons

3.30 Associate the following statements with either an electron, a proton or a neutron: (a) are found away from the nucleus (b) are atracted to a proton (c) are negatively charged (d) are found in the nucleus

3.31 Associate the following statements with either an electron, a proton or a neutron: (a) are atracted to a electron (b) have the smallest mass (c) are positively charged (d) are neutraly charged

3.32 Indicate whether the following statements are true or false: (a) Neutrons repel to each other (b) Protons and neutrons have opposite charges (c) Electrons repel to each other

3.33 Indicate whether the following statements are true or false: (a) Protons and electrons have opposite charges (b) Protons repel to each other (c) Neutrons and protons have very different mass

3.34 The atomic number of an atom is equal to the number of: (a) nuclei (b) neutrons (c) neutrons plus protons (d) electrons plus protons (e) electrons

3.35 The mass number of an atom is equal to the number of: (a) nuclei (b) neutrons (c) neutrons plus protons (d) electrons plus protons (e) electrons

3.36 Consider a neutral atom with 30 protons and 34 neutrons. The atomic number of the element is: (a) 30 (b) 32 (c) 34 (d) 64 (e) 94

3.37 Consider a neutral atom with 40 protons and 45 neutrons. The mass number of the element is: (a) 40 (b) 45 (c) 80 (d) 85 (e) 94

3.38 Identify the name of the following isotopes (a) ${}^{106}_{44}X$ (b) ${}^{93}_{40}X$ (c) ${}^{85}_{36}X$

3.39 Identify the name of the following isotopes (a) ${}^{155}_{63}X$ (b) ${}^{126}_{50}X$ (c) ${}^{107}_{46}X$

3.40 Select the isotopes from the list below: (a) ${}^{90}_{228}X$ (b) ${}^{231}_{90}X$ (c) ${}^{230}_{90}X$ (d) ${}^{90}_{229}X$ (e) ${}^{228}_{90}X$ (f) ${}^{230}_{100}X$ (g) ${}^{229}_{90}X$

3.41 Select the isotopes from the list below: (a) ${}^{233}_{92}X$ (b) ${}^{92}_{234}X$ (c) ${}^{234}_{92}X$ (d) ${}^{232}_{92}X$ (e) ${}^{92}_{233}X$ (f) ${}^{100}_{40}X$

3.42 Fill the table below indicating the number of electrons (n_e), neutrons (n_n) and protons (n_p) for the following neutral atoms:

	Isotope	n_e	n_p	n_n
a	$^{14}_{7}N$			
b	$^{15}_{7}N$			
c	$^{13}_{7}N$			

3.43 Fill the table below indicating the number of electrons (n_e), neutrons (n_n) and protons (n_p) for the following neutral atoms:

	Isotope	n_e	n_p	n_n
a	$^{134}_{55}Cs$			
b	$^{135}_{55}Cs$			
c	$^{137}_{55}Cs$			

3.44 Fill the table below indicating the number of electrons (n_e), neutrons (n_n) and protons (n_p) for the following charged atoms:

	Isotope	n_e	n_p	n_n
a	$^{48}_{20}Ca^{+2}$			
b	$^{18}_{9}F^{-}$			
c	$^{62}_{28}Ni^{3+}$			

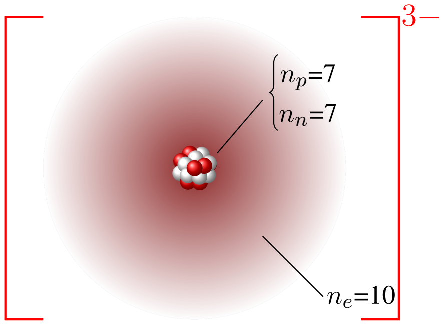
3.45 Fill the table below indicating the number of electrons (n_e), neutrons (n_n) and protons (n_p) for the following charged atoms:

	Isotope	n_e	n_p	n_n
a	$^{147}_{61}Pm^{+}$			
b	$^{192}_{77}Ir^{-}$			
c	$^{209}_{83}Bi^{-2}$			

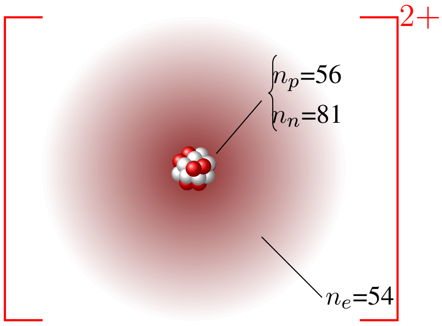
3.46 Calculate the number of protons, neutrons and electrons in the following isotopes: (a) $^{15}_{7}N$ (b) $^{64}_{29}Cu$ (c) $^{93}_{40}Zr$

3.47 Calculate the number of protons, neutrons and electrons in the following isotopes: (a) $^{79}_{34}Se$ (b) $^{85}_{36}Kr$ (c) $^{13}_{7}N$

3.48 Write down the isotopic symbol for the atom or ion represented below indicating the name of the element.



3.49 Write down the isotopic symbol for the atom or ion represented below indicating the name of the element.



3.50 Write down the isotopic symbol given the number of protons, neutrons and electrons: (a) 43 protons, 43 electrons and 60 neutrons (b) 48 protons, 48 electrons and 65 neutrons (c) 77 protons, 77 electrons and 115 neutrons

3.51 Write down the isotopic symbol given the number of protons, neutrons and electrons: (a) 93 protons, 93 electrons and 142 neutrons (b) 55 protons, 55 electrons and 79 neutrons (c) 52 protons, 52 electrons and 72 neutrons

3.52 The atomic mass of Ga is 69.72 amu. There are only two naturally occurring isotopes of gallium: 69Ga, with a mass of 69.0 amu, and 71Ga, with a mass of 71.0 amu. Calculate the natural abundance of the 69Ga isotope.

3.53 The atomic mass of Xe is 131.293 amu. There are only two naturally occurring isotopes of Xe: Xenon-133, with a mass of 133.0 amu, and Xenon-135, with a mass of 135 amu. Calculate the natural abundance of the Xenon-135 isotope.

3.54 Magnesium contains three different isotopes: magnesium-24 with an abundance of 79% and a mass of 23.9850423 amu, magnesium-25 with an abundance of 10% and a mass of 24.9858374 amu, and magnesium-26 with a mass of 25.9825937 amu. Calculate the abundance of magnesium-26 and the average atomic mass of a sample of magnesium.

3.55 Silicon contains three different isotopes: Si-28 with a mass 27.976927amu and abundace of 92.2297%, Si-29 with a mass 28.976495amu and abundace of 4.6832% and Si-30 with a mass 29.973770amu. Calculate the abundance of Si-30 and the average atomic mass of a sample of Si.

3.56 There are only two naturally occurring isotopes of Lutetium: Lu-175 with a mass of 174.94 amu and abundace of 97.41% , and Lu-176 with a mass of 175.94 amu and abundace of 2.59%. Calculate the average mass of the element using the table below, givem m_i is the isotopic mass and f_i is the fractional abundance of the isotope.

Isotope	m_i, amu	f_i	$m_i \times f_i$
Average mass = amu			$\sum m_i \times f_i =$

3.57 There are only two naturally occurring isotopes of Thallium: Tl-203 with mass of 202.972 amu and abundace of 29.52%, and Tl-205 with mass of 204.974 amu and abundance of 70.48%. Calculate the average mass of the element using the table below, givem m_i is the isotopic mass and f_i is the fractional abundance of the isotope.

Isotope	m_i (amu)	f_i	$m_i \times f_i$
Average atomic mass (amu)			$= \sum m_i \times f_i =$

3.58 Fill the table below and calculate the missing variable.

Isotope	m_i (amu)	f_i	$m_i \times f_i$
^{107}Ag	106.9050	0.5184	?
^{109}Ag	108.9047	0.4816	?
Average atomic mass (amu)			$= \sum m_i \times f_i =?$

3.59 Fill the table below and calculate the missing variable.

Isotope	m_i (amu)	f_i	$m_i \times f_i$
^{137}Ba	136.9058	?	?
^{138}Ba	137.9052	?	?
Average atomic mass (amu)		$= \sum m_i \times f_i =114.38$	

Electronic structure

3.60 Describe the shape of the following orbitals using the terms lobular, spherical, or lobular(dumbbell shape): (a) 3d (b) 1s (c) 2p (d) 2s

3.61 Describe the shape of the following orbitals using the terms lobular, spherical, or lobular(dumbbell shape): (a) 4p (b) 4d (c) 3p (d) 3s

3.62 Match the following statements with: (1) have the same shape (2) can fit the same number of electrons (3) belong to the same energy level (4) can fit the same number of electrons (a) 1s and 2s orbitals (b) five 3d orbitals (c) 2p and 3p sublevels (d) 3s and 3p sublevels

3.63 Match the following statements with: (1) have the same shape (2) can fit the same number of electrons (3) belong to the same energy level (4) can fit the same number of electrons (a) 4p and 3p orbitals (b) three 3p orbitals (c) 3d and 3s sublevels (d) 3p and 5p sublevels

3.64 Indicate the following: (a) orbitals in a 1s sublevel (b) sublevels in the fourth energy level (c) orbitals in the fourth energy level

3.65 Indicate the following: (a) orbitals in a 3p sublevel (b) sublevels in the third energy level (c) orbitals in the third energy level

3.66 Indicate the following: (a) the maximum number of electrons in a 3d sublevel (b) the maximum number of electrons in a 4f sublevel (c) the maximum number of electrons in a 1s sublevel

3.67 Indicate the following: (a) the maximum number of electrons in a 1s orbital (b) the maximum number of electrons in a 3d orbital (c) the maximum number of electrons in a 4f orbital

3.68 Indicate the number of orbitals that can have the following designations: (a) 2s (b) 3p (c) 0p (d) $n = 4$

3.69 Indicate the number of orbitals that can have the following designations: (a) 1d (b) $n = 1$ (c) 3d (d) 4f

3.70 What is the element with the electron configuration (a) $1s^2 2s^2 2p^6 3s^1$ (b) $[Ar]3d^5 4s^1$ (c) $1s^2 2s^2 2p^6 3s^1$ (d) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^2 4s^2$

3.71 What is the element with the electron configuration (a) $1s^2 2s^2 2p^6 3s^2 3p^5$ (b) $1s^2 2s^2 2p^6 3s^2 3p^4$ (c) $[Kr]5s^2 4d^8$

3.72 Write down the abbreviated electron configuration for the following elements: (a) Cobalt (b) Chromium (c) Argon (d) Potassium (e) Chlorine (f) Vanadium

3.73 Write down the abbreviated electron configuration for the following elements: (a) Calcium (b) Manganese (c) Iron (d) Nickel (e) Copper (f) Titanium

3.74 Write down the full electron configuration for the following elements: (a) Barium (b) Samarium (c) Vanadium

3.75 Write down the full electron configuration for the following elements: (a) Aluminum (b) Potassium (c) Nickel

3.76 Write down the orbital diagram for the following elements: (a) Argon (b) Silicon (c) Sulfur (d) Chlorine

3.77 Write down the orbital diagram for the following elements: (a) Phosphorus (b) Boron (c) Carbon (d) Nitrogen

3.78 Indicate whether the following orbital diagram is possible for an atom on its fundamental state:

$$[He]\ \underbrace{\boxed{\uparrow\downarrow}}_{2s}\ \underbrace{\boxed{\uparrow}\ \boxed{\uparrow}\ \boxed{\downarrow}}_{2p}$$

3.79 Indicate whether the following orbital diagram is possible for an atom on its fundamental state:

$$[He]\ \underbrace{\boxed{\uparrow}}_{2s}\ \underbrace{\boxed{\uparrow\downarrow}\ \boxed{\uparrow}\ \boxed{\uparrow}}_{2p}$$

3.80 Indicate whether the following orbital diagram is possible for an atom on its fundamental state:

$$[He]\ \underbrace{\boxed{\uparrow\downarrow}}_{2s}\ \underbrace{\boxed{\uparrow\downarrow}\ \boxed{\downarrow}\ \boxed{}}_{2p}$$

3.81 Indicate whether the following orbital diagram is possible for an atom on its fundamental state:

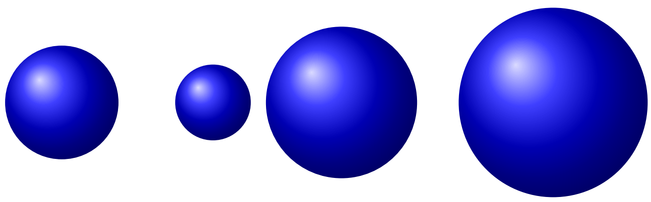
$$[He]\ \underbrace{\boxed{\uparrow}}_{2s}\ \underbrace{\boxed{\uparrow}\ \boxed{\uparrow}\ \boxed{}}_{2p}$$

Periodic Properties

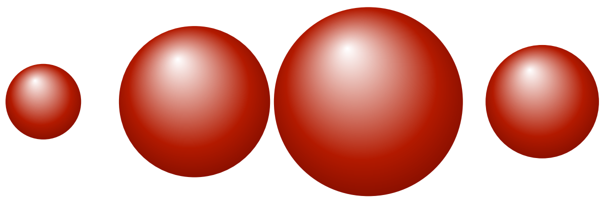
3.82 Among the elements, indicate the element with the largest atomic radius (a) B (b) C (c) F (d) Li (e) Na

3.83 Among the elements, indicate the element with the smallest atomic radius (a) C (b) N (c) O (d) S (e) Se

3.84 Match the spheres (far left, center left, center right, far right) with the atoms: Be, Li, Na, B



3.85 Match the spheres (far left, center left, center right, far right) with the atoms: P, S, Si ,Cl.



3.86 Among the elements, indicate the element with the largest electronegativity (a) B (b) C (c) F (d) Li

3.87 Among the elements, indicate the element with the largest electronegativity (a) Si (b) P (c) S (d) Se

3.88 Among the elements, indicate the element with the smallest ionization energy (a) B (b) C (c) F (d) Li (e) Na

3.89 Among the elements, indicate the element with the largest ionization energy (a) Al (b) Si (c) P (d) As (e) Sb

3.90 Among the elements, indicate the element with the largest metallic character (a) B (b) C (c) F (d) Li (e) Na

3.91 Among the elements, indicate the element with the largest metallic character (a) K (b) Rb (c) Cs (d) Ca

Ch. 4. Nuclear Chemistry

THE field of nuclear medicine was first established in 1934 with the production of artificial radioactive substances. This field uses the power of nuclear chemistry to cure cancer and other diseases or simply to visualize organs. In 1937, the first radioactive isotope was used to treat a person with leukemia at the University of California at Berkeley. Radioactive substances are now used to produce images of organs, such as the liver, spleen, thyroid gland, kidneys, and brain, and to detect heart disease. Today, procedures in nuclear medicine provide information about the function and structure of every organ in the body, which allows the nuclear physician to diagnose and treat diseases early. This chapter covers the basic principles of nuclear chemistry. You will learn the real meaning of radioactivity and how to quantify the effects of radiation or measure the duration of a radioactive chemical.

> “Nuclear power is one hell of a way to boil water.”
> *Einstein*

4.1 Radiation, particles & radioisotopes

Light elements have normally stable nuclei. Differently, heavier elements with atomic numbers larger than 20 tend to often have several isotopes–remember these are atoms of the element with a different number of neutrons–that have unstable nuclei. For these unstable isotopes, the forces that keep the nucleus together are not strong enough to stabilize the nuclei. An unstable nucleus is radioactive, which means that it will spontaneously emit radiation in the form of small particles. Not all radioactivity is the same and there exist different types of radiation, which we will address in the following. Table 4.1 reports common nuclear symbols.

`alpha radiation` Alpha radiation–referred to as α–is a type of radiation that contains alpha particles. These particles are indeed helium nuclei, with 2 protons, 2 neutrons, and a (2+) positive charge. Alpha particles are often represented as α or ${}^{4}_{2}\mathrm{He}$.

`beta radiation` Beta radiation–referred to as β–is a type of radiation that contains beta particles. These particles are indeed high-energy electrons with ($-$) negative charge. Beta particles are often represented as β or ${}^{0}_{-1}\mathrm{e}$.

`gamma radiation` Gamma radiation–referred to as γ–is a type of radiation that contains high-energy photons. These particles are indeed photons with no mass or charge. Gamma particles are often represented as γ or ${}^{0}_{0}\gamma$.

`protons` Protons in this chapter are often referred to as p or ${}^{1}_{1}\mathrm{H}^{+}$. These are positive charges.

`positrons` Positrons are the electron antiparticle, often referred to as β^{+} or ${}^{0}_{+1}\mathrm{e}$. They do have a positive charge.

▼Strawberries are normally treated with radiation

`neutrons` Neutrons are nuclear particles with no charge, often referred to as n or ${}^{1}_{0}\mathrm{n}$.

Sample Problem 29

Name or give the symbols for the following nuclear particles: beta particle, β^{+}, p and ${}^{0}_{0}\gamma$.

SOLUTION

Beta particles are represented by β or ${}^{0}_{-1}\mathrm{e}$. These particles are indeed simply electrons ejected during a nuclear decay. β^{+} represents a positron, an anti-electron. p stands for protons, a nuclear particle with positive charge. Finally, ${}^{0}_{0}\gamma$ represents gamma radiation.

STUDY CHECK

Name or give the symbols for the following nuclear particles: n, α and ${}^{1}_{1}\mathrm{H}^{+}$.

▼Ionization chamber smoke detectors contain a small amount of americium-241, a radioactive material

`Radioisotope notation` Radioisotopes–atomic isotopes that produce radiation–are written as ${}^{A}_{Z}\mathrm{X}$. For example, ${}^{14}_{6}\mathrm{C}$ is referred to as carbon-14. The number on the top is the mass number A, that is represented the total number of neutrons and protons in the isotope. The number on the bottom refers to the atomic number Z, that is, the total number of electrons in the atom. For example, the mass number of ${}^{14}_{6}\mathrm{C}$ is 14 whereas its atomic number is 6. ${}^{14}_{6}\mathrm{C}$ has 14 neutrons and protons and 6 electrons.

Sample Problem 30

Calculate the number of protons, neutrons and electrons of the following isotopes: ${}^{238}_{92}\mathrm{U}$, ${}^{24}_{13}\mathrm{Al}$ and ${}^{14}_{6}\mathrm{C}$.

SOLUTION

According to the isotope notation (${}^{A}_{Z}\mathrm{X}$), the number of top of the radioisotope is the mass number A that represents the number of protons plus neutrons, whereas the number of the bottom is the atomic number Z that represents the number of electrons. According to this, the number of electrons in an atom is Z. If an atom is neutral, the number of electrons and protons are the same, so the number of protons is also Z. The number of neutrons would hence be $A - Z$, as A is the number of protons+neutrons, and the number of protons is Z. We'll use a table below to obtain the electrons, protons and neutrons from A and Z.

▼An operating nuclear power plant produces very small amounts of radioactive gases

Radioisotope	A	Z	Electrons	protons	neutrons
${}^{238}_{92}\mathrm{U}$	238	92	92	92	146
${}^{24}_{13}\mathrm{Al}$	24	13	13	13	11
${}^{14}_{6}\mathrm{C}$	14	6	6	6	8

STUDY CHECK

Calculate the number of protons, neutrons and electrons of ${}^{99}_{43}\mathrm{Tc}$.

▼There are radioactive gases in the air we breathe

Table 4.1 Nuclear symbols

Particle Name		Symbol	Charge	Identity	Penetrating power	Discovery
Alpha	(α)	$^4_2$He	2+	Helium nucleus	Minimal	1899
Beta	(β)	$^{\ 0}_{-1}$e	-1	Electrons	Short	1899
Gamma	(γ)	$^0_0\gamma$	0	Electromagnetic radiation	Deep	1900
Neutrons	(n)	$^1_0$n	0	nuclear particle	Maximal	1932
Proton	(p)	$^1_1$H$^+$	$+1$	nuclear particle		1919
Positrons	(β^+)	$^{\ 0}_{+1}$e	$+1$	antiparticle		1932

4.2 Nuclear reactions

Isotopes–called emitters–spontaneously decompose producing new isotopes in a process called radioactive decay. In this decay, radiation is also emitted.

$$\text{Emitter} \longrightarrow \text{radiation} + \text{New isotope}$$

In the following, we will discuss the most important type of radioactive decay.

`alpha decay` Some isotopes produce alpha radiation, that is, they produce α particles on its decay. A nuclear reaction that produces an α particle ($^4_2$He) is called alpha decay. In alpha decay, the emitter decreases its mass number A four units and its atomic number Z two units.

$$\text{Emitter} \longrightarrow {}^4_2\text{He} + \text{New isotope}$$

`beta decay` Other isotopes produce beta radiation, that is, they produce β particles on its decay. A nuclear reaction that produces a β particle ($^{\ 0}_{-1}$e) is called beta decay. In beta decay, the emitter has the same mass number A as the product isotope. However, its atomic number Z decreases by one unit.

$$\text{Emitter} \longrightarrow {}^{\ 0}_{-1}\text{e} + \text{New isotope}$$

`positron emission` Certain isotopes decay by producing a positron, that is, they produce $^{\ 0}_{+1}$e particles on its decay. A nuclear reaction that produces $^{\ 0}_{+1}$e is called positron emission. In a positron emission, the emitter has the same mass number A as the product isotope. However, its atomic number Z increases by one unit.

$$\text{Emitter} \longrightarrow {}^{\ 0}_{+1}\text{e} + \text{New isotope}$$

`gamma decay` Some other isotopes produce gamma radiation in the form of γ particles on its decay. A nuclear reaction that produces a γ particle ($^0_0\gamma$) is called gamma decay. In this type of decay, no new isotope is produced. Gamma emitters are normally excited, that is they have higher energy than normal; we denote this with a * symbol. Exited particles tend to lose energy to become more stable. In gamma decay, the emitter and the product isotope, both have the same mass and atomic number.

$$\text{Emitter}^* \longrightarrow {}^0_0\gamma + \text{Emitter}^*$$

Sample Problem 31

Label the following nuclear reactions as: α, β or γ decay, or positron emission:

(a) $^{238}_{92}\text{U} \longrightarrow {}^{234}_{90}\text{Th} + {}^{4}_{2}\text{He}$

(b) $^{24}_{13}\text{Al} \longrightarrow {}^{24}_{12}\text{Mg} + {}^{0}_{+1}\text{e}$

(c) $^{14}_{6}\text{C} \longrightarrow {}^{14}_{7}\text{N} + {}^{0}_{-1}\text{e}$

(d) $^{99}_{43}\text{Tc}^{*} \longrightarrow {}^{99}_{43}\text{Tc} + {}^{0}_{0}\gamma$

SOLUTION

(a) This process produces $^{4}_{2}$He and therefore is alpha emission. (b) This process generates $^{0}_{+1}$e and therefore is positron emission. (c) This process produces $^{0}_{-1}$e and therefore is beta emission. (d) This process produces $^{0}_{0}\gamma$ and therefore is gamma emission.

STUDY CHECK

Label the following nuclear reactions as: α, β or γ decay, or positron emission:

(a) $^{127}_{55}\text{Cs} \longrightarrow {}^{127}_{54}\text{Xe} + {}^{0}_{+1}\text{e}$

(b) $^{27}_{13}\text{Al}^{*} \longrightarrow {}^{27}_{13}\text{Al} + {}^{0}_{0}\gamma$

(c) $^{218}_{85}\text{At} \longrightarrow {}^{214}_{83}\text{Bi} + {}^{4}_{2}\text{He}$

4.3 Unknown isotopes in nuclear reactions

Sometimes one needs to identify an unknown isotope $^{A}_{Z}$X in a chemical reaction. This means identifying the name X of the isotope, the atomic number Z as well as the mass number A. To do this, we will use the fact that the total mass number as well as the total atomic number should stay constant before and after the nuclear reaction. Let's break this idea down into an example.

Sample Problem 32

Identify the unknown isotope in the following nuclear reaction:

$$^{4}_{2}\text{He} + {}^{14}_{7}\text{N} \longrightarrow {}^{A}_{Z}\text{X} + {}^{1}_{1}\text{H}$$

SOLUTION

We will solve this problem by using the fact that the total mass number as well as the total atomic number should stay constant before and after the nuclear reaction. In order to do this, we will calculate the total atomic number before the reaction (in the left) and the total atomic number after the reaction (in the right) and equal both values. We will do the same for the mass number.

	$^{4}_{2}$**He**	$^{14}_{7}$**N**	$\longrightarrow$	$^{A}_{Z}$**X**	$^{1}_{1}$**H**
A	4	14		A	1
Z	2	7		Z	1

Now we build up two equations, one for A and another for Z, from each of the columns (column 2 and 3) of the table:

$$4 + 14 = A + 1 \qquad \text{Equation for A from column 2 of the table}$$

$$2 + 7 = Z + 1 \qquad \text{Equation for Z from column 3 of the table}$$

We now solve for A geting a value of $A = 17$ and for Z, getting $Z = 8$. With the atomic number Z we can go to the periodic table and identify the name of the isotope. The element with $Z = 8$ is called oxygen, and the final answer would be: $^{17}_{8}$O.

STUDY CHECK

Identify the unknown isotope in the following nuclear reaction:

$$^{241}_{95}Am \longrightarrow {}^{A}_{Z}X + {}^{4}_{2}He$$

Table 4.1 Half-life for various isotopes and chemicals

Americium-241	432.2 years	Lutetium-177	6.71 days	Hydrogen-3	12.35 years
Barium-133	10.74 years	Molybdenum-99	66 hours	Technetium-99	213,000 years
Bismuth-212	60.55 minutes	Nickel-63	96 years	Indium-111	2.83 days
Cadmium-109	464 days	Phosphorus-32	14.29 days	Technetium-99m	6.02 hours
Calcium-45	163 days	Phosphorus-33	25.4 days	Indium-113m	1.658 hours
Carbon-14	5730 years	Plutonium-239	24,065 years	Tin-113 115.1	days
Cesium-137	30 years	Polonium-210	138.38 days	Iodine-123	13.2 hours
Chlorine-36	301,000 years	Radium-226	1600 years	Tungsten-188	69.4 days
Chromium-51	27.704 days	Radon-222	3.8235 days	Iodine-125	60.14 days
Cobalt-57	270.9 days	Rhenium-188	16.98 hours	Uranium-235	703,800,000 years
Cobalt-58	70.8 days	Rubidium-81	4.58 hours	Iodine-129	15,700,000 years
Cobalt-60	5.271 years	Selenium-75	119.8 days	Uranium-238	4,468,000,000 years
Copper-62	9.74 minutes	Sodium-22	2.602 years	Iodine-131	8.04 days
Copper-64	12.701 hours	Sodium-24	15 hours	Xenon-127	6.41 days
Copper-67	61.86 hours	Strontium-85	64.84 days	Iron-55	2.7 years
Gallium-67	78.26 hours	Strontium-89	50.5 days	Xenon-133	5.245 days
Gold-195	183 days	Sulfur-35	87.44 days	Iron-59	44.529 days
Ondansetron	360 min	Capecitabine	2400s	Carmustine	0.25h

4.4 Half-life of a radioisotope

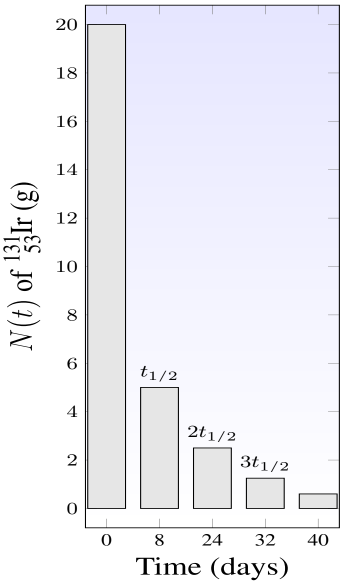
Radioisotopes–isotopes that decay producing radiation–are unstable and with time they eventually disappear given a more stable isotope. Some radioisotopes decay very quickly, such as the ones used in nuclear medicine to fight cancer. Other radioisotopes take longer to disappear.

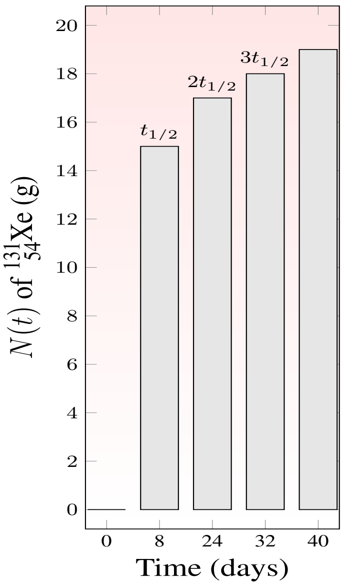
`The concept of half-live` The half-life of an isotope represented as $t_{1/2}$ is the time it takes for an isotope to disappear reducing the sample mass to half the initial value. For example, $t_{1/2}$ for chromium-51 is 28 days and that means that after 28 days a sample of 1 gram of the radioisotope will indeed weigh 0.5 g. Table 4.1 reports half-lives of numerous isotopes. Samples of radioisotopes weigh less and less with time as they decompose producing more stable isotopes. Similarly, $t_{1/2}$ for strontium-90 is 38 years which means that a one-gram sample will take 38 years to reduce its mass to 0.5g. We can use the concept of half-life to compare the speed of decomposition of different radioisotopes. For example $t_{1/2}$ for strontium-90 is 38 years whereas $t_{1/2}$ for chromium-51 is 28 years. Hence, strontium-90 will exist longer than chromium-51.

Sample Problem 33

On one hand, Docetaxel is a chemotherapy medication used to treat a number of types of cancer with a half-life of 309600 seconds. On the other hand, Vandetanib is an anti-cancer medication that is used for the treatment of certain tumors

$^{131}_{53}Ir \longrightarrow ^{131}_{54}Xe + ^{0}_{-1}e$





of the thyroid gland with a half-life of 1641600 seconds. Which medication will remain longer in the body?

SOLUTION

We will compare the half-lives of both medication using scientific notation to express the numbers so that we can clearly see how the numbers compare. $t_{1/2}$ for Docetaxel is 3.096×10^5 seconds, whereas $t_{1/2}$ for Vandetanib is 1.6416×10^6 seconds. We have that Vandetanib has a larger half-life and hence it will remain longer in the body.

STUDY CHECK

On one hand, Methadone is a synthetic Opioid agonist used for opioid maintenance therapy in opioid dependence and for chronic pain management with a half-life of 2×10^5 seconds. On the other hand, Fluoxetine, sold under the brand names Prozac, is an antidepressant with a half-life of 4 days. Which medication will remain longer in the body and how many time is the half-life of the largest remaining drug in comparison with the other?

Quantifying half-live The formula that related the amount of radioisotope with $t_{1/2}$ is:

$$\boxed{N(t) = N_o \cdot 0.5^{\left(\frac{t}{t_{1/2}}\right)}} \tag{4.1}$$

where $N(t)$ is the amount of isotope at a given time t, N_o is the initial amount of isotope, t is the time and $t_{1/2}$ is the half-life. $N(t)$ is often referred to as the activity of the radioisotope at a given time t. At the same time, while the radioisotope disappears, a new isotope–this time more stable than the radioisotope–starts forming. The amount of product formed $F(t)$ at a given time is:

$$\boxed{F(t) = N_o \cdot \left[1 - 0.5^{\left(\frac{t}{t_{1/2}}\right)}\right]} \tag{4.2}$$

After several half-lives So if the half-life is the time it takes for a radioisotope to decompose in half, what would happen after several half-lives? For example, imagine we have 20 grams of iridium-131 with a half-life of 8 days. When we prepare or hypothetically unseal the sample, we will have 20 grams of ^{131}Ir. After one half-life (8 days) we'll have 10 grams of ^{131}Ir. After two half-lives (16 days), we'll have 5 grams of ^{131}Ir. Similarly, after three half-lives (22 days), we'll have 2.5 grams.

Sample Problem 34

$^{131}_{53}$I has a half-life of 8days. How many milligrams of a 50mg sample will remain after 10 days.

SOLUTION

We will following steps to solve this problem:

1. **Step one:** list of the given variables.

Given	Asking

Analyze the Problem	G	A
	$t_{1/2} = 8$days $N_o = 50mg$ $t = 10days$	$N(t)$

② **Step two:** use the half-life formula $N(t) = N_o \cdot 0.5^{\frac{t}{t_{1/2}}}$ to obtain the mass remaining of the radioisotope

10d

$$N(t) = N_0 \cdot 0.5^{\frac{t}{t_{1/2}}}$$

50mg

8d

③ **Step three:** solve for $N = 50 \cdot 0.5^{\frac{10}{8}} = 50 \cdot 0.5^{1.25} = 21mg$. The result means that after 10 days from the 50mg sample of radioisotope, only 21mg will remain due to radioactive decay.

STUDY CHECK $^{222}_{86}$Rn has a half-life of 3.8 days. How many milligrams of a 25mg sample will remain after 15 days.

4.5 Radiation measurement, units and radiation effects

Beta and gamma radiation can be detected with a Geiger counter, which consists of a detector tube with a specific ionizing gas. When radiation enters the Geiger counter, it generates charged particles that produce a detectable electrical current. The larger the current the stronger the radioactive source. In the following, we will address the different units of radioactivity–activity, adsorbed dose, and biological damage– reported in Table 4.2 as well as the effects of radiation.

`Activity units` The radioactivity of an isotope often referred to as *activity*, can be measured in two different units: Curies (Ci) or becquerel (Bq). Curie was somehow the original unit employed to measure the radioactivity of radium and becquerel is a more modern unit of radioactivity. Bq is the SI unit of activity. Both units are related by:

$$\boxed{1Ci = 3.7 \times 10^{10} Bq} \quad \text{or} \quad \boxed{\frac{1Ci}{3.7 \times 10^{10} Bq}} \quad \text{or} \quad \boxed{\frac{3.7 \times 10^{10} Bq}{1Ci}} \tag{4.3}$$

Activity refers to the isotope and is also measured in disintegration per minute (cpm).

`Adsorbed dose` Whereas activity refers to the isotope, the adsorbed dose refers to the body that receives radiation. The unit for adsorbed dose is called rad (radiation adsorbed dose). This unit refers to the amount of radiation adsorbed per gram of material. The SI unit for adsorbed dose is called the gray (Gy).

▼An old Geiger counter used to measure radiation

`Radiation equivalent in humans, rem` Not all radiations have the same impact on the human body. The radiation equivalent in humans takes into account the different types of radiation to adjust the biological damage of radiation. The rem is the number of rads times a factor that depends on the radiation. This factor is one for beta and gamma radiation, being 20 for alpha particles. The following formula related REM with RAD:

$$\boxed{rem = rad \times Factor} \tag{4.4}$$

`Exposure to radiation` We are all somehow exposed to radiation every day. The reason for this background radiation is that many natural radioisotopes form the atoms of many materials such as brick, concrete, water, or even the air. Still, the daily exposure is very low and you should not be concerned by the effect of this background radiation.

Table 4.2 Radiation Measurement

Activity	curie (Ci)
	Becquerel (Bq) = 2.7×10^{-11} Ci
	1Ci = 2.22×10^{12} dpm
Dosage (D)	rad
	gray (Gy) =100rad
Damage (H)	rem
	Sievert (Sv) = 100rem

▼Nurses expose to radiation wear film badges to detect radiation exposure

`Dangers of radiation` Radiation units different than Ci or Bq are used to measure the impact of radiation on humans. The rem (radiation equivalent in humans) is a radiation unit that measures the direct biological effects of different kinds of radiation. The number of rems a person receives would determine the impact of the radiation on this person's health. As an example, radiation exposure under 25 rem is harmless and they cannot be detected. If a victim is exposed to 100 rem or higher, the person will suffer the symptoms of radiation sickness and will feel nausea, vomit, fatigue, and a reduction in white cell count. If a person is exposed to a dosage greater than 300 rem, that can lower the white-cell count to zero; the victim will suffer diarrhea, hair loss, and infection. Exposure to radiation of about 500 rem is expected to cause death in half of the people receiving that dose. Radiation dosages of about 600 rem would be fatal to all humans within a few weeks.

▼New Geiger counters

Sample Problem 35

Ioflupane is a radiopharmaceutical that helps visualize the brain of Parkinson patients. An injection of this drug has a 5mCi activity. Convert this value to MBq.

SOLUTION

You need to use the conversion factor between Bq and Ci, $\frac{1Ci}{3.7 \times 10^{10} Bq}$, as well as the conversion factor between mCi and Ci $\frac{1mCi}{1 \times 10^{-3} Ci}$, and Bq and MBq

$\frac{1MBq}{1\times10^6 Bq}$.

$$5\cancel{mCi} \times \frac{1\times10^{-3}\cancel{Ci}}{1\cancel{mCi}} \times \frac{3.7\times10^{10}\cancel{Bq}}{1\cancel{Ci}} \times \frac{1MBq}{1\times10^{6}\cancel{Bq}} = 185MBq$$

STUDY CHECK

Quadramet is a radiopharmaceutical used treat pain when cancer has spread to the bone. A injection of this drug has a 740 MBq activity. Convert this value to mCi.

4.6 Radiation protection

Radioactivity results from the emission of very energetic and small particles. It can be extremely harmful when no proper protection is used. Therefore, all hospital personnel working with radioactive isotopes–radiologists, doctors, and nurses–need to be protected against radiation. Table 4.3 reports some useful information regarding radiation protection.

`alpha particles` Alpha radiation is made of very heavy particles (He nuclei) that can only travel between 2-4cm in the air before disappearing Inside your body they can penetrate only 0.05mm. A simple piece of thin clothing, a lab coat, gloves, or even our skin can protect us against alpha particles.

`beta particles` Beta radiation is made of lighter particles (electrons) that move much faster than alpha particles. Beta particles travel between 200-300cm in the air and between 4-5mm in body tissue. Heavy clothing such as lab coats or gloves is needed to protect you against this radiation.

`gamma particles` Gamma radiation can pass through many materials including body tissues. Gamma rays travel around 500 m in the air and more than 50cm in tissue. Only very dense shielding, such as lead or concrete, will protect you from this radiation.

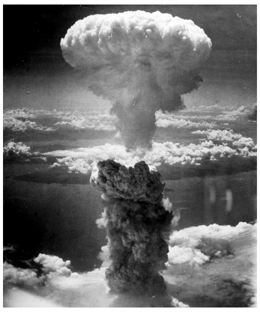
Table 4.3 Radiation protection

Particle	Travels in air	Travels in tissue	Protected with
α	2-4cm	0.05mm	thin clothing
β	200-300cm	4-5mm	heavy clothing
γ	500m	50cm	lead or concrete

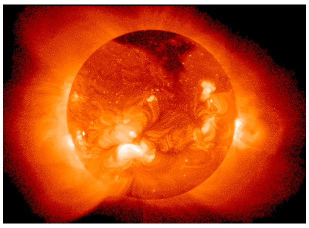
4.7 Radioactive gases: radon

Radon is a colorless, odorless radioactive gas produced by the radioactive decay of uranium. It is present in nearly all soils and very small levels of radon are found in the air we breathe every day. The problem occurs when radon gas enters our home and gets trapped. If you are breathing in too much radon, you will not feel sick right away. Only long-term exposure to high levels of radon can cause lung cancer and the risk is higher for those who smoke. While questions remain over the quantities and length of exposure, radon concerns are a fact of homeownership. Most residential real estate transactions require radon testing, and many states require radon mitigation for new construction. The recommended reference radon level is 100 $Bq \cdot m^3$ in dwellings. Testing your home for radon is easy and doesn't cost very much. You can test for radon yourself or hire a professional to do it for you. There are relatively simple tests for radon gas. Radon detection devices are commercially available. Digital radon detectors provide ongoing measurements giving daily, weekly, short-term and long-term average readouts via a digital display. Short-term radon test devices used for initial screening purposes are inexpensive, in some cases free.

▼Nuclear weapons are based on the principles of fission



▼The Sun generates its energy by nuclear fusion of hydrogen nuclei into helium.



4.8 Fusion and fission

This section will address nuclear fusion and fission. These two processes involve the liberation of a very large amount of energy and are the base for using nuclear processes as a source of energy.

`Nuclear fusion` Large energy quantities are released when two light isotopes combine to produce a heavier isotope. Nuclear fusion is the mechanism of energy production in the stars. Very high temperatures are required to initiate nuclear fusion and that is the reason why this source of energy has not been exploited on the earth yet. An example of fusion reactions found in the stars are:

$$ {}^{1}_{1}\mathrm{H} + {}^{2}_{1}\mathrm{H} \longrightarrow {}^{3}_{2}\mathrm{He} $$
$$ {}^{3}_{2}\mathrm{He} + {}^{1}_{1}\mathrm{H} \longrightarrow {}^{4}_{2}\mathrm{He} + {}^{0}_{1}\mathrm{e} $$

`Nuclear fission` Nuclear fission was discovered before the second world war when ${}^{235}_{92}\mathrm{U}$ was bombarded with neutrons. The result is the split of the atom into two different isotopes and the release of more neutrons:

$$ {}^{1}_{0}\mathrm{n} + {}^{235}_{92}\mathrm{U} \longrightarrow {}^{141}_{56}\mathrm{Ba} + {}^{92}_{36}\mathrm{Kr} + 3\,{}^{1}_{0}\mathrm{n} $$

This process releases 26 million times more energy than the combustion of methane. As neutrons are produced in a fission process, they can already activate another uranium atom producing more neutrons. This is the essence of a chain reaction: a self-sustained fission process. If less than one neutron causes a new fission process the fission process will stop and the reaction is said to be subcritical. Differently, when exactly one neutron from each fission even produces another fission the process will sustain and the reaction is known as critical. When more than one neutron produced generates a new fission the fission process will escalate and the reaction is known as supercritical. During World War II, the Manhattan project was a united states research project to build a bomb based on the principles of fission. A fission bomb operates by suddenly combining subcritical masses of uranium, producing an enormous explosion.

Sample Problem 36

Identify the following reactions as fusion or fission:

$$ {}^{1}_{1}\mathrm{H} + {}^{1}_{1}\mathrm{H} \longrightarrow {}^{2}_{1}\mathrm{H} + {}^{0}_{1}\mathrm{e} $$
$$ {}^{1}_{0}\mathrm{n} + {}^{235}_{92}\mathrm{U} \longrightarrow {}^{137}_{52}\mathrm{Te} + {}^{97}_{40}\mathrm{Zr} + 2\,{}^{1}_{0}\mathrm{n} $$

SOLUTION

The first nuclear reaction combines two hydrogen isotopes and hence it will be a fusion reaction. The second reaction results of the fragmentation–fission–of uranium and it will be fission.

STUDY CHECK

Identify the following reaction as fusion or fission:

$$^{3}_{2}\text{He} + {}^{3}_{2}\text{He} \longrightarrow {}^{4}_{2}\text{He} + 2\,{}^{1}_{1}\text{H}$$

CHAPTER 4

Radiation, particles & radioisotopes

4.1 Indicate the name of the following nuclear symbols: (a) ${}^{0}_{+1}e^{+}$ (b) ${}^{0}_{0}\gamma$ (c) ${}^{1}_{1}H$

4.2 Indicate the name of the following nuclear symbols: (a) ${}^{4}_{2}He$ (b) ${}^{1}_{0}n$ (c) ${}^{0}_{-1}e$

4.3 Indicate the nuclear symbol for (a) Oxygen-18 (b) Magnesium-24 (c) Lithium-7

4.4 Indicate the nuclear symbol for (a) Iodine-134 (b) Lithium-4 (c) Carbon-14

4.5 Calculate the number of electrons for the following isotopes: (a) ${}^{24}_{17}Mg$ (b) ${}^{129}_{53}I$

4.6 Calculate the number of electrons for the following isotopes: (a) ${}^{40}_{19}K$ (b) ${}^{87}_{37}Rb$

Nuclear Reactions

4.7 Classify the following nuclear reactions as: (a) α decay (b) β decay (c) γ decay (d) positron emission (e) electron capture

(a) ${}^{238}_{92}U \longrightarrow {}^{234}_{90}Th + {}^{4}_{2}He$

(b) ${}^{42}_{19}K \longrightarrow {}^{42}_{20}Ca + {}^{0}_{-1}e$

(c) ${}^{15}_{8}O \longrightarrow {}^{15}_{7}N + {}^{0}_{+1}e$

(d) ${}^{228}_{88}Ra \longrightarrow {}^{228}_{89}Ac + {}^{0}_{-1}e$

(e) ${}^{13}_{6}C + {}^{1}_{1}H \longrightarrow {}^{14}_{7}N + {}^{0}_{0}\gamma$

4.8 Classify the following nuclear reactions as: (a) α decay (b) β decay (c) γ decay (d) positron emission (e) electron capture

(a) ${}^{14}_{6}C \longrightarrow {}^{14}_{7}N + {}^{0}_{-1}\beta$

(b) ${}^{11}_{6}C \longrightarrow {}^{11}_{5}B + {}^{0}_{+1}\beta^{+}$

(c) ${}^{55}_{26}Fe + {}^{0}_{-1}\beta \longrightarrow {}^{55}_{25}Mn + X-ray$

(d) ${}^{234}_{88}Th^{*} \longrightarrow {}^{234}_{88}Th + {}^{0}_{0}\gamma$

(e) ${}^{226}_{88}Ra \longrightarrow {}^{222}_{86}Rn + {}^{4}_{2}\alpha$

Unknown isotopes in nuclear reactions

4.9 Identify the unknown radioactive particle involved in the following nuclear equations: (a) ${}^{11}_{6}C \longrightarrow {}^{A}_{Z}X + {}^{0}_{+1}\beta^{+}$ (b) ${}^{A}_{Z}X + {}^{0}_{-1}\beta \longrightarrow {}^{55}_{25}Mn + X-ray$ (c) ${}^{234}_{88}Th^{*} \longrightarrow {}^{A}_{Z}X + {}^{0}_{0}\gamma$ (d) ${}^{226}_{88}Ra \longrightarrow {}^{222}_{86}Rn + {}^{A}_{Z}X$

4.10 Identify the unknown radioactive particle involved in the following nuclear equations: (a) ${}^{9}_{4}Be + {}^{A}_{Z}X \longrightarrow {}^{12}_{6}C + {}^{1}_{0}n$ (b) ${}^{31}_{15}P + {}^{1}_{1}H \longrightarrow {}^{31}_{16}S + {}^{A}_{Z}X$ (c) ${}^{3}_{1}H + {}^{2}_{1}H \longrightarrow {}^{A}_{Z}X + {}^{1}_{0}n$ (d) ${}^{14}_{6}C \longrightarrow {}^{A}_{Z}X + {}^{0}_{-1}\beta$

Half-life of a radioisotope

4.11 The following organic chemicals are used in chemotherapy. Indicate the chemical that will last longer in the body. (a) Ondansetron ($t_{\frac{1}{2}}$=360min) (b) Capecitabine ($t_{\frac{1}{2}}$=2400s)

4.12 The following organic chemicals are used in chemotherapy. Indicate the chemical that will last longer in the body. (a) Carmustine ($t_{\frac{1}{2}}$=0.5h) (b) Capecitabine ($t_{\frac{1}{2}}$=2400s)

4.13 Xenon-133, which is used for lung imaging, has a half-life of 5.2 days. If 50.0 mg of Xe-133 were prepared at 8:00 A.M. on Monday, how many mg remain at 8:00 A.M. on the following day?

4.14 Gold-198, which is used for liver disease diagnosis, has a half-life of 2.7 days. If 100.0 mg of Au-198 were prepared at 8:00 A.M. on Monday, how many mg remain at 8:00 A.M. on Wednesday? And at 2:00 P.M. on Wednesday?

4.15 The half-life of bromine-74 is 25 min. How much of a 100 mg sample is still active after 100 min?

4.16 Technetium-99m has a half-life of 6 h, being used to image the skeleton and heart muscle in particular. How much of a 5 mg sample is still active after 50 min?

4.17 The half-life of bromine-74 is 25 min. 20mg of the isotopes remain after 10 minutes of preparing the sample. Calculate the initial mass of the bromine-74 sample.

4.18 The half-life of Au-198 is 2.7 days. 100mg of the isotopes remain after 5days of preparing the sample. Calculate the initial mass of the isotope sample.

Radiation measurement, units and radiation effects

4.19 $^{199}Tc^{*}$ is a radioisotope used for liver disease diagnosis. The administered activity of the isotope is 740MBe. How much is this activity in mCi?

4.20 $^{201}Tl^{*}$ is a radioisotope used for myocardial scan. The administered activity of the isotope is 110MBe. How much is this activity in mCi?

4.21 Select a symptom of mild radiation sickness from the list below: (a) a lowered white cell count. (b) a lowered red blood cell count. (c) a raised white cell count. (d) a raised red blood cell count. (e) a white cell count of zero.

4.22 Select a symptom that does not result from acute radiation sickness from the list below: (a) permanent hair loss. (b) loss of appetite. (c) fatigue. (d) fever. (e) nausea. (f) diarrhea.

4.23 Alpha radiation is the most damaging because alpha particles (a) have the largest charge. (b) have the greatest energy. (c) have the greatest mass. (d) consist of high energy electrons.

4.24 Gamma radiation is the most penetrating because gamma particles (a) have the largest charge. (b) have the greatest energy. (c) have the greatest mass. (d) consist of high energy electrons.

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Ch. 5. Chemical naming

ALL elements in the periodic table except for the noble gases–He, Ne, Ar, Kr, Xe, and Rn–combine to produce chemical compounds. Most of these chemicals are useful in your everyday life, and you drink water to quench your thirst, use Clorox to clean your house, or baking soda to get rid of a stinky refrigerator. In this chapter you will learn not only how to name these chemicals but also read chemical formulas–we call this to formulate chemicals. Still, chemical elements such as hydrogen and oxygen do not combine randomly and they only choose specific elemental partners to form a compound. As an example, hydrogen combines with oxygen using specific proportions to produce H_2O and not HO_2. In this chapter, you will also learn the rules that chemical elements use to combine.

5.1 Ions & ionic charges

Atoms gain and lose electrons to produce ions. An ion is just an atom with a positive or negative charge. Ions result from an electron transfer. Positive ions have lost negatively charged electrons, whereas negative ions have gained electrons. The reason for this electron transfer is that atoms try to achieve a very stable electronic configuration with eight electrons in the valence, and this is called the octet electron configuration. Examples of ions are: H^+, Ca^{2+} or O^{2-}. This section covers the properties of ions and ionic charges.

Cations and anions Atoms that lose electrons become positively charged. These ions are called cations. Examples of cations are Li^+ or Mg^{2+} called lithium cation and magnesium cation, respectively. Atoms that gain electrons become negatively charged, as electrons have a negative charge. These ions are called anions. Examples of anions are F^- called fluoride or N^{3-} called nitride. The way to name anions is by using the name of the element and the suffix -ide.

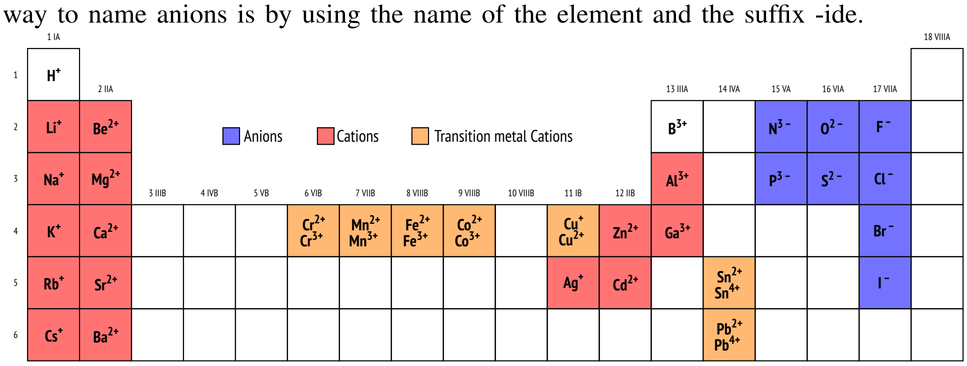


Figure 5.1 Ionic charges (valences) for different elements

Ionic charges: the valences How do we know that hydrogen produces a

H^+ ion and nitrogen a N^{3-} anion? The charge of an ion is called an ionic charge, and the numbers are coming from the periodic table. H, Na, or K are in group IA (left of the table) and hence the ionic charge will be $1+$. Similarly, Mg or Ca are in group IIA (left of the table) and hence the ionic charge will be $2+$. Differently, F, Cl, or Br are in group 7A (right of the table) and their charge will be $1-$. Oxygen is in group 6A (right of the table) and the ionic charge will be $2-$. Figure **??** contains all ionic charges. What if the element is not on this list such as in the case of Iron (Fe)? In that case, very probably it will have several ionic charges and this charge has to be indicated in the chemical name. An example would be Fe, which ionic charge is not in Figure **??** as iron can have several ionic charges.

Sample Problem 37

Identify the correct ionic state of the following elements: (a) Cl (b) K (c) O (d) C

SOLUTION

Cl is on the 7A group and hence its charge is $1-$, whereas potassium belongs to 1A and its charge will be $1+$. Oxygen and carbon will have $2-$ and $4-$ charges. The final ionic states are: Cl^-, K^+, O^{2-} and C^{4-}.

STUDY CHECK

Identify the correct ionic state of the following elements: (a) N (b) Br

5.2 Ionic compounds

Ionic compounds are chemicals resulting from the combination of a nonmetallic element with a metallic element. An example is NaCl, which results from combining sodium (a metal) with chloride (a nonmetal). Ionic compounds normally have high melting points and are solid under normal conditions. A typical ionic compound would be NaCl, cooking salt. Atoms of an ionic compound are connected through an ionic bond. In an ionic bond, one element gives away electrons (the cation) and the other one receives electrons (the anion). As an example, in the NaCl molecule Na gives away an electron to Cl, and the molecule results from the combinations of Na^+ and Cl^+. In an ionic compound, the element on the left is positive and the one on the right is negative.

`Combining ions` Ionic compounds are the result of combining two ions: a positive (cation) and a negative (anion) ion. Each ion has a charge, depending on its location on the table. When combining two atoms you first need to arrange the ions starting from positive and followed by negative. The charges of an ion would become the coefficient of the other ion. For example Mg^{2+} and N^{3-} are combined as Mg_3N_2:

$$3Mg^{2+} + 2N^{3-} \longrightarrow Mg_3N_2$$

Another example would be the combination of Na^+ and O^{2-} that would be Na_2O. You need to simplify the indexes of the formula by dividing by the smallest one, always using integer values. For example, Mg^{2+} and O^{2-} give Mg_2O_2 that should be written as MgO

$$Mg^{2+} \quad O^{2-} \longrightarrow Mg_{\cancel{2}}O_{\cancel{2}} \longrightarrow MgO$$

Another example that involves simplifying the formula is the chemical result of combining Ca^{2+} and C^{4-}. After combining the charges we obtain Ca_4C_2 that needs to be simplified by dividing by the smallest number leading to Ca_2C.

Sample Problem 38

Combine the following ions or give the ions given the final compound: (a) Li^+ and O^{2-} (b) Ca^{2+} and O^{2-} (c) Li_3N (d) Mg_2C

SOLUTION

The result of combining Li^+ and O^{2-} is Li_2O. For Ca^{2+} and O^{2-}, the resulting chemical is CaO. Li_3N results from the combination of Li^+ and N^{3-}, and Mg_2C results from Mg^{2+} and C^{4-}.

STUDY CHECK

Combine the following ions or give the ions given the final compound: (a) Na^+ and F^- (b) Na_3N

Simple ionic naming (type I ionic) Type I ionic compounds result from the combination of a metal with given valence (Li, Ca, Mg, etc, see Figure 5.1) and a nonmetal. To name an ionic compound (type I ionic) you need to (a) use the name of the first element in the compound, (b) use the first syllable of the second element, and (c) finish the name of the molecule in the suffix *-ide*. As an example, the formula NaCl is named sodium chloride, and $MgCl_2$ is named magnesium chloride. Another example would be:

calcium chloride	$CaCl_2$ (ionic)

To formulate an ionic compound based on a name, we need to combine both ions by exchanging the valences (the ionic charges). For example,$MgCl_2$ results from the combination of Mg^{2+} and Cl^- so that the number 2 in $MgCl_2$ near the Cl atom is coming from the Mg^{2+}. In other words:

$$Mg^{2+} \quad 2Cl^{-} \longrightarrow MgCl_2$$

The sign of the charges only indicates which element goes first in the formula: the positive element (cation) first followed by the negative element (anion). For example, the result of combining Na^+ and Cl^- is NaCl and not ClNa as Na has a positive ionic charge and has to appear first in the formula.

Sample Problem 39

Name or give the formula for the following ionic compounds: (a) MgO (b) Mg_3N_2 (c) Lithium nitride (d) Magnesium carbide

SOLUTION

The name for MgO is magnesium oxide. Mg_3N_2 is called magnesium nitride. The formula for Lithium nitride is Li_3N and the formula for Magnesium carbide is Mg_2C, result of simplifying Mg_4C_2 dividing by two, the smallest number.

STUDY CHECK

Name or give the formula for the following ionic compounds: (a) Sodium fluoride (b) Na_3N

The ionic chemical NaCl results from the combination of Na^+ and Cl^-. The ionic charges of Na and Cl are given in Figure **??** according to the group. If the ionic chemical contains a transition metal with variable ionic charge, that is, which is not in Figure **??** then the ionic naming becomes a bit more complex. The reason is that one needs to specify the charge of the metal, explicitly in the name of the chemical. An example would be $NiCl_2$ named as Nickel(II) chloride or Co_2O_3 named as Cobalt(III) oxide.

`Name complex ionic chemicals` This section covers how to name ionic chemicals containing a metal with variable charge. In this case, you need to specify the charge of the metal in the name. To calculate this number you will solve a simple math equation. For example, the name of Mn_2O_3 is Manganese(III) oxide. How do we get this name? Manganese has several charges as it is not in Figure **??**, let us use x for its charge Mn^x and oxygen has a charge of two O^{2-}. After combining Mn^x and O^{2-} the resulting formula would be Mn_2Ox. By comparison with the given formula, Mn_2O_3, x has to be three and hence the charge of Mn has to be three. Therefore, the final name would be Manganese(III) oxide.

`Properties of ionic compounds` Ionic compounds normally have high melting points and are solid under normal conditions. A typical ionic compound would be NaCl, cooking salt.

`The ionic bond` Atoms of an ionic compound are connected through an ionic bond. In an ionic bond, one element gives away electrons (the cation) and the other one receives electrons (the anion). As an example, in the NaCl molecule Na gives away an electron to Cl, and the molecule results from the combinations of Na^+ and Cl^+. In an ionic compound, the element on the left is positive and the one on the right is negative.

Sample Problem 40

Name or give the formula for the following ionic compounds: (a) MnO (b) Fe_3N_2 (c) Cobalt(II) carbide (d) Iron(II) oxide

SOLUTION

All the chemicals on this example contain a metal that can have several charges, and hence, we need to specify the ionic charge on the name. MnO results from Mn^x and O^{2-}. After combining the ions, the formula would be Mn_2Ox, a formula that needs to be compared to MnO. The formulas do not look similar, so lets make them more similar by dividing by two so that $MnO\frac{x}{2}$ resembles MnO. By comparing x has to be 2 and hence the name is Manganese(II) oxide. The name for Fe_3N_2 would be Iron(II) nitride. The valence of Iron comes from combining Fe^x and N^{3-} that gives Fe_3Nx. By comparison with Fe_3N_2 x has to be two and the name is Iron(II) nitride. the formula for Cobalt(II) carbide would be Co_2C as Cobalt(II) is Co^{2+} and carbide is C^{4-}. After combining the ions one obtains Co_4C_2 that gives Co_2C. Finally, the formula for Iron(II) oxide is FeO as Iron(II) is Fe^{2+} and oxide is O^{2-} that gives Fe_2O_2 and simplifying one obtains FeO.

STUDY CHECK

Name or give the formula for the following ionic compounds: (a) Manganese(IV) oxide (b) $AuCl_3$

5.3 Covalent compounds

Covalent compounds are chemicals resulting from the combination of nonmetallic elements. And example is CO_2, which results from combining carbon (a nonmetal) with oxygen (a nonmetal). At normal conditions, covalent compounds may exist as solids, liquids, or gases. Covalent compounds do not exhibit any electrical conductivity, either in pure form or when dissolved in water. A typical covalent compound would be H_2O, water. Atoms in a covalent compound are connected by means of a covalent chemical bond. In a covalent bond, both atoms connected share the electrons. As an example, the HCl molecule has an hydrogen and a chlorine atom connected by means of a covalent bond, in which each atoms share the electrons of the bond.

Covalent naming To name a covalent compound you need to (a) use the name of the first element in the compound, (b) use the first syllable of the second element, and (c) finish the name of the molecule in the suffix *-ide*. More importantly, you need to use prefixes (see Table **??**) that indicate the number of atoms in the molecule. See the Table below for a list of the different equivalencies between prefixes and numbers. As an example, the formula CH_4 is named carbon tetrahydride. Similarly, a covalent chemical name can be translated into a formula (we call this to formulate a chemical with a given name), and the formula for carbon monoxide would be CO. When the vowels *a* and *o* appear together, the first vowel is omitted as in carbon monoxide instead of carbon ~~monooxide~~. Another example would be N_2O named as dinitrogen oxide, and the name sulfur hexafluoride corresponds to the formula SF_6. The prefix mono is omitted in the first element of the name, and for example, you will not name the chemical CO as ~~monocarbon~~ monoxide, you would just say carbon monoxide. A final example of a covalent compound:

dinitrogen pentoxide N_2O_5 (covalent)

Table 5.3 Prefixes used to name covalent compounds

Prefix	number	Prefix	number	Prefix	number
Mono	1	Tetra	4	Hepta	7
Di	2	Penta	5	Octa	8
Tri	3	Hexa	6	Nona	9

Sample Problem 41

Name of give the name of the following covalent chemicals: (a) NO (b) CS_2 (c) Sulfur Dioxide (d) Nitrogen Trichloride

SOLUTION

All chemicals in this example are covalent as they result of the combination of nonmetals. In order to name them, we need to use prefixes and finish the sufix with -ide. The first chemical is called nitrogen monoxide. CS_2 is called carbon disulfide. The formula for sulfur dioxide and nitrogen trichoride are respectively SO_2 and NCl_3.

STUDY CHECK

Name of give the name of the following covalent chemicals: (a) SCl_2 (b) diboron thrioxide

5.4 Naming acids & bases

In this section, we will learn how to name acids and bases. Acids normally have common names (e.g. sulfuric acid) and their naming does not follow modern rules. Names and formulas of acids are listed in tables. Differently, bases (e.g. sodium hydroxide) are named in a standard way.

`Bases or hydroxides` Bases (hydroxides) result from the combination of metal and the hydroxide anion (OH^-). Examples are NaOH or $Ca(OH)_2$. The name of a base starts with the name of the cation finishing with the word *hydroxide*. An example is NaOH named as *sodium hydroxide*, or $Ca(OH)_2$, named as *calcium hydroxide*. The word *hydroxide* refers to the OH^- ion, and hence Sodium hydroxide results from combining Na^+ and OH^-, and Calcium hydroxide from combining Ca^{2+} and OH^-. More examples of hydroxides:

Magnesium hydroxide	$Mg(OH)_2$(hydroxide)

As a final note, not all bases are hydroxides. For example, ammonia NH_3 is a base even when it does not contain hydroxides in its structure. A way to remember this is sometimes to write ammonia like NH_4OH.

`Acids` Acids–in particular inorganic acids–are chemicals that normally contain hydrogen at the beginning of their formula. For example, HCl or H_2SO_4. HCl is a hydracid and is named as *hydrochloric acid*, whereas H_2SO_4 is an oxoacid that contains oxygen named as *sulfuric acid*. The names of acids are not standard and they come from common names employed in the field for many years. Table **??** contains a list of the most important oxoacids and hydracids. More examples of acids:

Nitric acid	HNO_3(oxoacid)
Hydrofluoric acid	HF(hydracid)

Sample Problem 42

Name or give the formula for the following acids and bases. Indicate whether the compound is an acid or a base. (a) HCN (b) KOH (c) Carbonic acid (d) Lithium hydroxide

SOLUTION

HCN is an acid named hydrocyanic acid. KOH is a base called potassium hydroxide. The formula for Carbonic acid is H_2CO_3, and Lithium hydroxide is a base with formula LiOH.

STUDY CHECK

Name or give the formula for the following ionic compounds: (a) phosphoric acid (b) $Mg(OH)_2$

5.5 Oxidation states

Isolated atoms tend to have a neutral state. However, the atoms of elements have the capacity of gaining and losing electrons forming cations and anions. The atoms that form a compound can have different states resulting in losing and gaining electronic charge. We refer to this as the oxidation state of an element in a compound.

`Oxidation states of oxoacids` Consider the following set of acids: HClO, $HClO_2$, $HClO_3$ and $HClO_4$. We say Cl in these acids have different oxidation states or different oxidation numbers. This section will cover the calculation of the oxidation state of the central atom of an oxoacid.

Let us address the oxoacid: $H\underline{Cl}O_3$. The goal is to calculate the oxidation number of the underlined element, Cl. To do this we will follow a set of simple rules. First, we will use the valences as the oxidation number of the elements (see Figure 5.1) to the right and the left of the central atom. Then, we will assign an unknown oxidation state of x to the central atom. After that, we will set up an equation so that the sum of all oxidation numbers equals the charge of the acid if any. In this formula, we will include the atomic coefficients. In the case of $H\underline{Cl}O_3$, the equation would be:

$$1 + x + 3 \cdot (-2) = 0$$

as the number of oxygens is three, we will have to time by three the valence of oxygen. The number zero results from the charge of the acid. If we solve for x, we obtain $x = 5$. That is, the oxidation state of Cl on $H\underline{Cl}O_3$ is 5 and this is represented as$H\underline{Cl}^{V}O_3$.

`Oxidizing and reducing character of oxoacids` The importance of the oxidation state of the central elements of oxoacid results from the fact that acids with high or low oxidation states, tend to be very reactive, sometimes capable of completely dissolving metals. We call this oxidizing (or reducing) acids. For example, HNO_3 and H_2SO_4 are both oxidizing acids, and these acids will solve for example a piece of copper. Similarly, acids with very small or negative oxidation numbers can be very reactive as well. These acids are called reducing acids or agents. Let us compare two oxoacids to elaborate more on the terminology used to describe redox numbers. For example, let us compare $H\underline{Cl}^{V}O_3$ and $H\underline{Cl}^{III}O_2$. We say Cl on $H\underline{Cl}^{V}O_3$ has a larger redox number than $H\underline{Cl}^{III}O_2$. We can also say, Cl in $H\underline{Cl}^{V}O_3$ is more oxidized than Cl on $H\underline{Cl}^{III}O_2$. Finally, we can also say, $H\underline{Cl}^{V}O_3$ is more reducing than $H\underline{Cl}^{III}O_2$. Again, the terms associated with high redox numbers are oxidized and reducing, and the terms associated with low redox numbers are reduced and oxidizing.

It is important to note that ultimately the oxidation state of an element is related to the number of electrons of the element. The more electrons the smaller–the more negative–the oxidation state. In other words, large oxidation states result from losing electrons.

Sample Problem 43

Calculate the redox number of S in the following acids and indicate the more oxidizing acid: $H_2\underline{S}_2O_6$ named dithionic acid and $H_2\underline{S}O_4$ named sulfuric acid.

SOLUTION

We will set up the redox formula for the first acid ($H_2\underline{S}_2O_6$), given that the redox number of H is $+1$ and the redox number of O is -2.

$$2 \cdot 1 + 2 \cdot x + 6 \cdot (-2) = 0$$

Solving for x:

$$2 + 2 \cdot x - 12 = 0 \quad \text{we have that } x = \frac{12 - 2}{2}$$

The oxidation state of S in $H_2\underline{S}_2O_3$ is $+5$. For the second acid ($H_2\underline{S}O_4$):

$$2 \cdot 1 + x + 4 \cdot (-2) = 0$$

Solving for x:

$$2 + x - 8 = 0 \quad \text{we have that } x = \frac{8 - 2}{1}$$

that gives a redox of 6. If we compare both acids the smaller the redox number the more reduced is the central element and the more oxidizing the acid is. Therefore, $H_2\underline{S}_2O_3$ is more oxidizing than $H_2\underline{S}O_4$.

STUDY CHECK

Calculate the redox number of the following acids: (a) H_2MnO_4 (b) $H_2Cr_2O_7$

5.6 Naming complex salts & common chemicals

At this point, we saw the naming and formulation of ionic (e.g. NaCl) and covalent compounds (e.g. CO_2). This section covers the naming of complex salts: oxosalts and hydrosalts. In general, salts (oxosalts or hydrosalts) are the result of mixing an oxoacid and a base. They tend to look more complex than simple ionic or covalent compounds as they have at least three different elements. An example of oxosalt would be $CaSO_4$ called calcium carbonate. An example of hydrosalt would be $NaHSO_4$ which is called sodium monohydrosulfate. This section will also cover the naming of hydrates (e.g. $CaSO_4 \cdot H_2O$), which are compounds containing water molecules inside their structure. Before being able to name these complex chemicals it is convenient to practice combining ions, without paying attention to the naming.

`Combining ions` To combine two ions, you first arrange the positive ion on the left followed by the negative ion on the right, to then cross the ionic charges from the top of the ion to the bottom of the opposite ion. The positive and negative charges are not carried. If the ions have more than one element we have to use parenthesis. An example would be combining Ca^{2+} and ${PO_4}^{3-}$ leading to $Ca_3(PO_4)_2$:

$$3Ca^{2+} + 2PO_4^{3-} \longrightarrow Ca_3(PO_4)_2$$

We would simplify in case the charges compensate for each other.

An example would be combining Mg^{2+} and $SO_4{}^{2-}$ leading to $MgSO_4$

$$Mg^{2+} \quad SO_4^{2-} \longrightarrow Mg_{\cancel{2}}(SO_4)_{\cancel{2}} \longrightarrow MgSO_4$$

Table 5.2 Names of oxoacids and oxosalts (top table) and hyracids (bottom table).*

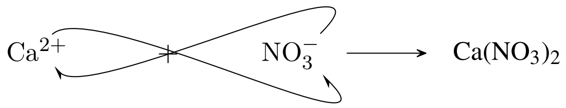
Element	Oxoacid	Oxoacid Name	Oxoasalt	Oxoasalt Name
Manganese	$HMnO_4$	Permanganic Acid	$MnO_4{}^-$	Permanganate
	H_2MnO_4	Manganic acid	$MnO_4{}^{-2}$	Manganate
Carbon	H_2CO_3	Carbonic Acid	$CO_3{}^{-2}$	Carbonate
Nitrogen	HNO_3	Nitric Acid	$NO_3{}^-$	Nitrate
	HNO_2	Nitrous Acid	$NO_2{}^-$	Nitrite
Phosphorus	H_3PO_4	Phosphoric Acid	$PO_4{}^{-3}$	Phosphate
Sulfur	H_2SO_4	Sulfuric Acid	$SO_4{}^{-2}$	Sulfate
	H_2SO_3	Sulfurous Acid	$SO_3{}^{-2}$	Sulfite
	$H_2S_2O_2$	Thiosulfurous Acid	$S_2O_2{}^{-2}$	Thiosulfite
	$H_2S_2O_3$	Thiosulfuric Acid	$S_2O_3{}^{-2}$	Thiosulfate
	$H_2S_2O_7$	Disulfuric acid	$S_2O_7{}^{-2}$	Disulfate
	$H_2S_2O_8$	Peroxodisulfuric acid	$S_2O_8{}^{-2}$	Peroxodisulfate
Chlorine	$HClO_4$	Perchloric Acid	$ClO_4{}^-$	Perchlorate
	$HClO_3$	Chloric acid	$ClO_3{}^-$	Chlorate
	$HClO_2$	Chlorous acid	$ClO_2{}^-$	Chlorite
	$HClO$	Hypochlorous acid	ClO^-	Hypochlorite
Iodine	HIO_4	Periodic Acid	$IO_4{}^-$	Periodate
Chromium	H_2CrO_4	Chromic acid	$CrO_4{}^{2-}$	Chromate
	$H_2Cr_2O_7$	Dichromic acid	$Cr_2O_7{}^{2-}$	Dichromate
Boron	H_3BO_3	Boric acid	$BO_3{}^{3-}$	Borate
Hydracid	**Hydracid Name**	**Hydracid**	**Hydracid Name**	
HCl	Hydrochloric acid	HBr	Hydrobromic acid	
HI	Hydroiodic acid	HF	Hydrofluoric acid	
HCN	Hydrocyanic acid	H_2S	Hydrosulfuric acid	

* Yellow indicate very important acids

`Naming Oxosalts` The names of the oxosalts are constructed by combining the name of the first element–you need to specify its charge in the case of a transition metal element with different possible charges–followed by the name of the oxosalt from Table **??**. For example, the name of $MgSO_4$ is magnesium sulfate, as Mg^{2+} is magnesium and $SO_4{}^{2-}$ is sulfate. Another example is$Fe_2(CO_3)_3$ called Iron(III) carbonate. A final example:

Litium nitrate	$LiNO_3$(oxosalt)

`Formulating Oxosalts` In the case that you know the name of an oxosalt and you want to obtain its formula, you first need to arrange the positive ion on the left followed by the negative ion on the right, to then cross the ionic charges from the top of the ion to the bottom of the opposite ion. For example, calcium nitrate results from the combination of Ca^{2+} calcium and $NO_3{}^-$, nitrate. By combining the two ions we obtain the final formula as $Ca(NO_3)_2$:



Sample Problem 44

Name of give the name of the following oxosalts: (a) K_2SO_4 (b) Na_2CO_3 (c) Magnesium carbonate (d) Sodium phosphate

SOLUTION

K_2SO_4 is named potassium sulfate, as K^+ is potassium and $SO_4{}^{2-}$ stands for sulfate. Na_2CO_3 is sodium carbonate. Magnesium carbonate is $MgCO_3$ and sodium phosphate is Na_3PO_4.

STUDY CHECK

Name of give the name of the following oxosalts: (a) $CaSO_4$ (b) Aluminum sulfate

Some oxosalts contain hydrogen atoms in their structure between the oxosalt cation and anion (e.g. $NaHSO_4$). For example, $NaHSO_4$ is named sodium monohydrogensulfate. To understand this name, we will first focus on the second part of the name, monohydrogensulfate which represents the anion. The name monohydrogensulfate ($HSO_4{}^-$) comes from adding a proton (H^+) to a sulfate cation ($SO_4{}^{2-}$). Mind that protons (H^+) are positively charged and therefore if we add a single H^+ to a sulfate cation ($SO_4{}^{2-}$) the charge will have to decrease a single unit, giving us $HSO_4{}^-$. As we can see, the name is directly related to the oxosalt anion and the number of hydrogens in the hydrosalt name. For example, phosphate ($PO_4{}^{3-}$) is an oxosalt anion whereas hydrogenphosphate ($HPO_4{}^{2-}$) and dihydrogenphosphate ($H_2PO_4{}^-$) are both hydrosalt anions. An explanation about the charges: as phosphate has three negative charges, hydrogenphosphate has to have one less charge (that is $2-$) and dihydrogenphosphate has to have two less negative charges (that is -1). Some final hydrosalt anions examples:

carbonate	$CO_3{}^{-2}$(oxosalt ion)
monohydrogen carbonate	$HCO_3{}^-$(hydrosalt ion)

Above we saw how to name just the ending of the oxosalt with hydrogen anion. Now we can move forward to the whole naming of the salt. We just need to add the name of the element in the first place, and for example, NaH_2BO_3 would be named sodium dihydrogenborate. If the first ion–the cation–is a transition metal cation (a type two cation) we need to include in parenthesis the valence of the cation. For example, $Fe(H_2BO_3)_2$ would be named iron(II) dihydrogenborate. More examples:

sodium carbonate	Na_2CO_3(oxosalt)
sodium monohydrogen carbonate	$NaHCO_3$ (hydrosalt)

Sample Problem 45

Name or formulate the following hydrosalts: (a) Magnesium hydrogensulfate (b) Sodium hydrogen carbonate (c) $LiHCO_3$ (d) $Mg(H_2PO_4)_2$

SOLUTION

The formula of Magnesium hydrogensulfate is $Mg(HSO_4)_2$ as the formula for monohydrogen sulfate is $HSO_4{}^-$ and the valence of magnesium is Mg^{2+}. The formula for Sodium monohydrogen carbonate is $NaHCO_3$ as it results from combining Na^+ and $HCO_3{}^-$. Mind monohydrogen carbonate results from adding a hydrogen ion H^+ to a carbonate $CO_3{}^{2-}$ ion. The name for $LiHCO_3$ is lithium monohydrogen carbonate, whereas the name for $Mg(H_2PO_4)_2$ is magnesium dihydrogenphosphate.

STUDY CHECK

Name or formulate the following hydrogensalts: (a) $LiHS_2O_3$ (b) LiH_2PO_4 (c) sodium hydrogenphosphate

Naming oxosalts with hydrogeHydrates Some chemicals contain water molecules trapped in their structure and therefore water molecules (H_2O) are often indicated in chemical formulas. These types of chemicals containing water are called *hydrates*, precisely because hydrate means water. Examples of hydrates are $BeSO_4 \cdot 4\,H_2O$ or $CuSO_4 \cdot 5\,H_2O$ called respectively beryllium sulfate tetrahydrate and cupper(II) sulfate pentahydrate. To formulate hydrates you just need to use prefixes such as mono, di, tetra–the same ones we use to name covalent chemicals, see Table **??**–to indicate the number of water molecules in the chemical and end the name with *hydrate*. As a note, warming up hydrates (e.g. $BeSO_4 \cdot 4\,H_2O$) results in the release of water producing a dehydrated or *anhydrous* compound (e.g. $BeSO_4$). A final example of hydrate naming:

Sodium sulfate pentahydrate	$Na_2SO_4 \cdot 5\,H_2O$ (a hydrate)

Sample Problem 46

Name or formulate the following hydrates: (a) Nickel(II) permanganate dihydrate (b) Sodium nitrate monohydrate (c) $Na_2CO_3 \cdot 10\,H_2O$ (d) $MgSO_4 \cdot 7\,H_2O$

SOLUTION

The formula for Nickel(II) permanganate is $Ni(MnO_4)_2$, therefore the formula for Nickel(II) permanganate dihydrate is $Ni(MnO_4)_2 \cdot 2\,H_2O$. The formula for Sodium nitrate is $NaNO_3$, therefore $NaNO_3 \cdot H_2O$ is Sodium nitrate monohydrate. The name for $Na_2CO_3 \cdot 10\,H_2O$ is sodium carbonate decahydrate and $MgSO_4 \cdot 7\,H_2O$ is magnesium sulfate heptahydrate.

STUDY CHECK

Name or formulate the following hydrates: (a) $LiNO_3 \cdot H_2O$ (b) $Na_3PO_4 \cdot 3\,H_2O$ (c) sodium sulfate tetrahydrate

Common naming Some of the chemicals are normally referred to by a common name that does not involve the use of any chemical naming rules. An example would be H_2O normally referred to as water instead of its standard name which is dihydrogen oxide. You can find more names in Table 5.3. Another example:

NaCl	Sodium chloride (standard name)	Table salt (common name)

Table 5.3 List of common chemicals

Chemical	Name	Chemical	Name
H_2O	Water	$Mg(OH)_2$	Milk of magnesia
NH_3	Ammonia	N_2O	Laughing gas
CH_4	Methane	$CaCO_3$	Marble
CO_2	Dry ice	CaO	Quicklime
NaCl	Table salt	$NaHCO_3$	Baking Soda
$NaHCO_3$	Sodium Bicarbonate	$MgSO_4 \cdot 7\,H_2O$	Epsom Salt

Sample Problem 47

Name or formulate the following common chemicals: milk of magnesia and dry ice.

SOLUTION

The formula for milk of magnesia is $Mg(OH)_2$ (magnesium hydroxide), whereas dry ice is the common name for CO_2, carbon dioxide.

STUDY CHECK

Name or formulate the following common chemicals: (a) ammonia (b) methane

5.7 Electron-dot structures of atoms & molecules

Protons, neutrons, and electrons make the atoms. Electrons–in particular valence electrons–are responsible for the main chemical properties of an atom. These electrons are loosely bound and can be exchanged easily with other atoms, in contrast to the strongly-tied core electrons. The electron-dot structure of an atom or a molecule–also called Lewis structures–is a visual representation of the electronic arrangement of the valence electrons. Atoms in a molecule will tend to be surrounded by eight electrons so that their electron configuration resembles a noble gas. This arrangement is known as the octet rule. This rule is responsible for the common negative charge of F, and the positive charge of Na: F ($[He]2s^22p^5$) can easily receive an extra electron producing ionic F^- ($[He]2s^22p^6$=[Ne]), and atomic Na ($[Ne]3s^1$) can lose an electron producing ionic Na^+ ($[He]2s^22p^6 = [Ne]$). There are a few exceptions. A remarkable one is the case of the hydrogen atom that follows the duet rule.

Valence electrons of atoms, and molecules and pairs of electrons

The electrons of an atom are divided into core electrons and valence electrons. The valence electrons of an atom are involved in chemical bonds as they are less bonded to the nucleus. *The number of valence electrons of an atom* corresponds to the group number. For example, hydrogen H belongs to the group IA, and hence it has one valence electron. Similarly, oxygen O belongs to the group VIA, having six valence electrons. Similarly, we can count the *number of valence electrons of a molecule* by adding the valence electrons of the atoms that make the molecule. For example, water (H_2O) has eight valence electrons as each oxygen has one valence electron and oxygen has six. The number of *pairs of electrons* is just the overall number of valence electrons divided by two. For example, water has eight valence electrons that correspond to 4 pairs of electrons.

Sample Problem 48

Indicate the number of valence electrons for the following atoms: N, O, C and S, and the number of pairs of electrons of the following molecules: NH_3, and CO_2.

SOLUTION

Nitrogen is in group VA and hence it has five valence electrons ($5e^-$). Oxygen belongs to the group VIA and C belong to IVA, hence they have wiz and four valence electrons, respectively. For the molecules, we have that ammonia has 8 electrons (nitrogen has five valence electrons and each hydrogen has one electron) that correspond to four pairs, whereas carbon dioxide has 16 electrons (carbon has four electrons and each oxygen has six) and eight pairs.

STUDY CHECK

Indicate the number of valence electrons for the following atoms: Cl and B.

▶Answer: Cl ($7e^-$), B ($3e^-$).

The octet rule Atoms exchange electrons when they combine to form molecules. This electron exchange is the driving force that drives the formation of molecules from single atoms. The octet rule states that each atom in a stable molecule should be surrounded by eight (octet) electrons achieving noble gas electron configurations. There are two important exceptions to this rule as H is surrounded only by two electrons (this is called the duet rule), and B by six. This rule comes from the experimental observation of numerous molecules.

Electron-dot structure of an atom The electron-dot structure of an atom is a visual representation of the arrangement of the valence electrons of the atom. To write the electron-dot structure of an atom, you just need to write down the symbol of the atom surrounded by the valence electrons located in the four directions of the space: top, bottom, right, and left. To place the electrons, you start in any of the directions and fill one electron at a time. For example, for the case of three electrons, we would have: $\cdot\dot{\mathrm{B}}\cdot$. After all four directions have been filled, you need to start pairing the electrons. For example, for the case of five electrons, we would have: $\cdot\dot{\underset{\cdot}{\mathrm{P}}}\colon$. Another example, oxygen has six valence electrons and hence, the electron-dot structure would be $\cdot\ddot{\underset{\cdot}{\mathrm{O}}}\colon$ Similarly, the electron-dot structure of fluorine would be $\colon\ddot{\underset{\cdot}{\mathrm{F}}}\colon$. For ions, you need to add (if its an anion) or subtract (if its a cation) valence electrons, and for example the electron-dot structure of the oxide anion O^{2-} is $\colon\ddot{\underset{\cdot\cdot}{\mathrm{O}}}\colon^{2-}$. The electron-dot structure of atoms is useful to predict–or make sense–of the atomic valence. Mind that the number of valence electrons of an atom is not the same as the valence of the atom. The valence of an atom is a number used to combine with other atoms forming compounds. For example, the electron-dot structure of nitrogen is $\cdot\dot{\mathrm{N}}\colon$ and this atom needs to gain three electrons to reach the noble gas configuration with eight electrons $\colon\ddot{\underset{\cdot\cdot}{\mathrm{N}}}\colon^{3-}$ hence its valence is -3.

Sample Problem 49

Write down the electron-dot structure for the following atoms: N, C and Cl^-.

SOLUTION

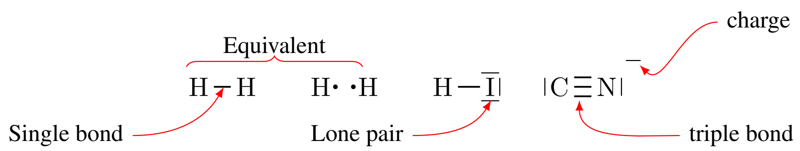
N has five valence electrons, whereas C has four. Hence the electron-dot for both will be: $\cdot\ddot{\mathrm{N}}\cdot$ and $\cdot\dot{\mathrm{C}}\cdot$. Cl^{-} has eight valence electrons, that is seven plus one, and hence its electron-dot structure will be $:\ddot{\underset{\cdot\cdot}{\mathrm{Cl}}}:^{-}$.

STUDY CHECK

Write down the electron-dot structure for N^{3-}

▶Answer: $:\ddot{\underset{\cdot\cdot}{\mathrm{N}}}:^{3-}$.

An introduction to electron-dot structures Below, you will find some examples of electron-dot structures. Mind that the lines represent pairs of electrons hence below there are two equivalent representations for the hydrogen molecule. In these structures, you will find two different types of lines. Some pairs of electrons connect atoms. We call these types of pairs bonds. Other pairs lay on atoms. We call these lone pairs. Each atom can have a different number of lone pairs. For example, in the Lewis structures below carbon has one lone pair whereas iodide has three pairs. Bonds can be simple or multiple, double or triple. Finally, some molecules are charged and the charge is normally indicated on the top right side of the representation.



Electron-dot structure of diatomic molecules Electron-dot structures–or Lewis structures–of diatomic molecules are the most simple electron-dot structures of molecules that you will see. To obtain these structures, you need to follow the next steps. The first step is (1) to set up the atoms in the molecule in the form of a line. After that, (2) you need to count the total number of valence electrons in the molecule by adding the valence electrons of each atom (remember the number of valence electrons corresponds to the group number in the A notation). Then (3) compute the pairs of electrons represented by lines–the total number of valence electrons divided by two. Finally, (4) you need to start distributing the electron pairs in the molecule in a very specific way, first connecting the atoms among themselves, and then placing the remaining pairs surrounding the atoms. Following the octet rule, each atom except for H and B should be surrounded by four pairs, counting as pairs the bonds and lone pairs.

Sample Problem 50

Construct the electron-dot structure of HCl.

SOLUTION

We first arrange the atoms in the molecule as indicated below and then we count the number of valence electrons: H(1) and Cl(7) that gives a total of eight electrons. We have four pairs of electrons.

$$\mathrm{H\ Cl}$$

Now we distribute the pair on each atoms knowing that each atom has to have four pairs with the exception of hydrogen that can only be surrounded by one pair. We can use lines instead of pairs

$$\mathrm{H}:\ddot{\underset{\cdot\cdot}{\mathrm{Cl}}}: \quad \text{or} \quad \mathrm{H}-\overline{\underline{\mathrm{Cl}}}|$$

STUDY CHECK

Construct the electron-dot structure of HF.

▸Answer: $H-\overline{\underline{F}}|$.

Electron-dot structure of general molecules Now we will address how to build up electron-dot structures of more complex molecules given that one of the atoms is the central atom and the others are connected to this central atom. The first step is (1) to arrange the atoms in the molecule, in the form of a central atom and the remaining atoms around it; the central atom is the one with a lower index in the molecule (e.g. in H_2O is O or in NH_3 is N). After that, (2) you need to count the total number of valence electrons in the molecule, dividing this number by two to obtain the number of pairs of electrons represented by lines. In the following (3) you need to connect the surrounding atoms to the central atom with electron pairs, and then (4) place electron pairs on top of the surrounding atoms, always placing a maximum of four atoms. Finally (5) place the remaining pairs in the central atom. Overall, each atom should be surrounded by four pairs (this is the octet rule) except H and B which should be surrounded by one and three pairs respectively. When drawing Lewis structures it is not important the atom arrangement (if the molecule looks like a line, a triangle or so) as long as the connectivity (which atom goes in the center and the surroundings) is correct.

Sample Problem 51

Construct the electron-dot structure of H_2O indicating the number of bonds and lone pairs.

SOLUTION

1. **Step one:** we first arrange the atoms in the molecule as H O H. The central atom is O as oxygen has the lower index in the H_2O molecule–the index for O is one and the index for H is two.
2. **Step two:** now we count the total number of valence electrons, including all atoms: 2xH(1) and O(6) that gives a total of eight electrons.
3. **Step three:** let us count the pairs of electrons; we have eight electrons and that is four pairs.
4. **Step four:** now we distribute the pair on each atoms knowing that each atom has to have four pairs with the exception of hydrogen that can only be surrounded by one pair. $H{:}\ddot{\underset{..}{O}}{:}H{:}$ and using lines instead of pairs (this is not necessary but makes the electron-dot structure look better) we obtain $H-\overline{\underline{O}}-H$. The molecule has two bonds, each one connecting and H to the oxygen atom and two lone pairs located on the oxygen atom.

STUDY CHECK

Construct the electron-dot structure of NH_3 indicating the number of bonds and lone pairs.

▸Answer: $H-\overset{}{\overline{N}}-H$ (with H bonded below N); three bonds and one lone pair.

Multiple bonds Often you are going to encounter electron-dot structures like the ones below

$$:N\equiv N: \text{ and } \ddot{:O}=\ddot{O:}$$

in which the atoms are connected through multiple bonds, double or triple bonds. Multiple bonds are formed while constructing electron-dot structures to impose the octet rule. Look for example the lewis structure for the HCN molecule below

$$H-C-\ddot{\underset{\cdot\cdot}{N}}:$$

In this structure, carbon does not follow the octet rule. We can enforce the octet rule by moving lone pairs from the atoms into the bond forming the structure below

$$H-C\equiv N:$$

In this structure both carbon and nitrogen follow the octet rule. Hence, we need to add one more step to the Lewis structure construction scheme: convert lone pairs of electrons into bonds to enforce the octet rule.

Sample Problem 52

Construct the electron-dot structure of O_2.

SOLUTION

1 **Step one:** We first arrange the atoms in the molecule as

$$O\ O$$

Now we count the total number of valence electrons, including all atoms: 2xO(6) that gives a total of twelve electrons. Let us count the pairs of electrons; we have twelve electrons and that is six pairs. Then we distribute the pairs, fist connecting the atoms $O-O$ (we have five extra pairs to distribute at this point), and we place the remaining pairs on top of the oxygen atoms

$$|\overline{\underline{O}}-O\rangle$$

The right oxygen do not follow the octet rule. In order to enforce the octet rule we move lone pairs into the bond

$$\langle O=O\rangle$$

STUDY CHECK

Construct the electron-dot structure of CO_2.

▶Answer: $\langle O=C=O\rangle$

Atomic charges in a molecule The electron-dot structure of a molecule results from counting the overall number of valence electrons of the molecule given that each atom brings a different number of valence electrons (n_e^{free}). For example, two H atoms bring one electron each, whereas O brings two electrons, giving a total of six valence electrons. When arranging the electron pairs in the molecule, the number of electrons surrounding an atom is called the valence of this atom in the molecule

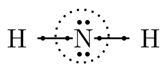
(n_e^{bonded}). We calculate the number of valence electrons of an atom in a molecule by accounting for the number of lone pairs on this atom and half the number of bonds:

$$n_e^{bonded} = \text{number of lone pairs} + 1/2\text{number of bonds}$$

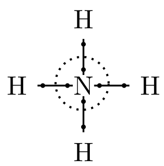
It is important to note that the valence of a free atom and the valence of this atom in a molecule is not necessarily the same. Indeed the difference between the valence electron of a free atom and the same atom in a molecule is the effective charge of that atom in the molecule, Q:

$$Q_{eff} = n_e^{free} - n_e^{bonded}$$

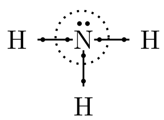
When the valence of an atom in a molecule is larger than the valence of the free atoms we have negative effective charges. In the example below



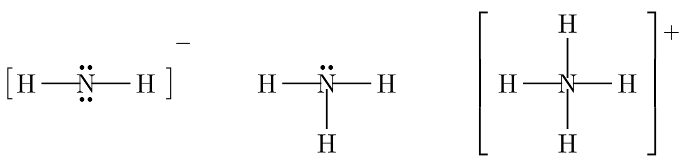
the number of electrons surrounding nitrogen is six electrons, more than the number of electrons originally brought to the molecule (five). We conclude that the atoms have a negative charge, and the effective atomic charge of nitrogen is Q=-1. In the next example,



the central atom, nitrogen, still has five valence electrons. After counting the electrons surrounding nitrogen, this time we find that this atom is surrounded by four electrons, less than the number of electrons originally brought to the molecule. We can conclude that nitrogen has a positive charge. In particular, the effective atomic charge of nitrogen in this molecule is the number of valence electrons minus the number of surrounding electrons. In this case, the atomic charge is Q=+1. When the valence of an atom in a molecule is the same as the valence of the free atoms we have zero effective charges. In this last example



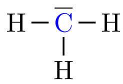
the central atom, nitrogen, has five valence electrons. After counting the electrons surrounding nitrogen–remember in a bond each atom shared an electron and hence each line around an atom counts as one electron–we find that this atom is surrounded by five electrons. As the number of valence electrons brought to the molecule is the same as the number of electrons surrounding the atom, we say the atomic charge of this atom is zero (Q=0). Hence, the atom is neutral. In all the molecules above, hydrogen remains neutral and hence the atomic charge of nitrogen corresponds to the molecular charge of the molecule. We can hence summarize the three scenarios indicated, as we have a neutral molecule in the center, a positive molecule on the right, and a negative molecule on the left.



The atomic charge of an isolated atom can be well-defined. However, the atomic charge of an atom in a molecule is arbitrarily defined, and more than one definition exists. Formal charges are one of the possible definitions of atomic charges in a molecule, whereas redox numbers are an alternative definition of atomic charges in molecules. None of these definitions is exactly correct. For example, redox numbers tend to overestimate the atomic charges, as they assume that all shared electrons in a bond belong to the most electronegative atom. Normally, negative formal charges tend to reside on electronegative atoms and not on electropositive atoms. At the same time, the sum of all effective charges needs to give the overall charge of the species. Furthermore, atoms in molecules tend to achieve formal charges as close to zero as possible. One can use formal charges to assess the validity of a Lewis structure. When comparing a series of equivalent Lewis structures for a molecule, the structures that best describe the bonding in the molecule tend to be those with small effective charges located on electronegative atoms.

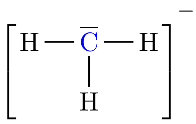
Sample Problem 53

Indicate the atomic charges of the blue highlighted atom



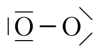
SOLUTION

The carbon atoms brings four electrons and in the molecule it is surrounded by eight electrons, five of which belongs to it. Hence the charge of C is -1; this means that carbon has one extra electron. Each hydrogen brings one electron and in the molecule each hydrogen has one electron (they share two electrons with C, one for C and one for H). The final lewis structure with the local charge of carbon can be indicated as:



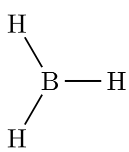
STUDY CHECK

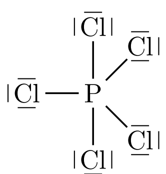
Indicate the atomic charges of all atoms in the Lewis structure below



▶Answer: left oxygen -1; right oxygen $+1$

Exceptions to the octet rule The octet rule assumes that all atoms are surrounded by eight pairs of electrons (either lone pairs or bonds) as atoms tend to achieve a stable noble gas configuration. However, not all atoms follow the octet rule. Is fair to say that most of the second-period atoms, C, N, O, and F follow the octet rule. Hydrogen does not follow the octet rule. It follows the duet rule of being surrounded by a single pair of electrons. Similarly, Boron does not follow the octet rule and tends to form a maximum of three bonds. The Lewis structure of a boron compound is presented below:





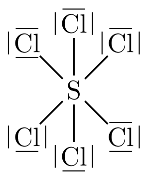
Boron and hydrogen have a reduced octet, with less than eight electrons. At the same time, atoms belonging to the third period (e.g. S, P) do not tend to follow the octet rule. They tend to have an expanded octet, surrounded by more than eight electrons. This is because the model behind the octet rule, the localized electron model, assumes that all atoms use their s and p orbitals to create bonds. However, atoms in the third period can also use the d orbitals with can fit extra pairs of electrons. The Lewis structure of a phosphorous compound is presented above. When dealing with expanded octets it is safe to assume that the central atom locates the expanded octet and hence the extra pairs of electrons.

Sample Problem 54

Obtain the Lewis structure of SCl_6.

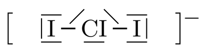
SOLUTION

We have that in the sulfur hexachloride molecule, sulfur is connected to six chlorine atoms. Hence, in this molecule, sulfur has an expanded octet. Let us first count the number of electrons and pairs of electrons. The valence of sulfur is six and the valence of chlorine is seven. Overall we have 48 electrons and 24 pairs. The Lewis structure of this molecule will be:



STUDY CHECK

Obtain the Lewis structure of $I_3{}^-$

▶Answer: 

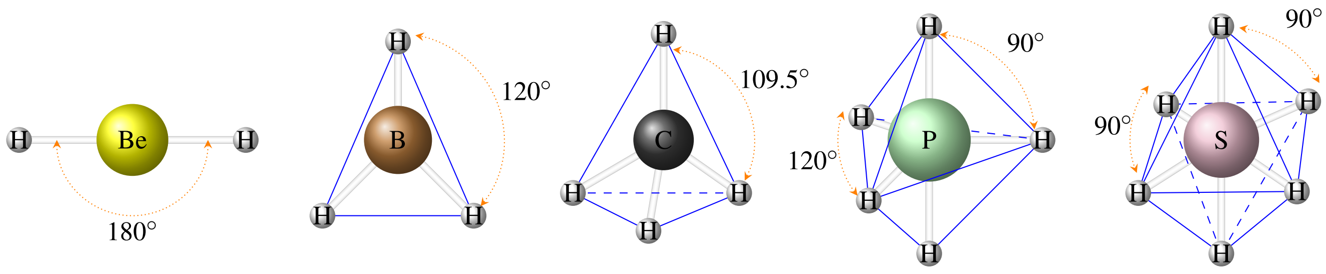
Steps to obtain Lewis structures The following steps can be used to obtain the Lewis structure of a general molecule:

1. **Step one:** Arrange the atoms in the molecule, in the form of a central atom and the surrounding atoms
2. **Step two:** Obtain the number of pairs of valence electrons by dividing the total number of valence electrons of the molecule by two
3. **Step three:** Connect the surrounding atoms to the central atom with electron pairs
4. **Step four:** Place electron pairs on top of the surrounding atoms, always placing a maximum of four atoms
5. **Step five:** Place the remaining pairs in the central atom.
6. **Step six:** Convert lone pairs of electrons into bonds to enforce the octet rule
7. **Step seven:** For extended octets place the extra electrons on the central atom
8. **Step seight:** When numerous equivalent Lewis structures exist, the best structures would have low formal charges, with negative charges located on electronegative atoms

5.8 Molecular shape

Molecules consist of arrangements of atoms presented in different forms. Let us use as an example the H_2O molecule, which contains two hydrogen atoms and one oxygen. Knowing that both hydrogens are connected to oxygen through a covalent bond, one can envision several molecular geometries such as a linear geometry or maybe a v-shaped geometry with oxygen at the point. The geometry of a molecule determines its properties, and small geometrical changes can have severe consequences on the functioning of molecules. For example, at high temperatures, when proteins in the body denaturalize losing their unique structure they also lose their functionality. The goal of this section is to identify the approximate shape of a given molecule.

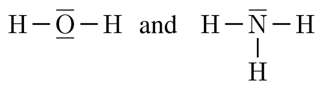
`The VSEP model` The VSEPR model, also known as the valence shell electron-pair repulsion model, is a model that predicts the geometries of molecules made of nonmetals. This model predicts the atomic arrangement of the molecules with an emphasis on the shape of the arrangement. However, it is not a very accurate model to predict geometries and there are better methods to obtain molecular geometries. The model predicts, for example, that water molecules have a v-shaped geometry and not a linear geometry while giving an estimate of the angle between the two O-H bonds. Still, VSEPR is not accurate enough to predict the O-H bond length and more advanced methods should be used for this purpose. The VSEPS model is based on the premise that the structure around a given atom results from minimizing the electron-pair repulsion. This way, the bonding and nonbonding pairs of electrons around a central atom are differently accounted for. Let us analyze a few cases in which the central atom only has bonding pairs of electrons. For example, the BeH_2 molecule has two bonding pairs around Be and the arrangement that maximizes the distance between both pairs hence minimizing repulsion is a linear arrangement. Hence, the BeH_2 molecule is linear with a 180° angle between both Be-H bonds. Another example would be the BH_3 molecules, a molecule with three bonding pairs. The geometry that maximizes the distance between the three pairs hence minimizing repulsion is a trigonal planar structure in which the three bonding pairs are in the same plane with an angle of 120° between the three bonds. A final example would be the methane molecule (CH_4), a molecule with four bonding pairs. A tetrahedral arrangement with 109.5° between the C-H bonds is the most stable arrangement for this case.



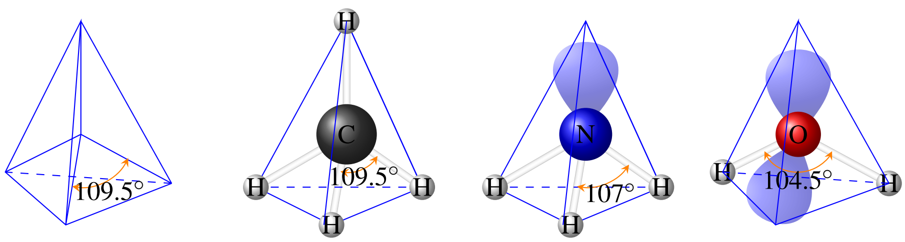
Linear | Planar trigonal | Tetrahedral | Trigonal bipyramidal | Octahedral

For the case of five bonds, the geometrical arrangement that minimizes the electron-pair interaction is a trigonal bipyramidal arrangement consisting of two pyramidal arrangements sharing a common base. The PH_5 molecule presents this arrangement. In this arrangement, there are two different bond angles: 90° between the vertical and in-plane and 120° for the in-plane bonds. Finally, the octahedral structure minimizes the pair repulsion in the case of six bonds and for example, the SH_6 molecule has an octahedral arrangement. In this arrangement, all bonds have a 90° angle. All atomic arrangements discussed above are presented in the diagram below.

`ABE Molecular code` We will use the ABE code to identify the molecular geometry of more complex molecules, with bonds and lone pairs. This code is based on the Lewis structure of the molecule, with B refers to the number of atoms connected to the central atom in the molecule (number of bonded atoms), and E is the number of lone pairs on the central atom. The overall number of bonded atoms and lone pairs is called the number of electron groups. Corresponding geometry for different ABE codes is tabulated. For example, an AB_2 molecule will be linear, whereas an AB_2E_2 is bent. The electron-dot structure of water and ammonia are:



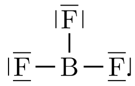
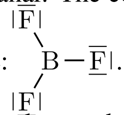
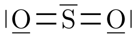
Water has two bonds with the central atom and hence two Bs and two lone pairs on top of the central atom and hence two Es. The ABE code of water is AB_2E_2 and its geometry is bent. The ABE code of ammonia is AB_3E, as the molecule has three atoms connected to the central nitrogen and N has a single lone pair. Its geometry would be trigonal pyramidal. Angles between the different bonds for the different atomic arrangements are also tabulated. For example, the angle between the two H-O bonds of water would be 104.5°, whereas the angle between two of the N-H bonds of ammonia would be 107°. The overall number of bonding and lone pairs is referred to as the number of electron regions and the molecular geometry of the molecule is not necessarily the geometry of the electron regions. For example, the molecule methane has four bonds and a tetrahedral geometry. Ammonia has two bonds and two lone pairs. The geometry of the electron regions is also tetrahedral with three bonds pointing toward the lower part of the tetrahedron and the lone pair pointing toward the upper part. At the same time, the molecular geometry of ammonia is trigonal pyramidal. For the case of water, we have that again the geometry of the four electron regions is tetrahedral whereas the molecular geometry is bent. We can also conclude that lone pairs require more room than bonding pairs and this has an impact on the molecular angles. For example, the angle between two bonds in a tetrahedron is 109.5° being this value is the same as the molecule angles of methane. Differently, the molecular angles of ammonia–a molecule with one lone pair–are 107°, and the molecular angle of water–a molecule with two lone pairs–is 104.5°. These results indicate that as the number of lone pairs increases the bonding pairs are more squeezed together.



Sample Problem 55

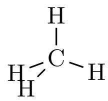
Identify the geometry of the following molecules: BF_3 and SO_2.

SOLUTION

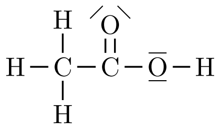
We need first the electron-dot structure of both molecules. For BF_3 
The code of this molecule is AB_3 and hence its geometry would be trigonal planar. The correct way to draw the molecule respecting its geometry would be: . The electron-dot structure for sulfur dioxide–remember this is covalent molecule–is  and its class is AB_2E. Hence the molecular geometry is linear.

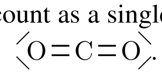
STUDY CHECK

Identify and draw the geometry of methane (CH_4).

▶Answer: Tetrahedral 

Complex molecules and multiple bonds To identify the ABE code we need to identify first the central and peripheral atoms. For some more complex molecules, there might not seem to be a central atom. An example of this case can be found in acetic acid.

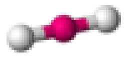
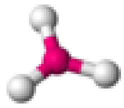
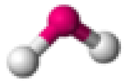
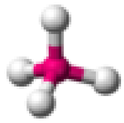
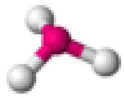
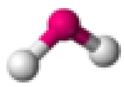
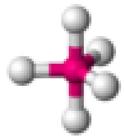
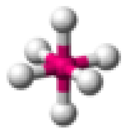
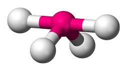
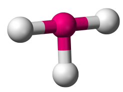
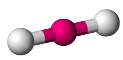
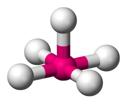
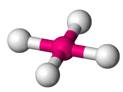


In the molecule we have a carbon atom connected to three hydrogens and to another carbon which at the same time is connected to two oxygens, one of these is also connected to the final hydrogen. For this case, we just need to identify different geometry centers, that is, atoms that are central to the other connections and which are connected to at least two atoms. For this molecule, we have three central atoms. For each central atom, there will be a molecular arrangement. For example, the ABE code for the left carbon is AB_4 whereas for the right carbon is AB_3. The geometry for the left-center would be tetrahedral, whereas for the right center would be planar trigonal. The right oxygen is also a central atom connected to carbon and hydrogen. The ABE code for this center is AB_2E_2 and the geometry is bent. Furthermore, in the ABE code double bonds count as a single bonding pair. For example, the Lewis structure of carbon dioxide is: . The ABE code for carbon dioxide would be AB_2 and its geometry would be linear with a 180° angle between the two C=O bonds. In the case of resonance, with several Lewis structures contributing to the chemical bond, any structure would predict the VSEPR geometry. Finally, it is important to stress that the VSPER model is just a simple model and we should be careful when predicting quantitative information. For example, we have that both molecules ammonia and phosphine, NH_3 and PH_3, both with code AB_3E have a trigonal pyramidal geometry. However, whereas the experimental bond angle of ammonia is 107°–close to the value predicted by the model–the bond angle for phosphine is 94°–very different than the angle predicted by the VSEPR model. Still, this model is good enough to predict general molecular structures.

Steps to use the VSEPR model The following steps can be used to obtain the molecular geometry using the VSEPR model:

1. **Step one:** Identify the central and the peripheral atoms.
2. **Step two:** Obtain the Lewis structure of the molecule
3. **Step three:**Obtain the ABE code with B being the number of peripheral atoms and E being the number of lone pairs. A represents the central atom. Multiple bonds (double, triple) count as a single B.
4. **Step four:** Use Table **??** to obtain the molecular geometry

Table ?? Molecular geometries

Electron groups (AEs)	Electron-group geometry	Bonded atoms (Bs)	Lone pairs (Es)	ABE Code	Molecular shape	Bond Angle	3D model
2	Linear	2	0	AB_2	Linear	180°	
3	Trigonal Planar	3	0	AB_3	Trigonal Planar	120°	
3	Trigonal Planar	2	1	AB_2E	Bent	120°	
4	Tetrahedral	4	0	AB_4	Tetrahedral	109°	
4	Tetrahedral	3	1	AB_3E	Trigonal pyramidal	109°	
4	Tetrahedral	2	2	AB_2E_2	Bent	109°	
5	trigonal bipyramidal	5	0	AB_5	trigonal bipyramidal	90°, 120°,180°	
6	octahedral	6	0	AB_6	octahedral	90°, 180°,180°	
5	trigonal bipyramidal	4	1	AB_4E	see-saw	180°,120°, 90°	
5	trigonal bipyramidal	3	2	AB_3E_2	T-shaped	90°, 180°	
5	trigonal bipyramidal	2	3	AB_2E_3	Linear	180°	
6	octahedral	5	1	AB_5E	square pyramidal	90°	
6	octahedral	4	2	AB_4E_2	square planar	90°, 180°	

5.9 Polarity of molecules

This section deals with bond and molecule polarity. A chemical bond will be polar or nonpolar depending on the tendency of the atoms in a bond to attract the electrons in the bond. Polar bonds result in the existence of a permanent dipole moment that makes a molecule polar. Polar molecules can interact with polar molecules and mix.

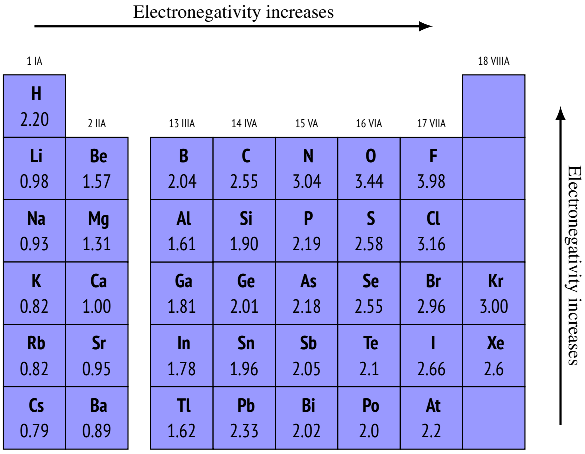
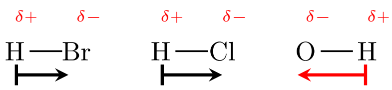


Figure 5.3 Electronegativity across the periodic table

`Bond polarity` Differences in electronegativity (see Figure 5.3) can be used to classify a chemical bond as covalent or ionic (see Table 5.4). On one hand, when the electronegativity difference (ΔEN) of the atoms in the bond is less than 0.4 we will say the bond is *nonpolar covalent*. Examples are the H–H (ΔEN=0) or C–H (ΔEN=0.2) bonds. In these cases, the electrons are shared equally in the bond leading to a lack of charge distribution. On the other hand, a bond is *polar covalent* when the atoms bonded are different or have an electronegativity difference between 0.4 and 1.8. Examples are H–Br (ΔEN=0.8), H–Cl (ΔEN=1.0), or O–H (ΔEN=1.2). Bonds become more polar as the electronegative difference in the bond increases. In polar covalent bonds, each atom shares the electrons in the bond unevenly which creates a *dipole moment*, a permanent charge separation. In the examples above, bromine or chlorine is more electronegative than hydrogen being more prone to attract the electrons than H. Br and Cl are partially negative $Cl^{\delta-}$ because of the negatively charged electrons and H is partially positive $H^{\delta+}$ because of the lack of electrons. We represent the excess of charge as on Cl or Br as $Cl^{\delta-}$ and electron deficiency in H as $H^{\delta+}$. The polarity of the bond is represented with an arrow pointing from the less electronegative, from positive, to the more electronegative atom so that the larger the electronegative difference the larger the dipole moment:



Finally, a bond is *ionic* when the atoms bonded have an electronegativity difference larger than 1.8. Examples are Na–Cl (ΔEN=2.2), K–Cl (ΔEN=2.3), or K–Br (ΔEN=2.1). For such a large electronegative difference, electrons are transferred from one atom to the other creating positive and negative ions in the bond. Ionic bonds are also polar due to the permanent charge distribution between the ions forming the bond. Mind that the variations in the type of bond are continuous and there is no defined point at which a bond stops being covalent to become ionic.

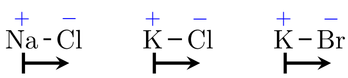


Table 5.4 Electronegative (EN) differences and types of bond

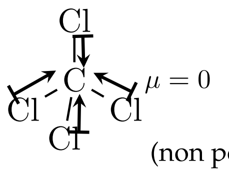
Electronegative Difference	0	0.4	1.8
Bond type	Nonpolar covalant	Polar covalent	Ionic
Electron Sharing	evenly	unevenly	electron transfer
	C − H	$\overset{\delta-}{\mathrm{O}}$ — $\overset{\delta+}{\mathrm{H}}$	$\overset{+}{\mathrm{Na}}$ - $\overset{-}{\mathrm{Cl}}$

`Polarity of molecules: diatomic molecules` Molecules can either be polar or nonpolar. Polar molecules have a permanent dipole moment resulting from one of more polar bonds. Nonpolar molecules have no permanent dipole moment resulting either from nonpolar bonds or from a combination of polar bonds that lead to no dipole moment. The polarity of diatomic molecules, small molecules with only two atoms, only depends on the nature of the atoms that form the molecule. If the atoms in the molecule are the same (e.g. H_2 or O_2) or have similar electronegativities, then the molecule would be nonpolar. If the atoms are different with electronegativity difference between 0.4 and 1.8 then the molecule would be polar. Examples are H_2 a nonpolar molecule whereas HCl or HBr are both polar molecules. See Table 5.5 for more examples.

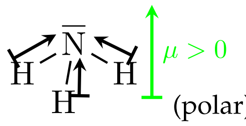
Table 5.5 Bond polarity differences in terms of electronegative (EN) differences

Bond	Electron Sharing	EN Difference	Type of bond
H − H	Evenly	0.0	Nonpolar covalent
C − H	Evenly	0.2	Nonpolar covalent
$\overset{\delta+}{\mathrm{H}}$ — $\overset{\delta-}{\mathrm{Br}}$	Unevenly	0.8	Polar covalent
$\overset{\delta+}{\mathrm{H}}$ — $\overset{\delta-}{\mathrm{Cl}}$	Unevenly	1.0	Polar covalent
$\overset{\delta+}{\mathrm{H}}$ — $\overset{\delta-}{\mathrm{O}}$	Unevenly	1.2	Polar covalent
$\overset{+}{\mathrm{K}}$ — $\overset{-}{\mathrm{Br}}$	Electron transfer	2.1	Ionic
$\overset{+}{\mathrm{Na}}$ — $\overset{-}{\mathrm{Cl}}$	Electron transfer	2.2	Ionic
$\overset{+}{\mathrm{K}}$ — $\overset{-}{\mathrm{Cl}}$	Electron transfer	2.3	Ionic

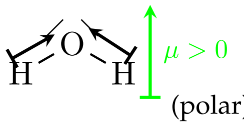
`Polarity of larger molecules` The polarity of larger molecules depends on the molecular geometry as well as the polar nature of the bonds that form the molecule. Let us analyze a few cases, H_2O, NH_3 and CCl_4. On one hand, for the H_2O case, the H−O bond is polar. The ABE type of water of AB_2E_2 and hence its geometry is bent; both H−O bonds do not compensate as they point in different directions and the directions do not cancel out what makes the water molecule polar having a permanent dipole moment ($\mu > 0$). Similarly, ammonia is a polar molecule made of polar bonds. The corresponding dipoles do not compensate each other ($\mu > 0$). On the other hand, CCl_4 is a nonpolar molecule made of polar bonds that do compensate each other ($\mu = 0$).



(non polar)

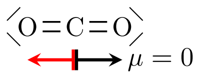


(polar)



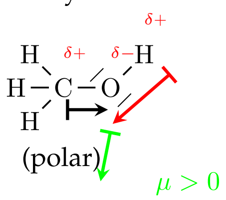
(polar)

On one hand, each of the C–O bonds on the carbon dioxide molecule is a polar bond. However, CO_2 is a linear molecule and the polarity of each C–O bonds compensate so that at the end the molecule is polar.



(non polar)

Finally, in some molecules, the dipoles of different bonds can have different directions, partially compensating for each other, as shown below for the case of methanol (CH_3OH). The O–H dipole is larger than the C–O dipole as hydrogen is less electronegative (EN=2.2) than carbon (EN=2.5). Therefore both dipole moments reduce giving an overall dipole. Note the C–O–H bonds should be arranged in a bent geometry.

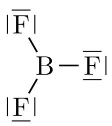


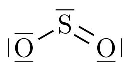
Sample Problem 56

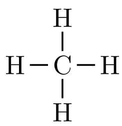
Identify the polar character (polar/nonpolar) of the following molecules: BF_3, SO_2, and CH_4.

SOLUTION

Let us analyze the geometries of the three molecules:







The bonds on SO_2 do not cancel out, as they do not point in opposite directions. Hence this molecule is polar. On the other hand, the bonds on methane and BF_3 cancel each other out, and hence even when the C-H and B-F bonds are polar, these two molecules would be non-polar.

STUDY CHECK

Identify the polar character (polar/nonpolar) of the following molecules: O_2 and NH_3.

▶Answer: O_2 is non-polar and NH_3 is polar.

`Polarity and mixing` Polarity is a key property when two liquids or two gases mix. If the mixing molecules have the same polar character they will be able to mix, whereas they will not mix when the polar character is different. For example, H_2O and HCl mix well as both are polar molecules. Differently, H_2O and CCl_4 do not mix well, as while water is a polar molecule, carbon tetrachloride (or chloro methane) is a nonpolar molecule. Differences in polarity create immiscibility in liquids and gases. Finally, methanol (CH_3OH) is a polar molecule as well. The central atom of a molecule (C) is connected to three hydrogens and an OH group. Hence this will be a polar molecule as one of the atoms attached to carbon is different. Both molecules, water, and methanol will mix as they have the same polarity. As a general rule: molecules with the same polarity (polar-polar or nonpolar-nonpolar) will mix.

`Molecules with the same polarity` Methanol (CH_3OH) is a polar molecule as well. The central atom of molecule (C) is connected to three hydrogens and an OH group. Hence this will be a polar molecule as one of the atoms attached to carbon is different. Both molecules, water, and methanol will mix as they have the same polarity. Methane (CH_4) is a nonpolar molecule, as the four polar C-H bonds compensate each other. Similarly, CCl_4, tetrachloro methene, is another nonpolar molecule, for the same reason. Both molecules, CH_4 and CCl_4 will mix together. As a general rule: molecules with the same polarity (polar-polar or nonpolar-nonpolar) will mix.

`Molecules with different polarity` CCl_4 is a nonpolar molecule, and H_2O is a polar molecule. As both have different polar characters they will not mix. If you mix water and CCl_4, two phases will remain instead of a single mixed liquid phase. As a general rule: molecules with different polarity (polar-nonpolar) will not mix. Another example will be water and oil. Water is polar, and oil is a nonpolar molecule. As a consequence, these two molecules will not mix. Soap has a polar and nonpolar parts. To remove oil from water, soap helps mix both polar water and nonpolar oil.

5.10 Intermolecular forces

Atoms in liquid or solid compounds are connected through chemical bonds, and bonds are forces within molecules. These bonds can be ionic or covalent depending on the nature of the elements that form the molecule. At the same time, the molecules of a liquid or solid compound interact with each other through intermolecular forces. The word intermolecular means between molecules. This section describes the three existing types of intermolecular forces as well as their nature and intensity.

`Intermolecular forces and intramolecular interactions`
Molecules are made of atoms that connect through *intra*molecular interactions such as covalent or ionic bonds. Differently, molecules interact with each other through *inter*molecular interactions. The prefix *inter* means "occurring between", whereas the prefix *intra* means "occurring within". Intermolecular forces are responsible for the melting and boiling point of a chemical. On one hand, the stronger the forces the higher the melting and boiling point. On the other hand, the more intermolecular interactions the higher the melting and boiling point. This is because to melt or boil a chemical we need to overcome the intermolecular forces that connect molecules to release the into a different state of matter. In the following, we will describe the three main types of intermolecular forces (some books describe four types of intermolecular forces counting the ion-molecule interaction).

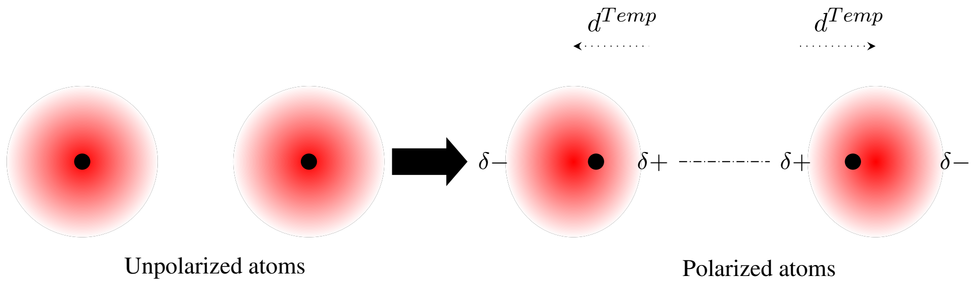


Figure 5.6 Dispersion forces result from instantaneous dipole moments resulting from the polarization of the electron density of atoms and molecules.

`Dispersion forces` All molecules are made of atoms that contain electrons. The electron density of an atom is distributed homogeneously without uneven charge distributions. As a consequence, atoms in general have no permanent dipole moment without negatively or positively charged regions. Still, when two atoms get close together, the presence of each other affects their electron density creating temporary dipole moments. We call this effect polarizability. This temporary dipole is responsible for London dispersion forces, also called Van der Waals forces or simply dispersion forces. Dispersion forces exist in all chemicals, as all chemicals can be polarized. The larger the atomic number, or the molar weight of the compound, the stronger these forces are. This is because in general the larger the atomic weight the more polarizable atoms are, and hence, they tend to generate stronger temporary dipoles, produced from charge polarization.

▼Polymer molecules interact by means of dispersion

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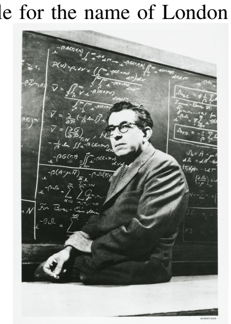
▼Fritz London a german physicist is responsible for the name of London forces



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▼geckos stick because of the van der Waals force

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The melting (or freezing) and boiling points of the noble gases are given in the Table below, where you can see how the larger the atomic mass of the gas the higher the melting–and boiling points. Mind that normally, the melting and freezing points of a substance are the same. Dispersion forces are common in chemicals made of hydrogen and carbon–we call these compounds hydrocarbons. The larger the size of the molecule the larger the effects of dispersion forces. For example:

C_6H_{14} (BP=69°C) C_8H_{18}(BP=125°C) $C_{10}H_{22}$(BP=174°C)

Table 5.4 Freezing and boiling point of the noble gases

Gas	Atomic Weight (amu)	Melting Point (°C)	Boiling Point (°C)
He	2	–	4
Ne	10	25	27
Ar	18	84	87
Kr	36	116	121
Xe	54	162	167
Rn	6	202	211

`Dipole-Dipole forces` Atoms have different electronegativity that is, a different tendency to attract the electron density in a bond. In the periodic table, electronegativity increases going from left to right and top to bottom. Elements in the top right of the table (Cl, F) tend to be very electronegative and hence they tend to strongly attract the electrons on chemical bonds. Differently, elements on the bottom left part of the table are electropositive and they tend to give away the electrons in the bond. Dipole moments result from differences in electronegativity. When an electronegative atom is connected to an electropositive atom in a bond, the electronegativity difference creates permanent dipole moments, and molecules with permanent dipole moments are called polar molecules. The dipole moment of a bond is a vector that points from the most electropositive atom to the most electronegative atom. For example for the O-H and N-H bonds,

$$\overset{\delta-}{\mathrm{N}}-\overset{\delta+}{\mathrm{H}} \qquad \overset{\delta+}{\mathrm{H}}-\overset{\delta-}{\mathrm{O}}$$

As such, the molecule HCl would be polar as a result of the combination of an electropositive atom (H) and an electronegative atom (Cl). Similarly, HF would be a polar molecule too. Dipole-dipole forces exist only in polar compounds, being the result of permanent dipole moments. These types of interactions are stronger than dispersion forces but weaker than normal interatomic covalent bonds. Molecules with dipole moment can attract each other through dipole forces, orienting themselves so that their positive side aligns with the negative side maximizing the electrostatic attraction. At the same type, dipole forces depend on the distance and at large distances are less effective. The dipole moment of a molecule is measured in Debye (D). For example, the dipole moment of HCl is 1.05D, whereas the dipole moment of HF is 1.82D. The stronger the dipole moment the stronger the dipole-dipole interactions. For example, when comparing C_3H_8 and CH_3OCH_3, the former has an almost null dipole moment, whereas the latter has a dipole moment of 1.3D. The boiling point of C_3H_8 is -42°C whereas the boiling point of CH_3OCH_3 is -25°C. The table below lists some dipole moments and boiling points showing the trend that the larger the dipole moment the higher (more positive) the boiling point.

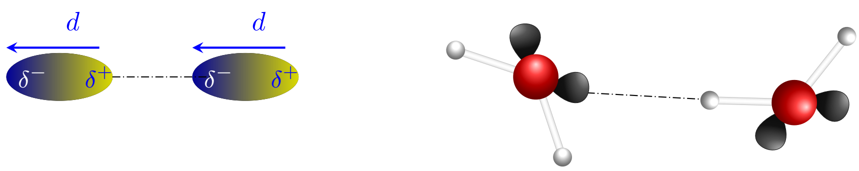


Figure 5.7 (Left) Dipole-dipole forces result from the interaction of permanent dipole moments existing in polar molecules. (Right) two water molecules interacting by means of hydrogen bonds

Table 5.5 Boiling (BP) point of a series of hydroacids

Compound	Dipole moment (Debyes)	Boiling Point (°C)
C_3H_8	0.1	-42
CH_3OCH_3	1.3	-25
CH_3Cl	2.0	-24
CH_3COH	2.7	-21
CH_3CN	3.9	-82

Hydrogen bonds Hydrogen bonds are the strongest of all intermolecular forces and exist only in molecules containing very specific bonds; in particular, they only exist in molecules containing H−F, H−N, or H−O bonds. An example of a molecule with hydrogen bonds is HF or NH_3. Hydrogen bonds are a specific type of dipole-dipole interaction responsible, among others, for some of the high boiling points of water. Due to the existence of hydrogen bonds water is liquid at room temperature (H_2O, BP=100°C), in comparison with similar molecules (H_2S, BP=-60°C). The anomalous character of oxygen, fluorine, and nitrogen results from the fact that these are very electronegative elements with lone pairs of electrons–these are non-bonding pairs of electrons–that enable the creation of hydrogen bonds. In particular, oxygen has two lone pairs, nitrogen one and fluorine three. The combination of high electronegativity, the presence of lone pairs, and the presence of hydrogen atoms makes hydrogen bonds possible.

Sample Problem 57

Indicate what types of intermolecular forces exist in the following molecules:

	HCl	CH_4	H_2O	CH_3Cl
Dispersion				
Dipole-Dipole				
H-bonds				

SOLUTION

All molecules can interact by means of dispersion forces. Differently, only polar molecules can interact by means of dipole-dipole forces. Finally, only molecules with a H−F, H−N or H−O bond can interact by means of hydrogen bonds. For these reason, from the table only HCl, H_2O and CH_3Cl has dipole forces, and only H_2O has hydrogen bonds.

	HCl	CH_4	H_2O	CH_3Cl
Dispersion	✓	✓	✓	✓
Dipole-Dipole	✓	✗	✓	✓
H-bonds	✗	✗	✓	✗

STUDY CHECK

Indicate what types of intermolecular forces exist in the following molecules: NH_3, HF, and CH_3-CH_3.

Intermolecular forces of liquids and boiling Boiling a liquid requires energy. This energy is invested in separating the molecules from the liquid until they are spread apart. To separate the molecules of a liquid, we need to overcome intermolecular forces. Imagine boiling CH_4. We know the molecules of methane only interact among themselves through weak dispersion forces. Imagine now boiling water. Water is polar and water has O−H bonds, hence water molecules interact through dispersion, dipole-dipole, and hydrogen bonds. The energy needed to separate the molecules of water will be larger than the energy required to separate the molecules of methane. The more intense the intermolecular forces, the higher the boiling point. Also, the more types of intermolecular forces present in a liquid the higher the boiling point. Finally, we can apply these ideas not only to liquids but also to solids.

Sample Problem 58

Compare the boiling point of these two molecules: HCl and H_2O.

SOLUTION

Let us build a table with the different types of intermolecular forces present in each liquid. The molecules of both liquids can interact by means of dispersion forces and also dipole-dipole forces, as both are polar molecules. Differently, only molecules with a H−F, H−N or H−O bond can interact by means of hydrogen bonds. For these reason, H_2O liquid contains hydrogen bonds.

	HCl	H_2O
Dispersion	✓	✓
D-D	✓	✓
H-bonds	✗	✓

Hence, water will boil at a higher temperature.

STUDY CHECK

Compare the boiling point of these two molecules: CH_3F and CH_4.

CHAPTER 5

Ions & Ionic charges

5.1 Indicate if the following chemical species represent an atom, and anion or a cation: (a) Fe^{2+} (b) Cl^- (c) Ag

5.2 Indicate if the following chemical species represent an atom, and anion or a cation: (a) Cs (b) Cs^+ (c) N^{-3}

5.3 Identify the ionic state of the following elements. If needed, indicate the existence of multiple ionic states: (a) H (b) O (c) N (d) F (e) Mn

5.4 Identify the ionic state of the following elements. If needed, indicate the existence of multiple ionic states: (a) Li (b) V (c) Cl (d) S (e) Cr (f) Sr (g) Ni

Covalent compounds

5.5 Name or formulate the following covalent compounds: (a) NO (b) Dichlorine monofluoride (c) NO_2

5.6 Name or formulate the following covalent compounds: (a) Chlorine Monofluoride (b) N_2O (c) Nitrogen trifluoride

5.7 Name or formulate the following covalent compounds: (a) SO_3 (b) Disulfur dichloride (c) SO_2 (d) Disulfur tetrachloride

5.8 Name or formulate the following covalent compounds: (a) P_4S_3 (b) Sulfur Tetrafluoride (c) As_2O_5 (d) Sulfur trioxide

Ionic compounds

5.9 Classify the following chemicals in two groups, justifying your classification: (a) NaCl (b) CO_2 (c) $FeCl_3$ (d) N_2O_3 (e) SO_3 (f) Ca_3N_2

5.10 Combine the following ions:

(a) $Na^+ + Cl^-$
(b) $Na^+ + Se^{2-}$
(c) $Na^+ + P^{3-}$
(d) $Mg^{2+} + Cl^-$
(e) $Mg^{2+} + O^{2-}$
(f) $Mg^{2+} + N^{3-}$

5.11 Name or formulate the following ionic (Type I) compounds: (a) Magnesium iodide (b) Ca_3P_2 (c) Lithium nitride (d) MgF

5.12 Name or formulate the following ionic (Type I) compounds: (a) Magnesium fluoride (b) CaS (c) Barium phosphide (d) Mg_3N_2

5.13 Name the following compounds:

(a) NaCl
(b) Ca_3N_2
(c) MgI_2
(d) SrS
(e) RbCl
(f) KF

5.14 Combine the following ions:

(a) $Cs^+ + F^-$
(b) $Sr^{2+} + O^{2-}$
(c) $Be^{2+} + C^{4-}$
(d) $Li^+ + I^-$

5.15 Classify the following chemicals in two groups. Justify your classification.

(a) NaCl
(b) MnO_2
(c) $FeCl_3$
(d) SrO
(e) Li_3N
(f) NiO

5.16 Formulate the following compounds:

(a) Copper(I) oxide
(b) Copper(II) nitride
(c) Nickel(III) oxide
(d) Manganese(IV) oxide

5.17 Name the following compounds:

(a) NiO
(b) Cr_2O_3
(c) VO
(d) MnO_4

5.18 Formulate the following compounds:

(a) Iron(II) nitride
(b) Copper(I) sulfide
(c) Chromium(III) iodide
(d) Palladium(IV) phosphide
(e) Manganese(VI) oxide

5.19 Name the following compounds:

(a) Ni_2O_3
(b) Fe_3N_2
(c) Cr_2O_3
(d) Ni_3P_2
(e) Ru_2Se_3

5.20 Name the following compounds:

(a) FeO
(b) CrN
(c) ZnI_2
(d) CoS
(e) MnF_3
(f) Cu_2C
(g) Ag_2O

5.21 Name or formulate the following ionic (Type II) compounds: (a) Fe_3P_2 (b) Copper(II) iodide (c) Fe_3N_2 (d) Iron(II) sulfide

5.22 Name or formulate the following ionic (Type II) compounds: (a) Fe_2S_3 (b) Gold(III) chloride (c) FeO (d) Vanadium(V) nitride

5.23 Name or formulate the following ionic (Type II) compounds: (a) FeI_2 (b) Lead(IV) sulfide (c) $FeBr_2$

5.24 Name or formulate the following ionic (Type II) compounds: (a) Manganese(IV) oxide (b) $FeCl_2$ (c) Copper(I) oxide

Acids and hydroxides

5.25 Name or formulate the following acids or bases: (a) HCl (b) Hydrofluoric Acid (c) $Mg(OH)_2$

5.26 Name or formulate the following acids or bases: (a) Sulfuric Acid (b) H_2CO_3 (c) Lithium hydroxide

5.27 From the following chemicals identify acids and bases: (a) KOH (b) $LiOH$ (c) CH_3OH

5.28 From the following chemicals identify acids and bases: (a) H_2SO_3 (b) NH_3 (c) $Ca(OH)_2$

5.29 From the following chemicals identify hydracids and oxoacids: (a) HF (b) H_2SO_3 (c) H_2S

5.30 From the following chemicals identify hydracids and oxoacids: (a) H_3BO_3 (b) HCl (c) HI

5.31 Identify the redox number of the central atom of the following oxoacids: (a) H_2CrO_4 (b) $H_2Cr_2O_7$ (c) $HMnO_4$

5.32 Identify the redox number of the central atom of the following oxoacids: (a) H_2MnO_4 (b) $HReO_3$ (c) H_2SiO_3

5.33 Identify the most oxidated acid:

(a) H_3AsO_4 or H_3AsO_3
(b) H_2XeO_4 or H_4XeO_6

5.34 Identify the most reduced acid:

(a) H_2RuO_4 or $HRuO_4$
(b) $HTcO_4$ or H_2TcO_4

5.35 Identify the most oxidant acid:

(a) H_2CrO_4 or $H_2Cr_2O_7$
(b) HNO_3 or HNO_4

5.36 Identify the most oxidant acid:

(a) $H_2S_2O_6$ or H_2SO_4
(b) H_2SeO_4 or H_2SeO_3

Naming of complex salts and common chemicals

5.37 Name or formulate the following oxoanions: (a) $ClO_4{}^-$ (b) $PO_4{}^{3-}$ (c) $SO_4{}^{2-}$ (d) $CO_3{}^{2-}$ (e) $NO_3{}^-$ (f) $CrO_4{}^{2-}$ (g) $BO_3{}^{3-}$

5.38 Name or formulate the following (Type I) oxosalts: (a) $Mg(NO_3)_2$ (b) Sodium permanganate (c) $KMnO_4$ (d) Calcium carbonate (e) Li_3PO_4

5.39 Name or formulate the following (Type I) oxosalts: (a) Lithium sulfate (b) Na_2CrO_4 (c) Lithium sulfite (d) $Cs_2Cr_2O_7$ (e) Calcium sulfate

5.40 Name or formulate the following compounds: (a) Na_2SO_4 (b) KNO_3 (c) $CaCO_3$ (d) $Ca(NO_2)_2$ (e) $SrSO_3$

5.41 Combine the following ions:

(a) $Na^+ + PO_4{}^{3-}$
(b) $Li^+ + MnO_4{}^-$
(c) $Mg^{2+} + NO_3{}^-$
(d) $Ca^{2+} + CO_3{}^{2-}$
(e) $Cs^+ + Cr_2O_7{}^{2-}$
(f) $K^+ + BO_3{}^{3-}$

5.42 Combine the following polyatomic ions:

(a) $Na^+ + NO_3^-$

(b) $Na^+ + CO_3^{2-}$

(c) $Na^+ + PO_4^{3-}$

(d) $Ca^{2+} + CO_3^{2-}$

(e) $Ca^{2+} + PO_4^{3-}$

5.43 Name or formulate the following (Type II) oxosalts: (a) $Cr_2(SO_4)_3$ (b) zinc(II) carbonate (c) $Fe(MnO_4)_3$

5.44 Name or formulate the following (Type II) oxosalts: (a) cobalt(III) carbonate (b) $Fe(ClO_4)_3$ (c) zinc(II) carbonate

5.45 Name or formulate the following hydrosalts: (a) $NaHCO_3$ (b) Calcium Hydrogencarbonate (c) $Al(HSO_4)_3$

5.46 Name or formulate the following hydrosalts: (a) Sodium dihydrogenphosphate (b) LiH_2PO_4 (c) Silver monohydrogenphosphate

5.47 Name or formulate the following hydrates: (a) $Al_2(SO_4)_3 \cdot 3\,H_2O$ (b) Silver phosphate dihydrate

5.48 Name or formulate the following hydrates: (a) $KMnO_4 \cdot 4\,H_2O$ (b) Lithium sulfate tetrahydrate

5.49 Name or formulate the following compounds: (a) $MgSO_4$ (b) $Ni(SO_4)_3$ (c) Cobalt(II) nitrate (d) Cobalt(II) sulfate dihydrate (e) $KHCO_3$

5.50 Name or formulate the following compounds: (a) $Ca(NO_3)_2$ (b) $Ca(HCO_3)_2$ (c) Nickel(II) sulfate (d) Nicke(II) sulfate tetrahydrate (e) NaH_2PO_4

5.51 Name or formulate the following compounds: (a) $MnSO_4$ (b) $CuNO_3$ (c) $Cr_2(CO_3)_3$ (d) $V(NO_2)_2$ (e) $FeSO_3$

5.52 Name or formulate the following pairs or ions: (a) carbonate and monohydrogencarbonate (b) sulfate and monohydrogensulfate (c) cromate and monohydrogenchromate (d) phosphate and dihydrogenphosphate (e) phosphate and monohydrogenphosphate (f) borate and dihydrogenphosborate

Electron-Dot structures of molecules

5.53 Draw electron-dot structures for the following diatomic molecules that obey the octet rule: (a) F_2 (b) Cl_2

5.54 Draw electron-dot structures for the following diatomic molecules that obey the octet rule: (a) HF (b) HCl

5.55 Draw electron-dot structures for the following diatomic molecules that obey the octet rule: (a) ICl (b) HI

5.56 Draw electron-dot structures for the following diatomic molecules: (a) CO (b) N_2 (c) O_2

5.57 Draw electron-dot structures for the following molecules that obey the octet rule, given that the first atom listed is the central atom: (a) CHN (b) CO_2

5.58 Draw electron-dot structures for the following molecules that obey the octet rule, given that the first atom listed is the central atom: (a) CH_4 (b) CH_3Cl (c) OH_2

5.59 Draw electron-dot structures for the following molecules that obey the octet rule, given that the first atom listed is the central atom: (a) NH_3 (b) NCl_3

5.60 Draw electron-dot structures for the following molecules that obey the octet rule, given that the first atom listed is the central atom: (a) $SeCl_2$ (b) CH_2O

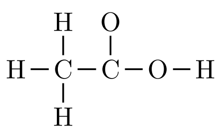
5.61 Draw electron-dot structures for the following molecules given that the first atom listed is the central atom. Some of the atoms might not obey the octet rule. If the species has a charge indicate the location of the charge: (a) BH_3 (b) BH_2F (c) $POCl_3$ (d) $ClO_4{}^-$

5.62 Draw electron-dot structures for the following molecules given that the first atom listed is the central atom. Some of the atoms might not obey the octet rule. If the species has a charge indicate the location of the charge: (a) BeH_2 (b) PCl_5 (c) SF_4 (d) ClF_3

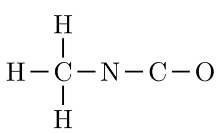
5.63 Draw electron-dot structures for the following molecules given that the first atom listed is the central atom. Some of the atoms might not obey the octet rule. If the species has a charge indicate the location of the charge: (a) $I_3{}^-$ (b) $Br_3{}^-$ (c) SF_6

5.64 Draw electron-dot structures for the following molecules given that the first atom listed is the central atom. Some of the atoms might not obey the octet rule. If the species has a charge indicate the location of the charge: (a) ClF_5 (b) XeF_4

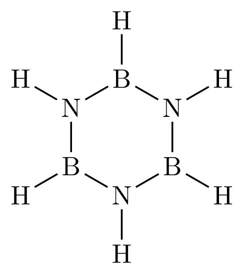
5.65 Given the skeletal structure below, draw the lewis structure of the molecule:



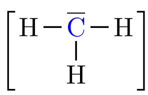
5.66 Given the skeletal structure below, draw the lewis structure of the molecule:

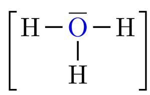


5.67 Given the skeletal structure below, draw the lewis structure of the molecule:

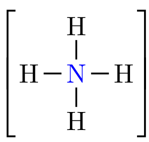


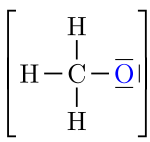
5.68 Indicate the charge of the atom marked blue in the following electron-dot structure:

(a) 

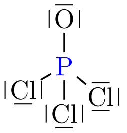
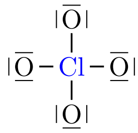
(b) 

5.69 Indicate the charge of the atom marked blue in the following electron-dot structure:

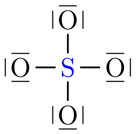
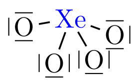
(a) 

(b) 

5.70 Indicate the charge of the atom marked blue in the following electron-dot structure that follow the octet rule:

(a)  (b) 

5.71 Indicate the charge of the atom marked blue in the following electron-dot structure that follow the octet rule:

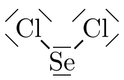
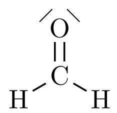
(a)  (b) 

Molecular shape

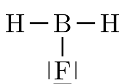
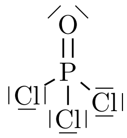
5.72 Identify the molecular shape of the molecules: (a) NH_3 (b) CH_4

5.73 Identify the molecular shape of the molecules: (a) H_2 (b) $BeCl_2$ (c) BF_3

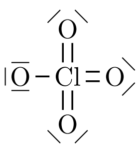
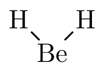
5.74 Given the following Lewis structures, predict the molecular geometry and angles:

(a)  (b) 

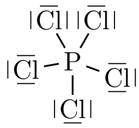
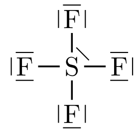
5.75 Given the following Lewis structures, predict the molecular geometry and angles:

(a)  (b) 

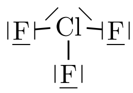
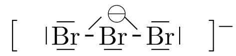
5.76 Given the following Lewis structures, predict the molecular geometry and angles:

(a)  (b) 

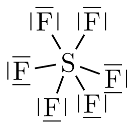
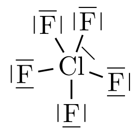
5.77 Given the following Lewis structures, predict the molecular geometry and angles:

(a)  (b) 

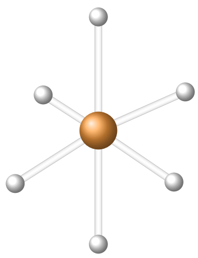
5.78 Given the following Lewis structures, predict the molecular geometry and angles:

(a)  (b) 

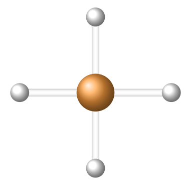
5.79 Given the following Lewis structures, predict the molecular geometry and angles:

(a)  (b) 

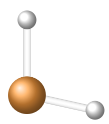
5.80 Identify the name of the following molecular structure:



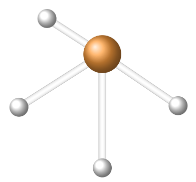
5.81 Identify the name of the following molecular structure:



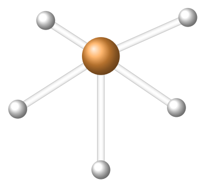
5.82 Identify the name of the following molecular structure:



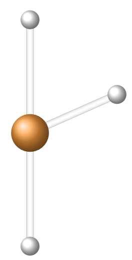
5.83 Identify the name of the following molecular structure:



5.84 Identify the name of the following molecular structure:



5.85 Identify the name of the following molecular structure:



POLARITY

5.86 Describe the trend in electronegativity (increases, decreases): (a) from Na to Mg (b) from Cs to Ba (c) from S to O (d) from P to C (e) from K to Rb (f) from B to C (g) from Li to Na

5.87 Describe the trend in electronegativity (increases, decreases): (a) from S to Cl (b) from Se to Br (c) from Rb to Cs (d) from P to O (e) from Cl to F (f) from Li to Be (g) from Se to Cl

5.88 Arrange the atoms in the following set in order of increasing electronegativity: (a) Li, Be, and B (b) Rb, Na, and K (c) N, C, and B

5.89 Arrange the atoms in the following set in order of increasing electronegativity: (a) P, Cl, and S (b) Br, F, and Cl (c) S, O, and Se

5.90 Classify the following bonds as ionic, polar colavent and nonpolar covalent by calculating the electronegative difference: (a) S–P (b) O–Mg (c) H–Na (d) H–N (e) Na–Cl (f) S–C

5.91 Classify the following bonds as ionic, polar colavent and nonpolar covalent by calculating the electronegative difference: (a) H–K (b) H–Li (c) O–Li (d) O–Na (e) Na–F (f) S–H

5.92 For the following ionic bonds indicate the positive and negative end of the dipole and whether the resulting dipole points to the left or to the right: (a) F — K (b) Se — Rb (c) Na — Cl

5.93 For the following ionic bonds indicate the positive and negative end of the dipole and whether the resulting dipole points to the left or to the right: (a) S — Ca (b) K — Br (c) Mg — S

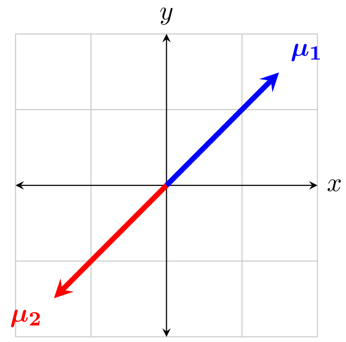
5.94 For the following covalent bonds indicate the partially positive and negative end of the dipole and whether the resulting dipole points to the left or to the right: (a) $\overset{\delta}{\mathrm{C}}-\mathrm{F}$ (b) $\overset{\delta}{\mathrm{N}}-\mathrm{F}$ (c) $\overset{\delta}{\mathrm{Si}}-\mathrm{O}$ (d) $\overset{\delta}{\mathrm{Br}}-\mathrm{S}$

5.95 For the following covalent bonds indicate the partially positive and negative end of the dipole and whether the resulting dipole points to the left or to the right: (a) N — O (b) S — C (c) N — C (d) Si — B

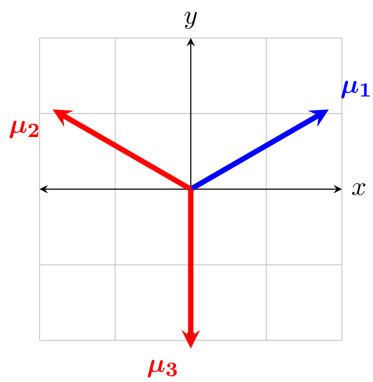
5.96 Select the more polar bond from the following set of bond pairs: (a) N–O or N–P (b) P–Si or P–S (c) S–Cl or S–Se

5.97 Select the more polar bond from the following set of bond pairs: (a) H–O or H–F (b) O–S or S–Se (c) H–Cl or H–F

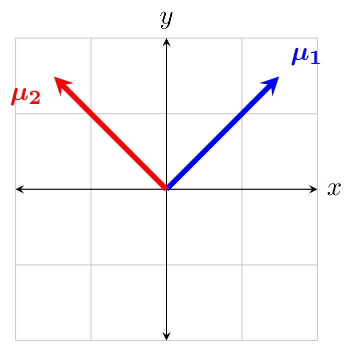
5.98 Predict if the following dipole distribution will lead to an overall dipole moment:



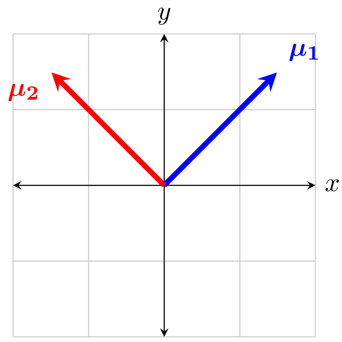
5.99 Predict if the following dipole distribution will lead to an overall dipole moment:



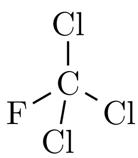
5.100 Predict if the following dipole distribution will lead to an overall dipole moment:



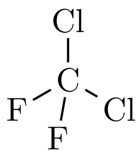
5.101 Predict if the following dipole distribution will lead to an overall dipole moment:



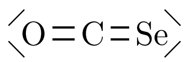
5.102 Predict if the following bond arrangement will lead to an overall dipole moment:



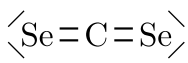
5.103 Predict if the following bond arrangement will lead to an overall dipole moment:



5.104 Predict if the following bond arrangement will lead to an overall dipole moment:



5.105 Predict if the following bond arrangement will lead to an overall dipole moment:



5.106 Indicate the polar character (polar, nonpolar) of the following molecules: (a) H_2O (b) HCl (c) CO_2 (d) CO (e) PF_3 (f) $SiCl_4$

5.107 Indicate the polar character (polar, nonpolar) of the following molecules: (a) CF_4 (b) NH_3 (c) SeF_2 (d) O_2(e) CH_4 (f) H_2

INTERMOLECULAR FORCES

5.108 Indicate the strongest intermolecular force existing between the molecules of the following compounds: (a) CH_3OH (b) H_2 (c) CCl_4

5.109 Indicate the strongest intermolecular force existing between the molecules of the following compounds: (a) CH_4 (b) CCl_3H (c) HF (d) HCl

5.110 From the following pair of molecules, which molecule forms intermolecular H bonds? (a) HF or H_2 (b) NH_3 or CH_4

5.111 From the following pair of molecules, which molecule forms intermolecular H bonds? (a) CH_3-O-CH_3 or H_2O (b) HCl or HF

5.112 From the following pair of molecules, which molecule forms stronger dispersion forces? (a) Ar or He (b) H_2O or H_2S

5.113 From the following pair of molecules, which molecule forms stronger dispersion forces? (a) CH_3CH_3 or CH_4 (b) CH_4 or CH_3Cl

5.114 From the following pair of molecules, which molecule forms stronger dipole forces? (a) HCl or HBr (b) H_2O or H_2S

5.115 From the following pair of molecules, which molecule forms stronger dipole forces? (a) NH_3 or H_2O (b) HI or HBr

5.116 From the following pair of molecules, which molecule has higher boiling point? (a) CH_3CH_3 or CH_4 (b) CO_2 or H_2O

5.117 From the following pair of molecules, which molecule has higher boiling point? (a) HF or HCl (b) Ar or He

Ch. 6. The Mole and Chemical Reactions

WHEN we buy eggs in the store, we buy them by the dozen, and the word dozen actually refers to the number twelve. Similarly, when we measure substances in a chemistry lab we measure them by the mole. This chapter will introduce the idea of mole and you will learn how to relate moles of a chemical to mass using a property called the molecular mass. This chapter also introduces chemical reactions. Chemicals react with each others and a chemical reaction is written in the form an equations. In this chapter you will learn how to balance those equations in order to predict the amount of chemicals produced.

6.1 The mole

Some of the terms you use in your everyday life refer to a number as shown in Figure 6.1. For example, you buy a pair of socks–two socks–or you buy a dozen of eggs from the grocery store–twelve eggs–and sometimes you buy a case of beer–24 cans. In a chemistry laboratory, we normally do not weigh small numbers of molecules of a chemical. In chemistry, molecules are counted by the mole, and the term mole–abbreviated as mol–refers to the 6.022×10^{23} number. For example, a mol of CO molecules contains 6.022×10^{23} molecules of CO, and a mol of water molecules contains 6.022×10^{23} molecules of water. This is because the word mole means the number 6.022×10^{23}, similarly as the word pair means the number two. The number 6.022×10^{23} is called Avogadro's number, about Amedeo Avogadro, the Italian physicist who coined the term. In chemistry labs, chemicals are often measured by weight. In this section, we will show how to convert moles into weight–into grams–by using a property called molecular weight. Finally, mind that the term mol and molecule even if they look similar they are not. Molecule refers to a combination of atoms and mole refers to a large number of molecules. As a note, the abbreviation of the mole is mol, and for example, we will say seven mol of H_2O.

From moles to molecules One mole of molecules contains 6.022×10^{23} molecules. This is because the term mole refers to Avogadro's number. Hence we can use the following unit equivalency:

$$\boxed{1\text{mol of } H_2O = 6.02 \times 10^{23}\text{molecules of } H_2O}$$

or a conversion factor to transform moles into molecules or molecules into moles as well:

$$\boxed{\frac{1\text{mol of } H_2O}{6.02 \times 10^{23}\text{molecules of } H_2O} \text{ or } \frac{6.02 \times 10^{23}\text{molecules of } H_2O}{1\text{mol of } H_2O}}$$

For example: how many molecules are there in 3 mol of H_2O? To calculate the number of molecules of H_2O in 3 moles of H_2O, (moles $\rightarrow$ molecules) you need to set up a conversion factor, starting with the given information (3 moles) and using the mol-to-molecule conversion factor with mol on the bottom:

$$3\,\cancel{\text{moles of H}_2\text{O}} \times \frac{6.02\times10^{23}\ \text{molecules of H}_2\text{O}}{1\,\cancel{\text{mol of H}_2\text{O}}} = 1.80\times10^{24}\text{molecules of H}_2\text{O}$$

If you need to convert molecules to moles (molecules $\rightarrow$ moles), you just need to follow the same procedure, using the conversion factor between mol-to-molecule with molecules in the bottom. For example 3×10^{20} H_2O molecules equals to 4.98×10^{-4}moles of H_2O as

$$3\times10^{20}\,\cancel{\text{molecules of H}_2\text{O}} \times \frac{1\text{moles of H}_2\text{O}}{6.02\times10^{23}\,\cancel{\text{molecules of H}_2\text{O}}} = 4.98\times10^{-4}\text{moles of H}_2\text{O}$$

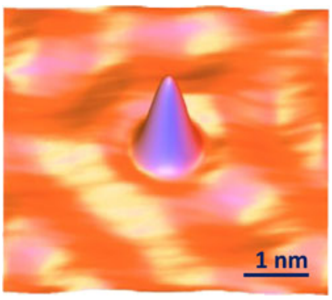
`From moles to atoms` Molecules are made of atoms, and for example, the CO_2 molecule contains an atom of C and two atoms of O. To convert from molecules to atoms (molecules $\rightarrow$ atoms) you need to use the coefficients in the molecular formula. For example, a H_2O molecule contains an atom of O and two atoms of H, and hence the relation between water molecules and H and O atoms is:

$$\boxed{\frac{1\ \text{molecule of H}_2\text{O}}{1\ \text{atom of O}}\ \text{and}\ \frac{1\ \text{molecule of H}_2\text{O}}{2\ \text{atom of H}}}$$

To convert from moles into atoms (moles $\rightarrow$ atoms) you need to use a two-step process in a single line. First, you convert from moles into molecules, to then convert molecules into atoms. For example, 3 moles of H_2O contains 1.6×10^{24} H atoms, as:

$$3\,\cancel{\text{moles of H}_2\text{O}} \times \frac{6.02\times10^{23}\ \cancel{\text{molecules of H}_2\text{O}}}{1\,\cancel{\text{moles of H}_2\text{O}}} \times \frac{2\text{H atoms}}{1\ \cancel{\text{molecules of H}_2\text{O}}}$$
$$= 1.6\times10^{24}\text{atoms of H}$$

▼Molecules are counted by the mole



© www.wallpaperflare.com

▼Eggs are bought by the dozen

© www.wallpaperflare.com

▼Socks are bought as pairs

© www.wallpaperflare.com

▼A ream of paper contains 500 sheets

© wikipedia

▼A six-pack contains 6 beers

© www.wallpaperflare.com

▼A gross is a dozen of dozens

© Flickr

Figure 6.1 Collections of items and their name

Sample Problem 59

Calculate: (a) the number of CuO molecules in 3.4 moles of CuO; (b) the number of moles of CO in 5×10^{20} CO molecules

SOLUTION

(a) 3.4 moles of CuO equals to 2.05×10^{22} molecules of CuO as:

$$3.4\cancel{\text{moles of CuO}} \times \frac{6.022 \times 10^{23} \text{ molecules of CuO}}{1\cancel{\text{mole of CuO}}} = $$
$$= 2.05 \times 10^{24} \text{molecules of CuO}$$

(b) 5×10^{20} CO molecules equals to 8.3×10^{-4} moles of CO, as

$$5 \times 10^{20}\cancel{\text{CO molecules}} \times \frac{1 \text{ mole of CO}}{6.022 \times 10^{23} \cancel{\text{CO molecules}}} = 8.3 \times 10^{-4}\text{moles of CO}$$

STUDY CHECK

Calculate the number of C, H, N and O atoms in 3.5 moles of caffeine ($C_8H_{10}N_4O_2$).

The following example shows how to relate atoms and moles.

Sample Problem 60

Calculate the number of O atoms in 4.5 moles of NO_2.

SOLUTION

(c) 4.5 moles of NO_2 contains 5.4×10^{24} O atoms, as

$$4.5\cancel{\text{moles of } NO_2} \times \frac{6.022 \times 10^{23}\, \cancel{NO_2 \text{ molecules}}}{1\cancel{\text{moles of } NO_2}} \times \frac{2\text{O atoms}}{1\cancel{NO_2 \text{ molecules}}}$$
$$= 5.4 \times 10^{24} \text{O atoms}$$

6.2 Converting moles into grams and into atoms

A standard way to measure chemicals in the lab is by weight. We can weigh different quantities and the larger the quantity the larger the weight. For a chemical, the weight of a mole is called the molar (or molecular) weight. For example, if we weigh a mole of water (H_2O) we will be weighing 18 grams of water, or if you weigh a mole of table salt (NaCl) the scale will show 58 grams. In this section, you will learn how to calculate the molar mass of a chemical and how to use this property to convert from weight to moles (and moles to weight).

`Molar mass of a chemical` Chemicals are made of atoms, and each atom has a specific atomic weight (AW) listed in the periodic table. For example, the atomic weight of Na is 23 grams whereas the atomic weight of Cl is 35 g. The weight of all the atoms of a molecule is called the molecular weight (we call this also molar weight or MW). For example, the molecular weight of NaCl is 58 g, as the weight of Na and Cl is 23 and 35g. Another example would be water, H_2O with a molecular weight of 18g–as the atomic weight of H and O is 1 and 16 g, respectively, and the molecule has two

▼The molecular mass of cinnamic acid ($C_9H_8O_2$), used in the manufacture of flavors, is $148.16\frac{g}{mol}$

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▼Car batteries contains sulfuric acid (H_2SO_4), a corrosive chemical with a molar mass of $96.07\frac{g}{mol}$

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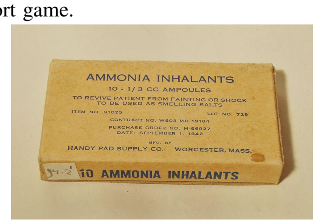
▼Ammonia smelling salts ($(NH_4)_2CO_3$, MW=$96g/mol$) were historically employed to wake up injured athlete during a sport game.



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▼Acetic acid is an organic acid with molar mass $60g/mol$

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H atoms. The units for molecular weight is $\frac{g}{mol}$, also written as g/mol. To compute the molar mass of a molecule you need to break down the molecule into atoms using the coefficients in the formula. For example, the formula for vinegar is $C_2H_4O_2$ which means a vinegar molecule contains 2C, 4H and 2O atoms. If you add the atomic masses of 2C, 4H, and 2O you will get $60g/mol$. If the chemical formula has a parenthesis, you need to open up the parenthesis to calculate the total number of atoms. As an example, $Ca(NO_3)_2$ contains 1Ca, 2N, and 6O, and its molar mass is $164.09g/mol$.

Sample Problem 61

Calculate: (a) The atomic weight of Mg; (b) the molecular mass of sulfuric acid, H_2SO_4

SOLUTION

(a) According to the periodic table the atomic weight (AW) of Mg is $24.31g/mol$. (b) The molar mass of H_2SO_4 is the result of adding the atomic masses of 2H (AW=$1g/mol$) atoms, 1 S (AW=$32g/mol$) and 4O (AW=$16g/mol$) atoms, that gives $98.08g/mol$.

STUDY CHECK

Calculate the molar mass of glucose $C_6H_{12}O_6$

`From moles to grams` The molar mass is used to convert moles to grams or grams to mol. For example, the molar mass of water is $18g/mol$. This means:

$$\boxed{1\text{mole of } H_2O = 18\text{g of } H_2O}$$

that is the same as

$$\boxed{\frac{1 \text{ mole of } H_2O}{18 \text{ grams of } H_2O} \text{ or } \frac{18 \text{ grams of } H_2O}{1 \text{ mole of } H_2O}}$$

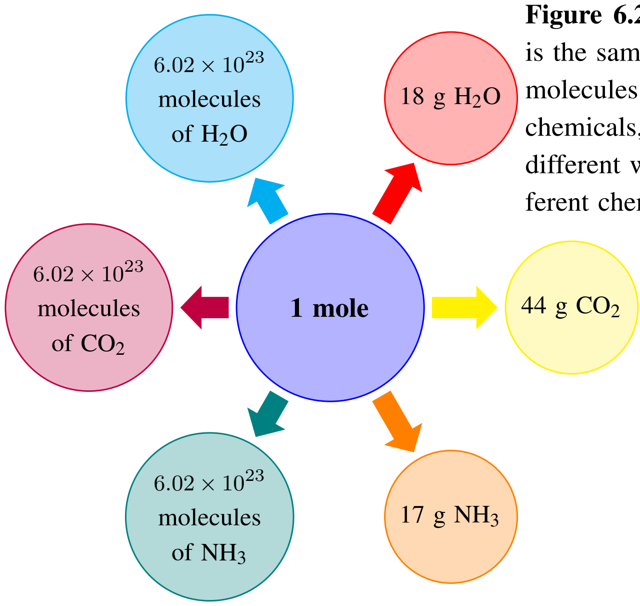

Figure 6.2 A mole is the same number of molecules for different chemicals, while being different weights for different chemicals.

Sample Problem 62

Smelling salts ($(NH_4)_2CO_3$) are chemical used to arouse consciousness. These are used by pro athletes to get into the zone before a game. How many moles of salt do you have in 100 grams of these salts?

SOLUTION

We first need to calculate the molar mass of $(NH_4)_2CO_3$, a chemical with 2N, 8H, 1C and 3O atoms. The molar mass hence would be: $2 \times 14 + 8 \times 1 + 1 \times 12 + 3 \times 16 = 96g/mol$. In order to calculate the moles given the gram, you need to use the molar mass a conversion factor:

$$100\,\cancel{\text{g of }(\text{NH}_4)_2\text{CO}_3} \times \frac{\text{moles of }(\text{NH}_4)_2\text{CO}_3}{96\ \cancel{\text{g of }(\text{NH}_4)_2\text{CO}_3}} = 1.04 \text{ moles of } (\text{NH}_4)_2\text{CO}_3$$

STUDY CHECK

Calculate the MW of table salt (NaCl) and the grams in 20 moles of this salt.

From grams to atoms In the previous sections, we covered how to convert grams to moles, moles to molecules, or molecules to atoms. You can follow the diagram below to switch from one of these properties (atoms, molecules, moles, grams) to another. For example, if you want to convert grams into moles, you will only need one step and you will only have to use a single property: the molar mass. Differently, if you need to convert grams into molecules you will have to use two different steps and use two different properties: the molar mass and Avogadro's number.

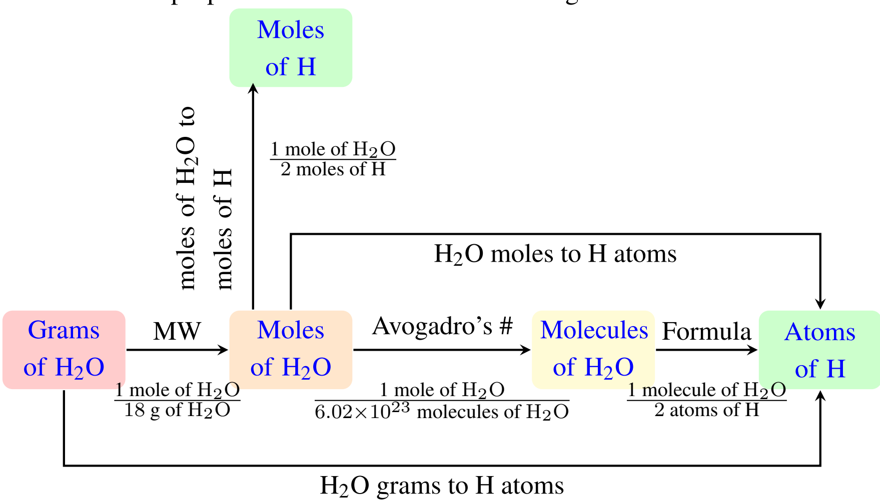

Figure 6.3 Diagrams relating grams, moles of molecules, moles of atoms, atoms and molecules.

Sample Problem 63

Convert 10 grams of ammonia (NH_3, MW=$17g/mol$) into H atoms.

SOLUTION

We will have to do this conversion in three different steps. First we will go from grams to moles, then from moles to molecules to finally transform molecules

▼A termite reaction between iron(III) oxide and Al: $Fe_2O_3 + 2\,Al \longrightarrow 2\,Fe + Al_2O_3$

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▼An image of the combustion of Mg: $2\,Mg(s) + O_2(g) \longrightarrow 2\,MgO(s)$

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▼Iron rust is the result of a combination reaction: $4\,Fe + 2\,O_2 \longrightarrow 2\,Fe_2O_3$

© wikipedia

▼Wood burning is a combustion reaction

© PngImg

into atoms:

$$10\,\cancel{\text{g of NH}_3} \times \frac{1 \text{ mole of NH}_3}{\cancel{17 \text{ g of NH}_3}} \times \frac{6.022 \times 10^{23} \text{ NH}_3 \text{ molecules}}{\cancel{1 \text{ mole of NH}_3}} \times \frac{3 \text{ H atoms}}{\cancel{1 \text{ NH}_3 \text{ molecule}}} = 1.8 \times 10^{24} \text{H atoms}$$

STUDY CHECK

Methane is a chemical used as a fuel. Calculate how many grams of methane CH_4 contains 5×10^{25} H atoms.

6.3 Chemical reactions

When we eat we burn food with molecular oxygen (O_2) to produce carbon dioxide and water. Similarly, when we start the engine of the car to go to work, gasoline burns to produce the same chemicals: CO_2 and H_2O. These are two examples of chemical reactions, but there are many other examples. Nitrogen from the air reacts with hydrogen to produce ammonia, a common chemical used in the production of fertilizers. This section covers the basics of chemical reactions. You will learn how to balance reactions and how to classify reactions.

Simple chemical reactions Magnesium is a metal that reacts with oxygen to produce magnesium oxide. Magnesium is solid $Mg_{(s)}$ whereas oxygen is gas and contains two oxygen atoms per molecule $O_{2(g)}$. Magnesium oxide, the result of the reaction, is solid $MgO_{(s)}$. The reaction between magnesium and oxygen to produce magnesium oxide

$$2\,Mg_{(s)} + 1\,O_{2(g)} \longrightarrow 2\,MgO_{(s)}$$

Mg and O_2 combine–that is why we use a plus sign– to produce MgO–we use an arrow to indicate that a chemical is being produced. Also, the symbols *(s)* or *(g)* indicates solid or gas state. The reactants are located before the arrow and the products are after the arrow. The numbers in front of the reactants and products (2,1 and 2) are called stoichiometric coefficients, and we will talk more about them in the following sections.

Reading a chemical reaction Chemical reactions can be read in words. To read a chemical reaction you need to connect the reactants with the word “react” and then use the words “to produce” and after that, you need to read the products. The numbers in front of the reactants and products represent the number of moles, and you need to include those numbers in the reading. For example, the following reaction

$$2\,Mg(s) + O_2(g) \longrightarrow 2\,MgO(s)$$

should be read as: “two moles of Mg react with one mole of O_2 to produce two moles of MgO”.

Balanced chemical reactions Chemical reactions contain molecules, which are made of atoms. Some chemical reactions are balanced, and others need to be balanced. To identify a balanced reaction, you should use the stoichiometric coefficients and the indexes in the molecular formulas to break down the reactants and products into atoms. In a balanced chemical reaction, the atoms of reactants should be the same as the atoms of the products. Consider the following reaction,

$$2\,Mg(s) + O_2(g) \longrightarrow 2\,MgO(s).$$

The table below shows all reactants and products in the form of atoms.

$2\,Mg(s) + O_2(g) \longrightarrow 2\,MgO(s)$		
Reactants	Products	
2Mg	2Mg	✓
O_2=2O	2O	✓

The number of Mg atoms in the reactants and products is the same and equals two. On the other hand, the number of O atoms in the reactants and products is the same, equal to two. For this reason, we say this reaction is *balanced*.
Now consider the following reaction:

$$C(s) + O_2(g) \longrightarrow CO(g),$$

The number of C atoms in the reactants and products is the same and equals one. In contrast, the number of O atoms in the reactants and products is not the same, and for this reason, we say this reaction is *not balanced*.

$C(s) + O_2(g) \longrightarrow CO(g)$		
Reactants	Products	
1C	1C	✓
O_2=2O	O	✗

`Balancing chemical reactions` To balance a reaction, we need to introduce the stoichiometric coefficients that make the number of atoms of reactants and products the same. To balance the number of oxygens, we will multiply CO by two, and that will give us two oxygens and two carbons as well. If we do this, now the carbon atoms of reactants and products will not be the same. We can solve this by multiplying C(s) by two. The following table summarizes the changes we made:

$2\,C(s) + O_2(g) \longrightarrow 2\,CO(g)$		
Reactants	Products	
2C	2C	✓
O_2=2O	2O	✓

The reaction is now balanced after introducing two stoichiometric coefficients and the number of C and O atoms in the reactant molecules and products is the same.

Sample Problem 64

Balance the following reaction:

$$CH_4(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

SOLUTION

We will break down each molecule into atoms. In the case of O, both CO_2 and H_2O contain oxygen and hence you will have to combine both oxygen atoms:

$$CH_4(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

Reactants	Products	
1C	1C	✓
4H	2H	✗
2O	3O	✗

The reaction is not balanced as the number of H and O atoms for the reactants and products is not the same. In order to balance the H, you can multiply by two H_2O, and that will balance H but also affect O.

$$CH_4(g) + O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g)$$

Reactants	Products	
1C	1C	✓
4H	4H	✓
2O	4O	✗

You can balance O by multiplying O_2 by two. That will give you the final balanced reaction in which all atoms (O, H and C) are the same in the product and reactant molecules.

$$CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g)$$

Reactants	Products	
1C	1C	✓
4H	4H	✓
4O	4O	✓

STUDY CHECK

Balance the following reaction: $Fe_2O_3(s) + C(s) \longrightarrow Fe(s) + CO(g)$

6.4 Five types of reactions

Most chemical reactions can be classified according to five types: combination, decomposition, single replacement, double replacement, and combustion.
In a *combinations reaction* two reactants combine to generate a product. An example of a combination is the reaction between Mg and oxygen to produce MgO:

$$2\,Mg(s) + O_2(g) \longrightarrow 2\,MgO(s) \text{ (combination)}$$

In a *decomposition reaction* a single reactant breaks down into several products. An example of a decomposition reaction is the thermal reaction of $CaCO_3$ to produce calcium oxide (CaO) and carbon dioxide

$$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g) \text{ (decomposition)}$$

In a single replacement reaction, an element replaces another element in a chemical. An example would be the reaction of Zn with HCl, in which Zn replaces hydrogen:

$$Zn(s) + 2\,HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g) \text{ (Single replacement)}$$

In a double replacement reaction, the first element in the reacting compounds switches places. An example is the reaction between $AgNO_3$ and NaCl, in which Ag from $AgNO_3$ replaces Na in NaCl:

$$AgNO_3(aq) + NaCl(aq) \longrightarrow NaNO_3(aq) + AgCl(s) \text{ (Double replacement)}$$

Finally, in a combustion reaction, a carbon-based chemical reacts with oxygen to produce carbon dioxide and water. An example would be the combustion of methane (CH_4):

$$CH_4(g) + 2\,O_2(g) \longrightarrow 2\,H_2O(g) + CO_2(g) \text{ (Combustion)}$$

6.5 Stoichiometry

In the previous section, you learned how to balance a chemical reaction. To do this, you had to find the stoichiometric coefficients that balance the atoms of the reactant and products. In this new section, we will learn how to use those coefficients to predict the amount of product formed. You will also learn how to predict the amount of reactant needed to react with another reactant. We will use the reaction between Mg and oxygen:

$$2\,Mg(s) + O_2(g) \longrightarrow 2\,MgO(s)$$

in which two moles of Mg and one mole of O_2 produce two moles of MgO. We will refer to this reaction in the following.

`Mole-Mole ratio` A chemical reaction can be expressed in the form of conversion factors. For example, the mole ratio between Mg (a reactant) and O_2 (another reactant) and between Mg and MgO is:

$$\boxed{\frac{\text{2 moles of Mg}}{\text{1 moles of O}_2} \text{ or } \frac{\text{1 moles of O}_2}{\text{2 moles of Mg}} \text{ and } \frac{\text{2 moles of Mg}}{\text{2 moles of MgO}} \text{ or } \frac{\text{2 moles of MgO}}{\text{2 moles of Mg}}}$$

Finally, the mole ratio between O_2 and MgO is:

$$\boxed{\frac{\text{1 moles of O}_2}{\text{2 moles of MgO}} \text{ or } \frac{\text{2 moles of MgO}}{\text{1 moles of O}_2}}$$

Mole ratios are used, for example, to transform the amount of reactant into the product.

`Reactants to products` We will calculate how much MgO will be produced from 5 moles of Mg by converting Mg into MgO using the conversion factor between both chemicals. As we want to transform the Mg into MgO we will use the conversion factor with Mg at the bottom of the fraction. This way the units will cancel out to give moles of MgO:

$$5 \cancel{\text{ moles of Mg}} \times \frac{\text{2 moles of MgO}}{2 \cancel{\text{ moles of Mg}}} = \text{5 moles of MgO.}$$

This result means that 5 moles of Mg will produce 5 moles of MgO.

`Reactant to a different reactant` Sometimes we will have to calculate how much reactant will be needed to react with another reactant. In those cases, we will use the conversion factor that relates both reactants. If we have 5 moles of Mg and we want to know how much oxygen we need to react with Mg, we will proceed as:

$$5 \cancel{\text{ moles of Mg}} \times \frac{\text{1 mole of O}_2}{2 \cancel{\text{ moles of Mg}}} = \text{2.5 moles of O}_2.$$

This result means that 2.5 moles of O_2 will react with 5 moles of Mg.

Sample Problem 65

Hydrogen reacts with oxygen to produce water according to the following reaction

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(g)$$

Calculate: (a) the number of moles of water produced from 5 moles of H_2; (b) the number of moles of oxygen needed to react with 7 moles of H_2.

SOLUTION

(a) we will first convert 5 moles of H_2 into water:

$$5\,\cancel{\text{moles of } H_2} \times \frac{2 \text{ moles of } H_2O}{2\,\cancel{\text{moles of } H_2}} = 5 \text{ moles of } H_2O,$$

that is: 5 moles of hydrogen produce 5 moles of water. (b) We will now calculate the amount of oxygen needed to react with 7 moles of hydrogen

$$7\,\cancel{\text{moles of } H_2} \times \frac{1 \text{ moles of } O_2}{2\,\cancel{\text{moles of } H_2}} = 3.5 \text{ moles of } O_2,$$

that is: 3.5 moles of O_2 will react with 7 moles of H_2.

STUDY CHECK

Hydrogen reacts with oxygen to produce water according to the following reaction

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(g)$$

Calculate the number of moles of water produced by 4 moles of oxygen.

6.6 Mass calculations

In the previous sections, given the moles of reactants, we learned how to use chemical reactions to predict the amount of product formed. In this section, we will learn how to do the same, but instead of starting with the number of moles, this time, we will work our way starting with a quantity given in grams. We will base the following examples on the reaction of Mg and O_2:

$$2\,Mg(s) + O_2(g) \longrightarrow 2\,MgO(s)$$

`Molar mass review` Remember that the molar mass (also known as molecular weight, MW. Atomic weight for the atoms, AW) of a chemical is a property used to convert grams into moles. For example, the molar mass of H_2O is $18g/mol$. If we need to convert 12 grams of water into mols we should do:

$$12\,\cancel{\text{g of } H_2O} \times \frac{\text{moles of } H_2O}{18\,\cancel{\text{g of } H_2O}} = 0.66 \text{ moles of } H_2O$$

To use the stoichiometric coefficients from a chemical reaction, the starting quantity must be the moles of a reactant of products. This is because these coefficients are expressed in moles and hence to operate with them you can only use moles.

`From grams to moles` If we want to calculate the grams of MgO produced from 3 moles of Mg, we will start with the moles of Mg and use a conversion factor that relates moles of Mg and moles of MgO, locating the moles of Mg on the bottom:

$$3\ \cancel{\text{moles of Mg}} \times \frac{2\ \text{moles of MgO}}{2\ \cancel{\text{moles of Mg}}} = 3\ \text{moles of MgO}.$$

Now we aim to calculate the number of MgO moles produced from 5 grams of Mg (AW=$24g/mol$). This time, we will have first to convert the grams of Mg into moles to then use the mole ratio between Mg and MgO:

$$5\ \cancel{\text{g of Mg}} \times \frac{1\ \cancel{\text{moles of Mg}}}{24\ \cancel{\text{g of Mg}}} \times \frac{2\ \text{moles of MgO}}{2\ \cancel{\text{moles of Mg}}} = 0.21\ \text{moles of MgO}.$$

`From grams to grams` We want to answer the following question: we have 6 moles of Mg (AW=$24g/mol$) and we want to calculate the mass of MgO(MW=$40g/mol$) produced. The importance of this type of calculation is based on the fact that in a chemistry laboratory chemicals are normally weighted in grams. To answer this example, we will start with the grams of Mg and convert this quantity to moles of Mg using the atomic mass. After that, we will use the mole ratio between Mg and MgO to calculate the moles of MgO. At this point, we will finish the calculation by converting the moles of MgO into grams using its molecular weight:

$$5\ \cancel{\text{grams of Mg}} \times \frac{1\ \cancel{\text{moles of Mg}}}{24\ \cancel{\text{grams of Mg}}} \times \frac{2\ \cancel{\text{moles of MgO}}}{2\ \cancel{\text{moles of Mg}}} \times \frac{40\ \text{grams of MgO}}{1\ \cancel{\text{moles of MgO}}}$$
$$= 10\ \text{grams of MgO}.$$

Overall, we have that 5 grams of Mg will produce 10 grams of MgO.

Sample Problem 66

For the reaction of hydrogen and nitrogen to produce ammonia (NH_3):

$$3\,H_2(g) + N_2(g) \longrightarrow 2\,NH_3(g)$$

Calculate: (a) the number of moles of NH_3 produced from 10 grams of hydrogen (MW=$2g/mol$); (b) Calculate the number of grams of NH_3 (MW=$17g/mol$) produced from 10 grams of nitrogen (MW=$28g/mol$)

SOLUTION

(a) You will solve this problem in a single line by using two steps: first convert the grams of hydrogen to moles, to then convert the moles of hydrogen into ammonia:

$$10\ \cancel{\text{grams of H}_2} \times \frac{1\ \cancel{\text{moles of H}_2}}{2\ \cancel{\text{grams of H}_2}} \times \frac{2\ \text{moles of NH}_3}{3\ \cancel{\text{moles of H}_2}} = 3.33\ \text{moles of NH}_3.$$

In other words, 10 grams of hydrogen produce 3.33 moles of ammonia. (b) To solve this question, we will use one additional step in order to convert the moles of ammonia into grams:

$$10\ \cancel{\text{grams of N}_2} \times \frac{1\ \cancel{\text{moles of N}_2}}{28\ \cancel{\text{grams of N}_2}} \times \frac{2\ \cancel{\text{moles of NH}_3}}{1\ \cancel{\text{moles of N}_2}} \times \frac{17\ \text{grams of NH}_3}{1\ \cancel{\text{moles of NH}_3}}$$
$$= 12.14\ \text{grams of NH}_3.$$

STUDY CHECK

For the reaction of hydrogen and nitrogen to produce ammonia (NH_3):

$$3\,H_2(g) + N_2(g) \longrightarrow 2\,NH_3(g)$$

Calculate the number of grams of nitrogen needed to react with 3 grams of hydrogen to produce ammonia.

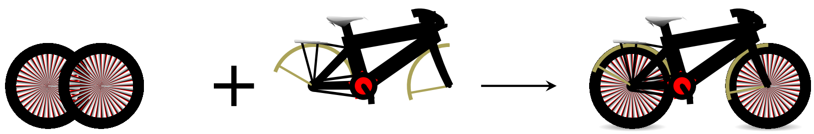
6.7 Percent yield

Often, reactions do not fully proceed to completion. Fewer amounts of products are obtained when a specific quantity of a chemical is expected from a reaction. The percent yield tells how much of the final chemical is produced. The larger the yield the larger the amount of products generated and therefore, less waste is generated. At the same time, when mixing two reactants sometimes leftovers remain. In other words, sometimes there is a reagent that limits the reaction and an excess reagent. This section will cover the very important ideas of percent yield and limiting (or excess) reagents.

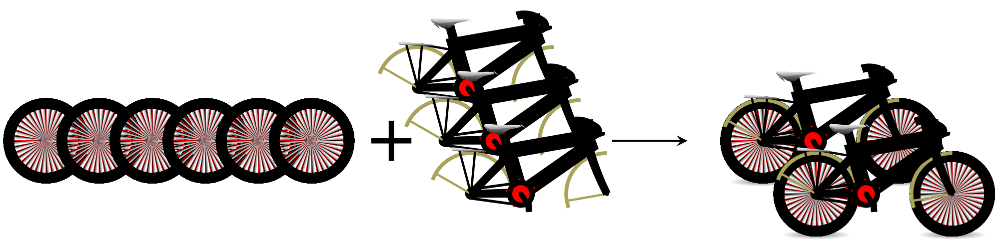
`Percent yield` Chemical reactions are less than perfect as the staring material does not fully convert into products. Furthermore, side reactions often occur in competition with the main reaction generating byproducts. The theoretical yield is the number of products one will expect in a hypothetically perfect chemical reaction, while the actual yield is the amount of product that is produced. Therefore, the percent yield is just the fraction between the actual yield and the theoretical yield in percent form:

$$\% \text{ Yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

To compute the percent yield, we need the theoretical yield, which is the amount of product that will be produced following the rules of stoichiometry, and the actual yield, normally known in an experiment from lab measurements. Look at the scenario below: we need two wheels and a body to ensemble a bike. We can think about the bike production yield based on the starting materials we have.



If we have three pairs of wheels and three bodies and we only produce two bikes, the yield will be 67%.



Sample Problem 67

Ammonia is produced by reacting molecular nitrogen and molecular hydrogen following the reaction:

$$3\,H_{2(g)} + N_{2(g)} \longrightarrow 2\,NH_{3(g)}$$

What is the percent yield of ammonia if 4 moles of hydrogen gives 2 moles of ammonia.

SOLUTION

We will first compute the theoretical yield, that is the moles of ammonia produced from 4 moles of hydrogen:

$$4 \cancel{\text{ moles of H}_2} \times \frac{2 \text{ moles of NH}_3}{3 \cancel{\text{ moles of H}_2}} = 2.66 \text{ moles of NH}_3.$$

Hence the theoretical yield is 2.66 moles of NH_3, and the actual yield given in the problem is 2 moles of NH_3. The actual yield is smaller than the theoretical yield. That is reasonable as the moles of ammonia produced in the real experiment should always be smaller than the amount of ammonia produced in theory. To calculate the percent yield we use the formula:

$$\text{Percent Yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \frac{2}{2.66} \times 100 = 75.18\%$$

STUDY CHECK

Ammonia is produced by reacting molecular nitrogen and molecular hydrogen following the reaction:

$$3\,H_{2(g)} + N_{2(g)} \longrightarrow 2\,NH_{3(g)}$$

Calculate the percent yield of ammonia if 1 moles of nitrogen gives 1.5 moles of ammonia

6.8 Limiting reagent

Let us consider the reaction between solid carbon and molecular oxygen to produce CO_2:

$$C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)}$$

This reaction should be read as "one mole of carbon reacts with one mole of molecular oxygen to produce one mole of carbon dioxide". Mind that to produce one mole of CO_2 you need one mole of C and one mole of molecular oxygen. Now, what would happen if you mix one mole of carbon with 0.5 moles of oxygen? In this scenario, when the 0.5 moles of oxygen are consumed the reaction will stop; half a mole of carbon dioxide will be formed and half a mole of C will remain. Differently, no oxygen will remain and the whole 0.5 moles will consume. We say that C is the excess reactant and oxygen is the limiting reactant. Often a reaction will be accompanied by the reagent quantities and you will have to identify the limiting reagent. This reagent will limit the amount of product formed, and hence, any stoichiometric calculation aimed to predict the amount of product formed should be based on the limiting reagent and never on the excess reagent. Next, we will explain how to systematically identify the limiting reagent.

Identify the limiting reagent Consider the following reaction, in which 3 moles of Mg react with 2 moles of oxygen to produce magnesium oxide:

$$\underset{3\text{ moles}}{2\,Mg} + \underset{2\text{ moles}}{O_2} \longrightarrow 2\,MgO$$

To identify the limiting reagent, we will choose one of the given reagent quantities and calculate the moles of the other reagent needed (we will call this n_{needed}). For example, if we choose to start with 3 moles of Mg, the amount of oxygen needed to react with this quantity will be:

$$n^{O_2}_{needed} = 3 \cancel{\text{ moles of Mg}} \times \frac{1 \text{ moles of } O_2}{2 \cancel{\text{ moles of Mg}}} = 1.5 \text{ moles of } O_2.$$

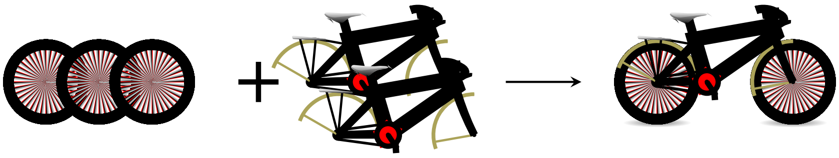
This means that to react with 3 moles of Mg you need $n^{O_2}_{needed} = 1.5$ moles of oxygen. On the other hand, you have two moles of oxygen ($n^{O_2}_{given} = 2$), and that is more than what you need to react with the 3 moles of Mg. Hence,

$$n^{O_2}_{needed} < n^{O_2}_{given}$$

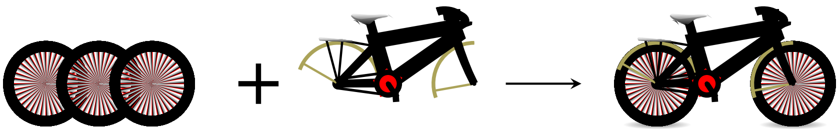
oxygen is the excess reagent and Mg is the limiting reagent. The leftovers will be:

$$n^{O_2}_{left} = |n^{O_2}_{needed} - n^{O_2}_{given}| = 0.5 \text{ moles of } O_2$$

Let us reinforce the idea of limiting reagents with the scenario below. If we have three wheels and two bodies, the number of wheels limits bike production.



Differently, if we have three wheels and one body, then the body of the bikes limit production.



Sample Problem 68

In the synthesis of ammonia (NH_3):

$$\underset{2\text{ moles}}{3\,H_2(g)} + \underset{2\text{ moles}}{N_2(g)} \longrightarrow 2\,NH_3(g)$$

you mix 2 moles of hydrogen with 2 moles of nitrogen. Identify the limiting and excess reagents and indicate the moles of leftover remaining.

SOLUTION

We will choose one of the reagents, for example the two moles of hydrogen, and calculate the amount of nitrogen needed to react with this amount of oxygen:

$$2 \cancel{\text{ moles of } H_2} \times \frac{1 \text{ moles of } N_2}{3 \cancel{\text{ moles of } H_2}} = 0.66 \text{ moles of } N_2.$$

Therefore, to react with 2 moles of hydrogen we need 0.66 moles of nitrogen, and we have 2 moles of nitrogen. This means we have more nitrogen than what we need and hence, nitrogen is the excess reagent and hydrogen the limiting reagent. As we need 0.66 moles of nitrogen and we have 2 moles, 1.33 moles of nitrogen will remain.

STUDY CHECK

In the synthesis of ammonia (NH_3):

$$3\,H_{2(g)} + N_{2(g)} \longrightarrow 2\,NH_{3(g)}$$

you mix 3 moles of hydrogen with 0.5 moles of nitrogen. Identify the limiting and excess reagents and indicate the moles of leftover remaining.

6.9 Reactions and energy

`Exothermic and endothermic reactions` Some reactions release heat; these are called exothermic reactions. Others absorb heat and are called endothermic reactions. Think for example the combustion of the gas in a cooking stove. This reaction produces heat and hence it is exothermic. Differently, bread needs heat to rise. Hence, the chemical reaction involved in bread making should be endothermic. Similarly, if you melt an ice cube you need to give energy to the cube so that it becomes water. The melting of ice is an endothermic reaction. Endothermic reactions have positive heat of reaction ΔH_R and are normally accompanied by an indication of heat in the reactants side of the reaction before the arrow.

$$\text{Reactants} + \quad \longrightarrow \text{Products} \qquad \Delta H_R > 0$$

Differently, exothermic reactions normally accompanied by an expression of heat in the products side have negative ΔH_R.

$$\text{Reactants} \longrightarrow \text{Products} + \qquad \Delta H_R < 0$$

Figure 6.4 represents endo and exothermic processes.

Sample Problem 69

Classify the following reactions as exothermic and endothermic:

$$2\,H_2\,(g) + O_{2(g)} \longrightarrow 2\,H_2O\,(g) + \qquad \Delta H_R = -572 kJ.$$

$$6\,CO_2\,(g) + 6\,H_2O_{(l)} + \quad \longrightarrow C_6H_{12}O_6\,(s) + 6\,O_2\,(g) \qquad \Delta H_R = +47 kJ.$$

SOLUTION

The first reaction is exothermic as it produces hear. Differently, the second reaction is endothermic and absorbs energy.

STUDY CHECK

Classify the following reactions as exothermic and endothermic:

$$C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)} + \qquad \Delta H_R = -393.5 kJ.$$

$$C_{(s)} + 2\,S_{(s)} + \quad \longrightarrow CS_{2(l)} \qquad \Delta H_R = +87.9 kJ.$$

$$S_{(s)} + O_{2(g)} \longrightarrow SO_{2(g)} + \qquad \Delta H_R = -296.8 kJ.$$

`Heat-Mole conversions` Remember that a chemical reaction can be translated into a series of conversion factors that relate the moles of reactants with the products or with other reactants. At the same time, a chemical reaction involving heat can be converted into a series of conversion factors that are related to energy and the moles of reactants and products. For the exothermic reaction:

$$2\,H_2\,(g) + O_{2(g)} \longrightarrow 2\,H_2O\,(g) \qquad \Delta H_R = -572KJ.$$

the moles of hydrogen are related to heat:

$$\boxed{\frac{\text{2 moles of } H_2}{\text{-572 KJ}} \text{ or } \frac{\text{-572 KJ}}{\text{2 moles of } H_2}}$$

Similarly, we can relate energy with moles of O_2 or moles of water:

$$\boxed{\frac{\text{1 moles of } O_2}{\text{-572 KJ}} \text{ or } \frac{\text{-572 KJ}}{\text{1 moles of } O_2}} \quad \boxed{\frac{\text{2 moles of } H_2O}{\text{-572 KJ}} \text{ or } \frac{\text{-572 KJ}}{\text{2 moles of } H_2O}}$$

We will use these relationships to convert moles of reactants or products into heat.

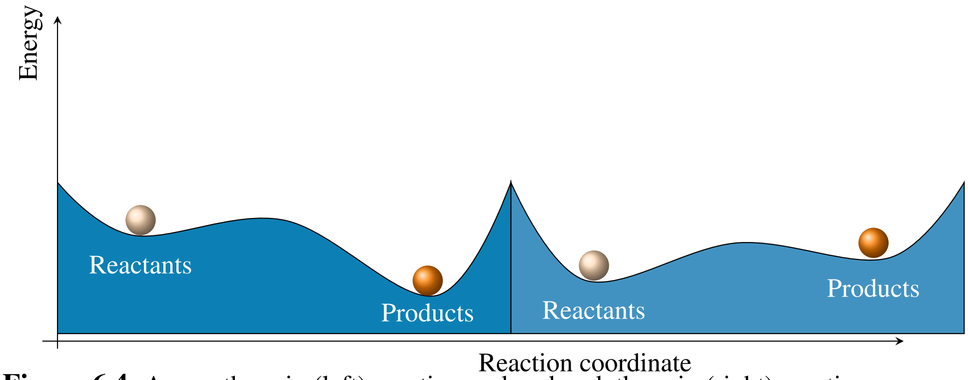


Figure 6.4 An exothermic (left) reaction and and endothermic (right) reaction.

Sample Problem 70

Hydrogen reacts with nitrogen to produce ammonia (NH_3) according to the following reaction

$$3\,H_{2(g)} + N_{2(g)} \longrightarrow 2\,NH_{3(g)} \qquad \Delta H_R = -92KJ.$$

Calculate: (a) the enthalpy of reaction; (b) indicate whether the reaction is endo or exothermic; (c) calculate the heat exchanged when produced 5 moles of ammonia.

SOLUTION

(a) the heat of reaction is -92KJ, and (b) the reaction is exothermic as the heat appears as a product. This means the reaction produces heat. (c) We will use the conversion factor that relates ammonia with heat and will set up the moles of ammonia on the bottom of the conversion factor so that the units will cancel and energy will remain

$$5\ \cancel{\text{moles of } NH_3} \times \frac{\text{-92KJ}}{2\ \cancel{\text{moles of } NH_3}} = -230KJ,$$

that is: 5 moles of ammonia produce -230KJ. The fact that this value is negative means that heat will be released.

STUDY CHECK

Given the reaction below, calculate the number of hydrogen moles needed to generate -200KJ.

$$3\,H_{2(g)} + N_{2(g)} \longrightarrow 2\,NH_{3(g)} \qquad \Delta H_R = -92KJ.$$

CHAPTER 6

The Mole

6.1 Calculate the number of molecules in: (a) 8 moles of CO (b) 10 moles of CO_2

6.2 Calculate the number of molecules in: (a) 4 moles of NH_3 (b) 50 moles of H_2SO_4

6.3 Calculate the number of moles in: (a) 6×10^{23} molecules of NO (b) 5×10^{15} molecules of NaCl(c) 3×10^{27} molecules of MgO

6.4 Calculate the number of moles in: (a) 3.2×10^{21} molecules of H_2O (b) 2×10^{23} molecules of CO_2

6.5 Fill the conversion factor that calculates the final property:

$$10^{24}\ \cancel{\text{molecules of NO}_2} \times \frac{\underline{\qquad}\ \text{moles of NO}_2}{\underline{\qquad}\ \cancel{\text{molecules of NO}_2}} = \underline{\qquad}\ \text{moles of NO}_2.$$

6.6 Fill the conversion factor that calculates the final property:

$$3\ \cancel{\text{moles of NO}} \times \frac{\underline{\qquad}\ \text{molecules of NO}}{\underline{\qquad}\ \cancel{\text{moles of NO}}} = \underline{\qquad}\ \text{molecules of NO}.$$

6.7 Fill the conversion factor that calculates the final property:

$$6\ \cancel{\text{moles of C}_6\text{H}_{12}\text{O}_6} \times \frac{\underline{\qquad}}{\underline{\qquad}} = \underline{\qquad}\ \text{molecules of C}_6\text{H}_{12}\text{O}_6.$$

6.8 Fill the conversion factor that calculates the final property:

$$10^{25}\ \cancel{\text{molecules of CH}_4\text{N}_2\text{O}} \times \frac{\underline{\qquad}}{\underline{\qquad}} = \underline{\qquad}\ \text{moles of CH}_4\text{N}_2\text{O}.$$

Converting moles into grams and into atoms

6.9 Calculate the molar weight of the following molecules: (a) CO (b) H_2 (c) $Fe_2(CO_3)_3$

6.10 Calculate the molar weight of the following molecules: (a) NH_3 (b) O_2

6.11 Fill the conversion factor that calculates the final property:

$$4\ \cancel{\text{moles of CO}_2} \times \frac{\underline{\qquad}\ \text{g of CO}_2}{\underline{\qquad}\ \cancel{\text{moles of CO}_2}} = \underline{\qquad}\ \text{g of CO}_2.$$

6.12 Fill the conversion factor that calculates the final property:

$$10\ \cancel{\text{g of NO}} \times \frac{\underline{\qquad}\ \text{moles of NO}}{\underline{\qquad}\ \cancel{\text{g of NO}}} = \underline{\qquad}\ \text{moles of NO}.$$

6.13 Fill the conversion factor that calculates the final property:

$$5\ \cancel{\text{moles of C}_6\text{H}_{12}\text{O}_6} \times \frac{\underline{\qquad}}{\underline{\qquad}} = \underline{\qquad}\ \text{g of C}_6\text{H}_{12}\text{O}_6.$$

6.14 Fill the conversion factor that calculates the final property:

$$7\ \cancel{\text{g of CH}_4\text{N}_2\text{O}} \times \frac{\underline{\qquad}}{\underline{\qquad}} = \underline{\qquad}\ \text{moles of CH}_4\text{N}_2\text{O}.$$

6.15 Fill the conversion factor that calculates the final property:

$$10^{26}\ \cancel{\text{molecules of NO}_2} \times \frac{\underline{\qquad}\ \text{atoms of O}}{\underline{\qquad}\ \cancel{\text{molecules of NO}_2}} = \underline{\qquad}\ \text{atoms of O}.$$

6.16 Fill the conversion factor that calculates the final property:

$$10^{22}\ \cancel{\text{atoms of O}} \times \frac{\underline{\hspace{3em}}\ \text{molecules of } H_2O}{\underline{\hspace{3em}}\ \cancel{\text{atoms of O}}}$$
$$= \underline{\hspace{3em}}\ \text{molecules of } H_2O$$

6.17 Fill the conversion factor that calculates the final property:

$$6\ \cancel{\text{molecules of } C_6H_{12}O_6} \times \frac{\underline{\hspace{3em}}}{\underline{\hspace{3em}}}$$
$$= \underline{\hspace{3em}}\ \text{atoms of C.}$$

6.18 Fill the conversion factor that calculates the final property:

$$10^{21}\ \cancel{\text{atoms of N}} \times \frac{\underline{\hspace{3em}}}{\underline{\hspace{3em}}}$$
$$= \underline{\hspace{3em}}\ \text{molecules of } CH_4N_2O.$$

6.19 Answer the following questions: (a) Calculate the number of C atoms in 3 moles of $C_{10}H_{14}N_2$? (b) Calculate the number of H atoms in 3 moles of $C_{10}H_{14}N_2$? (c) Calculate the number of N atoms in 3 moles of $C_{10}H_{14}N_2$

6.20 Answer the following questions: (a) How many grams are there in 4 moles of $C_6H_{12}O_6$? (b) How many C atoms are there in 3 moles of $C_6H_{12}O_6$?(c) How many O atoms are there in 3 moles of $C_6H_{12}O_6$?

6.21 Calculate the molar weight of the following molecules: (a) benzene, C_6H_6 (b) Carbon disulfide, CS_2 (c) Nitrogen tetroxide, N_2O_4

6.22 Calculate the molar weight of the following molecules: (a) Sulfur dioxide, SO_2 (b) Unsymmetrical dimethyl hydrazine, $(CH_3)_2NNH_2$ (c) Dimethyl sulfide, $(CH_3)_2S$

6.23 Fill the conversion factor that calculates the final property, given that the molar mass of C_2H_6 is 30g/mol:

$$7 \times 10^{21}\ \text{atoms of C} \times \frac{\underline{\hspace{3em}}}{\underline{\hspace{3em}}} \times \frac{\underline{\hspace{3em}}}{\underline{\hspace{3em}}}$$
$$= \underline{\hspace{3em}}\ \text{moles of } C_2H_6$$

6.24 Fill the conversion factor that calculates the final property, given that the molar mass of C_2H_6 is 30g/mol:

$$5 \times 10^{25}\ \text{atoms of H} \times \frac{\underline{\hspace{3em}}}{\underline{\hspace{3em}}} \times \frac{\underline{\hspace{3em}}}{\underline{\hspace{3em}}}$$
$$\times \frac{\underline{\hspace{3em}}}{\underline{\hspace{3em}}} = \underline{\hspace{3em}}\ \text{g of } C_2H_6$$

Chemical Reactions

6.25 Balance the following reactions: (a) $P_4(s) + O_2(g) \longrightarrow P_4O_{10}(s)$ (b) $Al(s) + O_2(g) \longrightarrow Al_2O_3(s)$

6.26 Balance the following reactions: (a) $FeS(s) + O_2(g) \longrightarrow Fe_2O_3(s) + SO_2(g)$ (b) $NH_3(g) + O_2(g) \longrightarrow NO(g) + H_2O(g)$

6.27 Classify next reaction as combination, decomposition, single replacement, double replacement, or combustion: (a) $Pb_{(s)} + FeSO_{4(s)} \longrightarrow PbSO_{4(s)} + Fe_{(s)}$ (b) $C_6H_{12(g)} + 9\,O_{2(g)} \longrightarrow 6\,CO_{2(g)} + 6\,H_2O_{(g)}$ (c) $2\,RbNO_{3(aq)} + BeF_{2(aq)} \longrightarrow Be(NO_3)_{2(aq)} + 2\,RbF_{(aq)}$

6.28 Balance the following reactions: (a) $H_{2(g)} + Br_{2(g)} \longrightarrow HBr_{(g)}$ (b) $C_{(g)} + O_{2(g)} \longrightarrow CO_{(g)}$ (c) $O_{3(g)} \longrightarrow O_{2(g)}$ (d) $NH_4NO_{2(aq)} \longrightarrow N_{2(g)} + H_2O_{(l)}$ (e) $Na_3PO_{4(aq)} + MgCl_{2(aq)} \longrightarrow Mg_3(PO_4)_{2(aq)} + NaCl_{(aq)}$

Stoichiometry and mass calculations

6.29 Fill the mole ratio for the following reaction:

$$C_6H_{12}O_6(s) + 6\,O_2(g) \longrightarrow 6\,CO_2(g) + 6\,H_2O(g)$$
$$\frac{\underline{\hspace{3em}}\ \text{moles of } C_6H_{12}O_6}{\underline{\hspace{3em}}\ \text{moles of } O_2}$$

6.30 Fill the mole ratio for the following reaction:

$$C_6H_{12}O_6(s) + 6\,O_2(g) \longrightarrow 6\,CO_2(g) + 6\,H_2O(g)$$
$$\frac{\underline{\hspace{3em}}\ \text{moles of } O_2}{\underline{\hspace{3em}}\ \text{moles of } CO_2}$$

6.31 Fill the conversion factor that calculates the moles of oxygen needed to react with 2 moles of Silver producing AgO:

$$2\,Ag(s) + O_2(g) \longrightarrow 2\,AgO(s)$$
$$2\ \cancel{\text{moles of Ag}} \times \frac{\underline{\hspace{3em}}\ \text{moles of } O_2}{\underline{\hspace{3em}}\ \cancel{\text{moles of Ag}}}$$
$$= 1\ \text{moles of } O_2.$$

6.32 Fill the conversion factor that calculates the moles of AgO produced from 2 moles of Silver:

$$2\,Ag(s) + O_2(g) \longrightarrow 2\,AgO(s)$$

$$2\ \cancel{\text{moles of Ag}} \times \frac{\underline{\qquad}\ \text{moles of AgO}}{\underline{\qquad}\ \cancel{\text{moles of Ag}}} = 2\ \text{moles of AgO}.$$

6.33 Fill the conversion factor that calculates the moles of AgO produced from 10 moles of oxygen:

$$2\,Ag(s) + O_2(g) \longrightarrow 2\,AgO(s)$$

$$10\ \cancel{\text{moles of } O_2} \times \frac{\underline{\qquad}\ \text{moles of AgO}}{\underline{\qquad}\ \cancel{\text{moles of } O_2}} = 20\ \text{moles of AgO}.$$

6.34 Fill the conversion factor that calculates the moles of AgO produced from 5 moles of oxygen:

$$2\,Ag(s) + O_2(g) \longrightarrow 2\,AgO(s)$$

$$5\ \cancel{\text{moles of } O_2} \times \frac{\underline{\qquad}\ \text{moles of AgO}}{\underline{\qquad}\ \cancel{\text{moles of } O_2}} = 10\ \text{moles of AgO}.$$

6.35 Calculate how many moles of nitrogen are needed to react with 5 moles of hydrogen, to produce ammonia:

$$\underset{5\text{ moles}}{3\,H_2(g)} + \underset{\underline{\qquad}\text{ moles}}{N_2(g)} \longrightarrow 2\,NH_3(g)$$

6.36 Calculate the number of grams of nitrogen needed to react with 4 moles of hydrogen, to produce ammonia:

$$\underset{4\text{ moles}}{3\,H_2(g)} + \underset{\underline{\qquad}\text{ grams}}{N_2(g)} \longrightarrow 2\,NH_3(g)$$

6.37 Calculate the number of grams of hydrogen needed to react with 0.3 moles of nitrogen, to produce ammonia:

$$\underset{\underline{\qquad}\text{ grams}}{3\,H_2(g)} + \underset{0.3\text{ moles}}{N_2(g)} \longrightarrow 2\,NH_3(g)$$

6.38 Calculate the number of grams of hydrogen needed to react with 5 moles of nitrogen, to produce ammonia:

$$\underset{\underline{\qquad}\text{ grams}}{3\,H_2(g)} + \underset{0.3\text{ moles}}{N_2(g)} \longrightarrow 2\,NH_3(g)$$

Percent Yield and Limiting Reagent

6.39 Six moles of nitrogen gas react to produce three moles of ammonia according to the following reaction:

$$3\,H_2(g) + \underset{6\text{ moles}}{N_2(g)} \longrightarrow \underset{3\text{ moles}}{2\,NH_3(g)}$$

6.40 Six moles of nitrogen gas react to produce two moles of ammonia according to the following reaction:

$$3\,H_2(g) + \underset{6\text{ moles}}{N_2(g)} \longrightarrow \underset{2\text{ moles}}{2\,NH_3(g)}$$

6.41 We mix three moles of hydrogen gas with three moles of nitrogen gas.

$$\underset{3\text{ moles}}{3\,H_2(g)} + \underset{3\text{ moles}}{N_2(g)} \longrightarrow 2\,NH_3(g)$$

Calculate the limiting reagent.

6.42 We mix three moles of hydrogen gas with half a mole of nitrogen gas.

$$\underset{3\text{ moles}}{3\,H_2(g)} + \underset{0.5\text{ moles}}{N_2(g)} \longrightarrow 2\,NH_3(g)$$

Calculate the limiting reagent.

6.43 We mix two moles of hydrogen gas with five moles of nitrogen gas.

$$\underset{2\text{ moles}}{3\,H_2(g)} + \underset{5\text{ moles}}{N_2(g)} \longrightarrow 2\,NH_3(g)$$

Calculate the limiting reagent.

6.44 We mix six moles of hydrogen gas with two moles of nitrogen gas.

$$\underset{6\text{ moles}}{3\,H_2(g)} + \underset{2\text{ moles}}{N_2(g)} \longrightarrow 2\,NH_3(g)$$

Calculate the limiting reagent.

6.45 Liquid mercury reacts with gas chlorine to produce mercury(II) choride, a white solid:

$$Hg(l) + Cl_2(g) \longrightarrow HgCl_2(s)$$

In an experiment, 5-mL of mercury (AW=200.59g/mol) with density 5g/mL, reactants with 4-g of chlorine to produce 6g of $HgCl_2$. What is the percent yield of the reaction.

6.46 Mercury(II) halides can be converted into mercury (I) halides by combination with metallic mercury. Mercury(I) halides are known as mercurous halides. When chlorine is the halide, the resulting mercury salt is known as calomel:

$$HgCl_2(s) + Hg(l) \longrightarrow Hg_2Cl_2(s)$$

When 2 grams of mercury(II) chloride reacts to produce 2 grams of calomel, calculate the percent yield of the reaction.

6.47 The Wurtz reaction results from the reaction of bromomethane (CH_3Br) with sodium to produce ethylene (C_2H_6)

$$2\,CH_3Br_{(g)} + 2\,Na_{(s)} \longrightarrow C_2H_{6(g)} + 2\,NaBr_{(s)}$$

How many grams of sodium are need to produce 3g of ethylene given that the yield of the reaction is 30%.

6.48 Nitriles with stannous chloride ($SnCl_2$) in the presence of hydrochloric acid produce an imine.

$$\underset{nitrile}{CH_3CN(aq)} + SnCl_{2(aq)} + HCl_{(aq)} \longrightarrow \underset{imine}{CH_3CHNH(aq)}$$

How many moles of imine (MW=43g/mol) are produced if we react 3g of nitrile (MW=41g/mol), with 2g of stannous chloride (MW=188g/mol) and 1g of hydrochloric acid (MW=36g/mol).

Reactions and energy

6.49 Identify the following reaction as endothermic or exothermic.

$$C_6H_{12}O_6(s) + 6\,O_{2(g)} \longrightarrow 6\,CO_{2(g)} + 6\,H_2O_{(g)} \qquad \Delta H^\circ_R = -2800KJ/mol$$

6.50 Identify the following reaction as endothermic or exothermic.

$$B_2O_3(s) + 3\,H_2O_{(g)} \longrightarrow 3\,O_{2(g)} + B_2H_{6(g)} \qquad \Delta H^\circ_R = 2035KJ/mol$$

6.51 For the following reaction:

$$C_6H_{12}O_6(s) + 6\,O_{2(g)} \longrightarrow 6\,CO_{2(g)} + 6\,H_2O_{(g)} \qquad \Delta H^\circ_R = -2800KJ/mol$$

Fill the conversion factor:

$$\frac{\_\_\_\_\_\_ \text{ moles of } O_2}{\text{-2800 KJ}}$$

6.52 In the following combustion reaction:

$$C_6H_{12}O_6(s) + 6\,O_{2(g)} \longrightarrow 6\,CO_{2(g)} + 6\,H_2O_{(g)} \qquad \Delta H^\circ_R = -2800KJ/mol$$

glucose ($C_6H_{12}O_6$) burns to produce carbon dioxide and water. Calculate the heat involved in the combustion of 3 moles of glucose.

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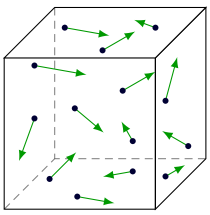
Ch. 7. Reactions in gase phase

THE air we all breathe contains numerous gases, such as oxygen, nitrogen, or carbon dioxide. Some of these gases are indeed essential for life. As an example, plants take up carbon dioxide to give off oxygen, and water is produced by the reaction of oxygen and hydrogen gas. Other gases are dangerous for life. An example is carbon monoxide, which results from gas stoves, heating systems, and fire. This is a colorless, odorless, and tasteless gas that can bind to the blood displacing oxygen. As a consequence, carbon monoxide can build up in closed environments causing death. This chapter deals with the properties of gases. You will learn how to calculate the volume or pressure of a gas, characterizing its state. You will also learn how to work with mixtures of gases and for example predict the pressure of oxygen in an atmosphere containing numerous gases.

7.1 Gases and its properties

Gases contain atomic or molecular particles. They have very different properties than liquids or solids. The particles of a gas are spread and far away from each other. Liquids, on the other hand, are made of loose particles that interact through weak forces. Solids, on the other hand, are packed materials, and their particles, atoms, or molecules, are closer together. This section covers the different properties of gases.

`Gases in the periodic table` Some of the elements in the periodic table are molecular gases, resulting from the combination of two atoms of the same element. For example, molecular oxygen (O_2) is a gas. Similarly, molecular nitrogen (N_2), molecular hydrogen (H_2), molecular chlorine (Cl_2), or molecular fluorine (F_2) are all diatomic gases–they contain two atoms of the same element. Other gases result from the combination of two different non-metals. Examples are carbon monoxide (CO) or dioxide (CO_2), and nitrogen monoxide (NO) or dioxide (NO_2). The noble gases (Ne, He, Ar) also exist in the gas state.



▲Microscopic picture of a gas

`Characteristics of gases` Gases have different properties compared to solids or liquids:

- Gases assume the volume and shape of their container. As they expand, they have no shape different than their container's shape.
- Gases are compressible: they can be compressed, reducing their volume. Differently, liquids and solids are incompressible.
- The density of gases is small, compared to the one for solids and liquids.

`Volume, V` The volume of a gas (V) is the amount of space it occupies, and gases fully occupy the volume of their container. Liters (L) and milliliters (mL) are units of volume. A liter is a cubic unit and one litter equals a cubic decimeter (dm^3).

`Temperature, T` The temperature (T) of a gas is related to the speed (the average velocity) of its particles. The higher the temperature the higher the particles' speed. Although there are different units of temperature such as Kelvin (K) or celsius (C°), in this chapter many formulas require the use of Kelvin temperature (T_K), that related to celsius (T_C) by the formula

$$\boxed{T_K = T_C + 273}$$

▼Chlorine reactive is a gas

▼Hydrogen is a flammable gas used to power some cars

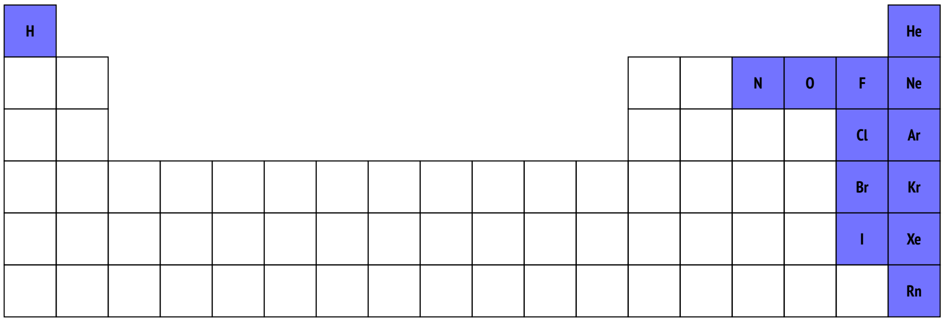


Figure 7.1 All gas elements; some monoatomic others diatomic molecules. For example, oxygen is found as O_2, and hydrogen as H_2. Differently, Neon is found as a monoatomic gas.

`Amount of gas, n` The amount of a gas (n) refers to the number of gas particles. The larger the amount of gas, the larger the number of gas particles. The amount of gas is measured in moles or grams. Figure 7.1 displays a list of all gas elements.

`Pressure, P` In general, pressure is defined as force divided by surface. In the international system, the unit of force is the Newton (N) and the unit of area (A) is the m^2. One newton is $1 kg \cdot m/s^2$.

$$\boxed{P = \frac{F}{A}}$$

The particles of a gas are constantly moving. On their movement, they frequently hit the walls of their container, like raindrops hitting the ceiling when it rains. When they hit the walls they exert pressure, and pressure is defined as the force acting on a certain area. The larger pressure the stronger the collisions with the walls and the higher the frequency of collision–the stronger the force applied to the walls. Imagine you are driving a motorcycle. While you drive you can feel the collision of the air's particles with your face. The faster you go the higher pressure. The value of air pressure is measured with a barometer and depends on your location on the earth–in particular your altitude–as well as the weather. If you are at sea level the atmospheric pressure is one unit of pressure (one atm), due to the air that you have on top of you. If you climb a mountain, the pressure decreases, as there is less air on top of you. The higher you are from the sea level, the lower the air pressure. The weather also affects pressure, and on hot days the pressure of air is higher, whereas on cold days pressure is lower.
Units of pressure are bars, atmospheres (atm), torr, pascals (Pa), or millimeters of mercury (mmHg). To covert pressure units, you can use the following conversion factors:

$$\boxed{\frac{1 \text{ atm}}{1.01325 \text{ bar}}} \quad \boxed{\frac{1 \text{ atm}}{760 \text{ mmHg}}} \quad \boxed{\frac{1 \text{ torr}}{1 \text{ mmHg}}} \quad \boxed{\frac{1 \text{ atm}}{101325 \text{ Pa}}} \quad \boxed{\frac{1 \text{ atm}}{14.7 \text{ lb/in}^2 \text{(psi)}}}$$

one millimeter of mercury (mmHg) is the same as 1 torr. As a note, the name torr acknowledges the person who invented the barometer: Torricelli, an Italian physicist. The US equivalent to 1 atm is 14.7 psi. For example, the pressure inside the tires of a car ranges from 30 to 35 psi.

Sample Problem 71

An oxygen sample has a pressure of 2 atm. Convert this value to: (a) mmHg and (b) Pascals.

SOLUTION

(a) we start by placing the given data (2 atm) and using the conversion factor between atm and mmHg, with the atm unit on the bottom, so that the units cancel

$$2\,\cancel{\text{atm}} \times \frac{760\text{ mmHg}}{1\,\cancel{\text{atm}}} = 1520\text{mmHg}$$

(b) we proceed as in (a) but using the conversion between atm and Pa:

$$2\,\cancel{\text{atm}} \times \frac{101325\text{ Pa}}{1\,\cancel{\text{atm}}} = 2.02 \times 10^5\text{Pa}$$

STUDY CHECK

An oxygen sample has a pressure of 730 mmHg. Convert this value to atmospheres.

Ideal gas law in terms of moles The ideal gas law says:

$$\boxed{PV = nRT}$$ Ideal Gas Law

where:

P is the pressure of the gas in atm

V is the volume of the gas in L

n is the number of moles of the gas

T is the temperature of the gas in K

R is the constant of the gas $0.082\frac{\text{atm}\cdot\text{L}}{\text{mol}\cdot\text{K}}$

Imagine for example that you inflate a balloon with your mouth, introducing air particles into the balloon. While the number of particles inside the balloon grows, its volume will grow too. More particles will collide with the walls of the balloon and hence, the pressure inside the balloon will also increase.

Sample Problem 72

Helium gas is used to inflate blimps, scientific balloons and party balloons. What is the volume in litters of a 0.2 moles Helium balloon at 300K and 2 atm.

SOLUTION

Given	Asking

Analyze the Problem		
	$T = 300K$ $P = 2atm$ $n = 0.2mol$ $R = 0.082\frac{\text{atm}\cdot\text{L}}{\text{mol}\cdot\text{K}}$	V

Using now the ideal gases formula: $PV = nRT$, we have

$$2\cancel{\text{atm}} \cdot V = 0.2\cancel{\text{mol}} \cdot 0.082\frac{\cancel{\text{atm}} \cdot \text{L}}{\text{mol} \cdot \cancel{\text{K}}} \cdot 300\cancel{K}$$

All units but L cancel out. Solving for V we have 2.46 L.

STUDY CHECK

What is the pressure in atmospheres of a 1 L balloon containing 3 moles of Helium at 40C°.

`Ideal gas law in terms of density` The ideal gas law in terms of density is:

$$\boxed{P \cdot MW = DRT} \quad \text{Ideal Gas Law in terms of D}$$

where:

P is the pressure of the gas in atm

MW is the molecular weight (or atomic weight, AW) of the gas in g/mol

D is the density in $g \cdot L^{-1}$

T is the temperature of the gas in K

R is the constant of the gas $0.082\frac{\text{atm}\cdot\text{L}}{\text{mol}\cdot\text{K}}$

We use this formula when we are questioned about the molar mass or density of the gas.

Sample Problem 73

What is the density of a Helium balloon at 400K and 3 atm.

SOLUTION

Besides the data in the problem, as the gas is He we already know its atomic mass from the periodic table:

	Given	Asking
Analyze the Problem	$T = 400K$ $P = 3atm$ $AW = 4g \cdot mol^{-1}$ $R = 0.082\frac{\text{atm}\cdot\text{L}}{\text{mol}\cdot\text{K}}$	D

Using now the ideal gases formula in terms of density: $P \cdot MW = DRT$, we have

$$3\cancel{\text{atm}} \cdot 4\frac{g}{\cancel{\text{mol}}} = D \cdot 0.082\frac{\cancel{\text{atm}} \cdot \text{L}}{\cancel{\text{mol}} \cdot \cancel{\text{K}}} \cdot 400\cancel{K}$$

Solving for D we have 0.36 $g \cdot L^{-1}$.

STUDY CHECK

What is the molecular mass of a 4 $g \cdot L^{-1}$ density gas at 30C° and 5 atm.

STP conditions STP conditions refer to standard temperature (273K) and pressure (1 atm) conditions. Working at STP conditions means the pressure will be fixed at 1 atm and temperature at 273K.

$\boxed{\text{1 atm and 273K}}$ STP Conditions

Sample Problem 74

Calculate the volume in litters of 5 moles of nitrogen at STP conditions.

SOLUTION

From the problem we have the following data:

Analyze the Problem	Given	Asking
	$n = 5moles$ $P = 1atm$ $T = 273K$	V

We need to apply the ideal gas formula with the set of given variables:

$$1\cancel{\text{atm}} \cdot V = 5\cancel{\text{mol}} \cdot 0.082\frac{\cancel{\text{atm}} \cdot \text{L}}{\text{mol} \cdot \cancel{\text{K}}} \cdot 273\cancel{K}$$

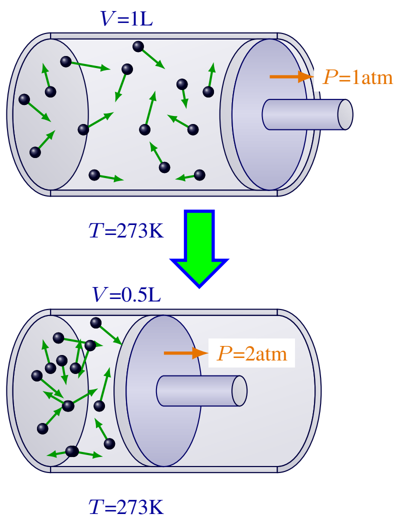
and solving for V we have a final volume of 112L.

STUDY CHECK

Calculate the grams in 4L of N_2 at STP conditions.

7.2 Change of gas properties

The previous section addressed the properties of an ideal gas. However, as all properties of a gas are related, if we modify one the others will change too. This section covers situations in which one of the gas properties changes (e.g. V changes) and you need to predict the change of another gas property (e.g. P). For example, imagine you compress a balloon with your hand. The temperature and number of moles of the gas inside the balloon are constant, as the balloon is closed and in contact with the atmosphere. Differently, the pressure and volume will change. In particular, the volume will decrease and the pressure will increase. This means that the gas particles will hit the balloon harder and with more frequency.



▲ Boyle's Law

Solving problems with an initial and final state To solve problems in which two of the gas variables are kept fixed and the other two are fixed, one needs to apply the ideal gas law at the initial and final state to then divide both formulas. Imagine a situation in which you have a 1L hot air balloon with 1 mole of gas and you add gas to a total of 5 moles. You want to calculate the final volume after you inflate the volume, knowing the temperature and pressure are kept constant. The initial state corresponds to 1L and 1 mole of gas and the final estate corresponds to an unknown volume and 5 moles. Using the ideal gas formula twice you have:

$$\left.\begin{array}{r} PV_1 = n_1RT \\ \\ PV_2 = n_2RT \end{array}\right\} \quad \frac{PV_1}{PV_2} = \frac{n_1RT}{n_2RT} \tag{7.1}$$

as some of the variables the cancel out:

$$\frac{\cancel{P}V_1}{\cancel{P}V_2} = \frac{n_1\cancel{RT}}{n_2\cancel{RT}} \tag{7.2}$$

and you end up with Avogadros' law. If you plug the numbers into the formula:

$$\frac{1L}{V_2} = \frac{1 \text{ mol}}{5 \text{ mol}} \tag{7.3}$$

and you get a final volume of 5L.

Sample Problem 75

A 3L gas sample has a pressure of 5 atm. If the pressure increases to 10 atm at fixed temperature and number of moles, calculate the final volume of the gas.

SOLUTION

From the problem we have the following data:

	Given	Asking
Analyze the Problem	$V_1 = 3L$ $P_1 = 5atm$ $P_2 = 10atm$	V_2

We need to apply the ideal gas formula to the initial state and final state and divide both formulas. The number of moles and the temperature are constant and will cancel out from both equations:

$$\left.\begin{array}{r} P_1V_1 = nRT \\ \\ P_2V_2 = nRT \end{array}\right\} \quad \frac{P_1V_1}{P_2V_2} = \frac{nRT}{nRT} \tag{7.4}$$

Plugging the values:

$$\frac{P_1V_1}{P_2V_2} = \frac{nRT}{nRT} \tag{7.5}$$

and solving:

$$\frac{3 \cdot 5}{10 \cdot V_2} = 1 \tag{7.6}$$

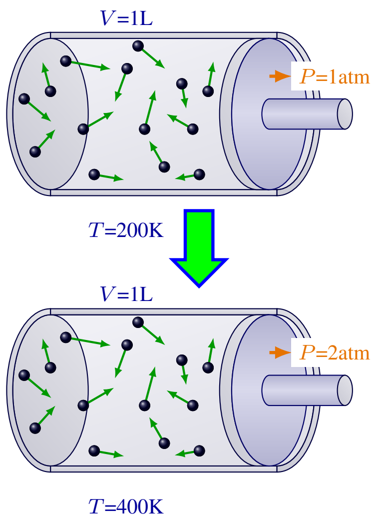
the final volume will be 1.5 L.

STUDY CHECK

A 4 atm gas sample has a temperature of 300K. If we decrease its temperature to 200K at fixed volume and number of moles, calculate the final pressure of the gas.

`Pressure-Volume change` If temperature and the number of moles of gas are kept constant the product of pressure and volume will remain constant too. This is the case of the balloon-pressing example. We call this Boyle's Law:

$$\boxed{\frac{P}{V} = c \quad or \quad P_1 \cdot V_1 = P_2 \cdot V_2} \quad \text{Boyle's law}$$



▲ Gay-Lusscac's Law

where:

P_1, V_1 are the initial pressure and volume

P_2, V_2 are the final pressure and volume

c is a constant

`Pressure-Temperature change` Imagine you cool down a balloon at a fixed volume. What would happen to the balloon's pressure? Based on Gay-Lussac's law, its pressure will decrease:

$$\boxed{\frac{V}{T} = c \quad or \quad \frac{P_1}{T_1} = \frac{P_2}{T_2}} \quad \text{Gay-Lussca's law}$$

where:

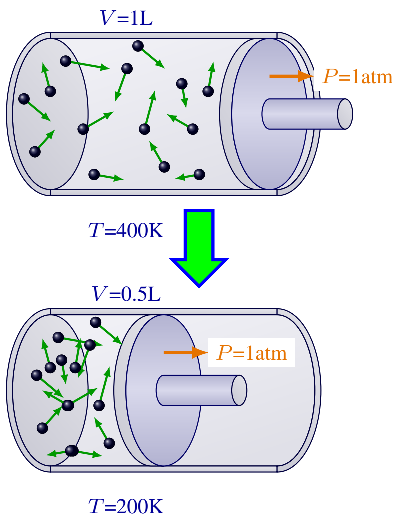
P_1, T_1 are the initial pressure and temperature

P_2, T_2 are the final pressure and temperature

c is a constant

`Volume-Temperature change` Imagine you cool down a balloon at a fixed pressure (under the atmosphere). What would happen to the balloon's volume? Based on Charle's law, its volume will decrease:

$$\boxed{\frac{V}{T} = c \quad or \quad \frac{V_1}{T_1} = \frac{V_2}{T_2}} \quad \text{Charle's law}$$



▲ Charle's Law

where:

V_1, T_1 are the initial volume and temperature

V_2, T_2 are the final volume and temperature

c is a constant

`Volume-Moles change` Imagine a hot air balloon. Air comes in and out of the balloon as the balloon is not closed. Hence the pressure inside the balloon is just the atmospheric pressure. Also as the balloon is in contact with the air, its temperature will be constant, resulting from the thermal equilibrium between the inside of the balloon and the atmosphere. If you inflate the balloon with hot air, the volume of the balloon and the number of moles are related by Avogadro's law:

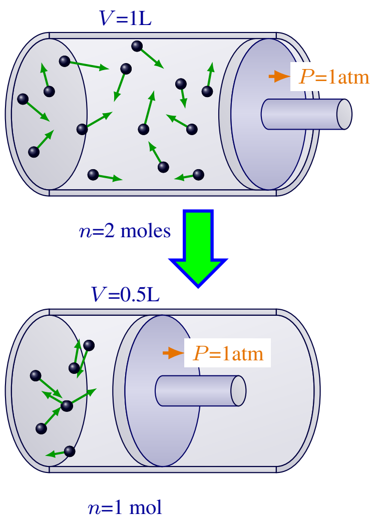
$$\boxed{\frac{V}{n} = c \quad or \quad \frac{V_1}{n_1} = \frac{V_2}{n_2}} \quad \text{Avogadro's law}$$

where:

V_1, n_1 are the initial volume and number of moles

V_2, n_2 are the final volume and number of moles

c is a constant



▲ Avogadro's Law

`Relating the different variables of a gas` The question is now if we increase the pressure at a fixed number of moles and pressure, how do we know if the volume will increase or perhaps decrease? Similarly, if for example, the number of gas moles increases at fixed pressure and volume, will the temperature of the gas increase or perhaps decrease? We can answer these questions by employing the ideal gas law. If the variables that we need to relate are on the same side of the equation (e.g. P and V) then if one of the variables increases the other will decrease. Differently, If the gas variables to relate are located on opposite sides of the gas law (e.g. P and T) then both will change in the same direction. For example, let us consider the changes of P and V (at fixed n and T). As they are on the same side of the ideal gas law ($PV = nRT$) if P increases V will decrease. Differently, for the change of P and T (at fixed V and n), as both variables are on opposite sides of the ideal gas law ($PV = nRT$), if P increases, T will increase as well.

7.3 Mixtures of gases and gas stoichiometry

The air is a mixture of different gases. It contains oxygen (O_2) and nitrogen (N_2) as well as other gases such as carbon dioxide, argon, or water vapor. Only 21% of the air is made of oxygen and 78.2% of nitrogen. The other gases represent 0.8% of the air. The atmospheric pressure is 1 atm and results from the pressure of all the components of the air. Each gas exerts a partial pressure and all combined exert the total atmospheric pressure. In this section, you will learn how to work with mixtures of gases. This section also covers the use of the molar volume to relate moles and volume at standard conditions.

`Molar volume` If we work at STP conditions the volume of one mole of gas equals 22.4L, and we refer to this relationship as the molar volume.

$$\boxed{\frac{1\text{mol}}{22.4\text{ L at STP}}} \quad \text{Molar Volume}$$

This relationship allows us to carry out stoichiometric calculations in a chemical reaction involving gases.

`Stoichiometry and gases` If you encounter chemical reactions with gases, the molar volume relation allows you to carry out stoichiometric calculations. Why is this important? Imagine you have this reaction:

$$\underset{3\text{L}}{2\,H_2(g)} + O_2(g) \longrightarrow \underset{x\text{L}}{2\,H_2O(g)}$$

Gases are measured using their pressure and are more convenient to speak about liters of hydrogen than moles of hydrogen or grams of hydrogen, as hydrogen is a gas. This way, if we start by mixing 3L of H_2 we would like to know how much water is being produced. To calculate this, we will use the stoichiometric coefficients. In previous chapters, we saw that these numbers represent moles and the units of these numbers are mol. If the reaction deals with gases you want to interpret the stoichiometric coefficients in terms of liters. This way:

$$x = 3\,\cancel{\text{L of H}_2} \times \frac{2\text{ L of H}_2\text{O}}{2\,\cancel{\text{L of H}_2}} = 3\text{ L of H}_2\text{O}.$$

Overall, if we mix three liters of hydrogen we obtain 3L of water. In case we know the litters of any of the reactants and we need to calculate the moles of product, then we have to add an extra step to transform litters into moles.

Sample Problem 76

Hydrogen gas reacts with oxygen gas to produce water vapor according to the following equation:

$$\underset{2L}{2\,H_2(g) + O_2(g)} \longrightarrow \underset{x\ mol}{2\,H_2O(g)}$$

Calculate the number of moles of water produced from 2L of oxygen at STP conditions.

SOLUTION

We will solve the problem in a single line, first relating the liters of oxygen and litters of water produced and finally converting litters of water into moles of water using the molar volume. Remember when there are gases in the reaction, the stoichiometric coefficients can be interpreted in terms of liters:

$$x = 2L\,\cancel{\text{L of } O_2} \times \frac{2\,\cancel{\text{L of } H_2O}}{2\,\cancel{\text{L of } O_2}} \times \frac{1 \text{ mol of } H_2O}{22.4\cancel{\text{L of } H_2O}} = 0.178 \text{ mol of } H_2O.$$

We have that two liters of oxygen produce four liters of water. At the same time, 22.4L of water–or any other gas–is 1 moles of that gas. So four L of water are 0.17moles of water.

STUDY CHECK

Hydrogen gas reacts with nitrogen (MW=28 g/mol) gas to produce ammonia (MW=17 g/mol) at STP conditions according to the following equation:

$$3\,H_2(g) + \underset{0.5L}{N_2(g)} \longrightarrow \underset{x\ g}{2\,NH_3(g)}$$

Calculate the number of grams of ammonia produced from 0.5L of nitrogen.

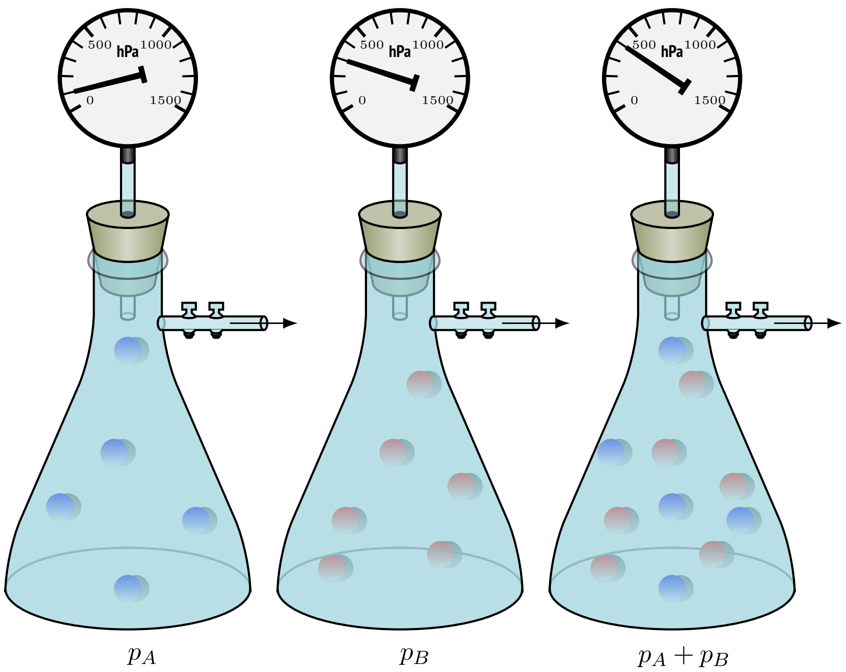


Figure 7.2 A visual representation of Dalton's law of partial pressure: after adding two different gases with different partial pressured, the final pressure is the result of adding both partial pressures.

`Partial and total pressure` Imagine you have a container with 1atm of Ar and another container of the same volume containing 1 atm of Ne. If you combine the containers into a single container (and the temperature does not change), hence the pressure in the container will result from both gases and will be 2 atm. Inside the mixed container, 2 atm will the total pressure (P_{Total}), whereas the partial pressure of each gas (p_1 and p_2) will be 1 atm. Dalton's Law says that the total pressure results from adding the partial pressure of each gas. For a gas mixture with n components:

$$\boxed{P_{Total} = p_1 + p_2 + \ldots p_n}$$ Dalton's Law

Sample Problem 77

Medical Air is a odorless gas made mostly of nitrogen and oxygen, administer by ventilator in hospital settings with an operating gauge pressure of 3 atm. If the oxygen pressure inside a container is 2.37 atm, calculate the partial pressure of nitrogen in the mixture.

SOLUTION

The problem gives the total pressure of the mixture and the partial pressure of one of the components. By using Dalton's law, we know that if the total pressure is 3atm and the partial pressure of oxygen is 2.37, hence the partial pressure of the other component has to be 0.63 atm.

STUDY CHECK

Entonox is a medicinal mixture of dinitrogen oxide (N_2O) and oxygen (O_2). The pressure N_2O in a entonox container is 2 atm and the oxygen pressure is 1520 mmHg as well. Calculate the total pressure in atm in a Entonox container.

`Partial pressure of a gas in a mixture` For a mixture with different gases, the partial pressure of a given gas (A) will depend on the number of moles of that particular gas and the overall volume of the mixture

$$\boxed{p_A = \frac{n_A RT}{V}}$$

`Mole fraction` The mole fraction (X_A) of a gas (A) in a mixture of gas is just the number of moles of this gas over the total number of moles in the mixture. The larger the mole fraction of a gas in a mixture the more molecules of that specific gas are there in the mixture. One can express the morel fraction in terms of partial pressures also, as the pressure of a given gas over the total pressure. For a mixture with n components:

$$\boxed{X_A = \frac{n_A}{n_A + n_B + \cdots + n_n} \quad or \quad \frac{p_A}{p_A + p_B + \cdots + p_n}}$$

For a mixture of gases, the partial pressure of a gas (p_A) is related to the mole fraction of that gas (X_A) and the total pressure of the mixture of gases (P_{Total}):

$$\boxed{p_A = X_A \cdot P_{Total}}$$

Sample Problem 78

A mixture of gases with a total pressure of 2 atm contains 3 moles of Ar, 3 moles of He and 1 moles of Ne. Calculate the partial pressure of each component on the mixture.

SOLUTION

We calculate first the mole fraction for each component of the mixture. As the total number of moles if 7 moles and there are 3 moles of Ar, its mole fraction is 0.43. Similarly, the mole fraction for He is 0.43 and for Ne is 0.14. To calculate the partial pressure of each gas you just need to multiply its mole fraction by the total pressure (2 atm). Hence: p_{Ar}=0.86atm, p_{He}=0.86atm and p_{Ne}=0.28atm

STUDY CHECK

A mixture of gases with a total pressure of 5 atm contains 1 mol of Ar and 1 mol of He. Calculate the partial pressure of each component on the mixture.

7.4 Kinetic molecular theory of gases

At this point, we know enough about the properties of gases to be able to condense all these pieces of information into a quantitative model that could generate numerical predictions. The kinetic model of gases can predict among other properties the particle average velocity–this is technically called root mean square velocity, v_{RMS}.

`Kinetic theory of gases` The kinetic theory of gases is a model that explains the properties of gases. This theory envisions a gas in the form of a set of moving particles. Some of the ideas behind this model are:

- The particles of a gas are in constant motion and move very fast.
- On its movement, gas particles collide with each other changing paths and colliding with the walls of their container exerting pressure.
- Gas particles are far apart from each other, barely interacting.
- The average kinetic energy of the particles of gas (this is the energy of the particles due to movement) is proportional to the temperature of the gas.

Using the kinetic theory we can rationalize the different properties of a gas. As the particles of a gas are in constant motion and apart from each other they fill and occupy the same volume of their container. The temperature of a gas is related to its kinetic energy, that is, the average speed of the gas particles. Also, as the gas particles collide with the container's wall, they exert pressure. The kinetic theory of gases explains for example how room fresheners work. As you spray the room, the molecules of the perfume in a gas state move fast and occupy the room. The kinetic molecular theory of the gases gives a molecular-based description of the temperature of a gas–among other properties. The ideal gas law is experimental; this means is a law that comes from measuring and carrying experiments. However, this law does not provide any reasons behind the behavior of gases, ideal or real. The kinetic molecular theory provides a molecular description of temperature. In particular one of the outcomes of this theory is that the average velocity of a gas particle depends on the square root of the temperature of the gas. More precisely, the way this theory describes velocity is in the form of a *root mean square velocity* v_{RMS}, that is, as an average of the velocity of each particle. The formula that connects the root mean square velocity with temperature is:

$$\boxed{v_{RMS} = \sqrt{\frac{3000RT}{MW}}}$$ root mean square velocity formula

where:

MW is the molecular weight of the gas in g/mol

T is the temperature of the gas in K

R is the constant of the gas in energy units 8.314 $J \cdot K^{-1} \cdot \text{mol}^{-1}$

v_{RMS} is the root mean square velocity in m/s

It is important to notice that the root mean square velocity depends on temperature–the more temperature the more velocity– and is inversely proportional to the molecular weight of the gas–the heavier the mass the lower velocity.

Sample Problem 79

Order the following molecules in increasing order of root mean square velocity: Ne, CO_2 and H_2O.

SOLUTION

Root mean square velocity is inversely proportional to the molecular weight of the gas; hence, the larger the mass the lower velocity. If we compare the molecular weight of the gases: Ne(MW=20g/mol), CO_2(MW=44g/mol) and H_2O(MW=18g/mol). The root mean square velocity of water is the largest and the root mean square velocity of carbon dioxide is the smallest.

STUDY CHECK

Calculate the root mean square velocity of the molecules of water at 25C°.

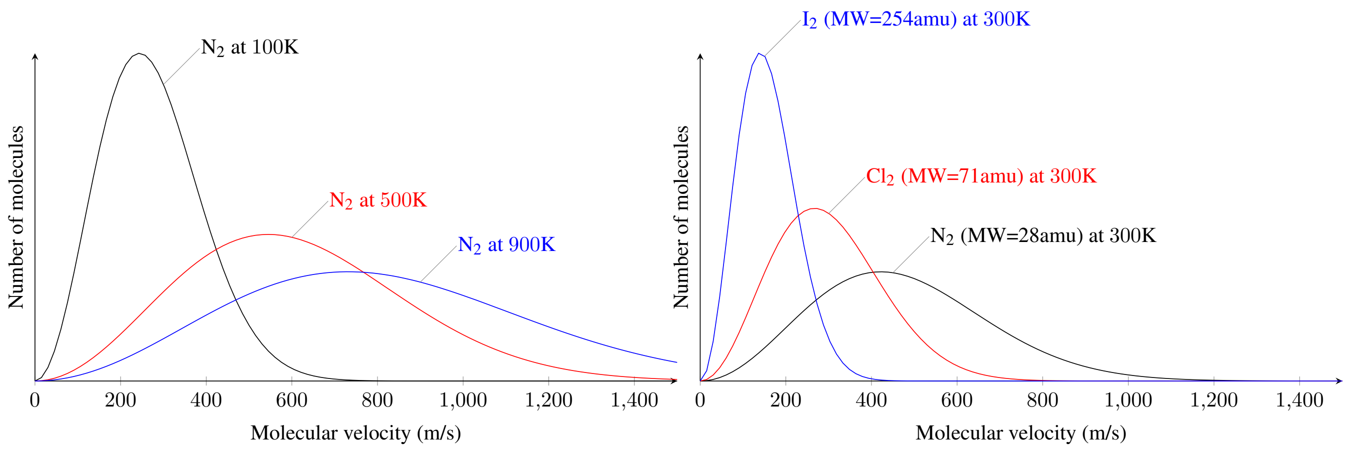


Figure 7.3 Effect of temperature and mass on the distribution of molecular speeds

`Distribution of velocities` The root means square velocity v_{RMS} is just an average of the square velocities of the gas particles. Still, some particles will have faster velocity than v_{RMS}, and others will have slower velocity. The molecular velocities of the particles of gas follow a distribution that is mass and temperature dependent. As shown in Figure 7.3, the higher temperature the larger the root square velocity, with a wider distribution of velocities. At the same time, the larger the molar mass of the gas, the smaller the root square velocity with a thinner distribution of velocities.

CHAPTER 7

Gases And Its Properties

7.1 Convert the following properties: (a) A pressure value of 2 atm into mmHg (b) A pressure value of 3000 Pa into atm (c) A temperature value of 25°C into K

7.2 Convert the following properties: (a) A pressure value of 900 mmHg into torr (b) A temperature value of 400K into °C

Ideal Gas Law

7.3 A gas contained in a 3L tank has a pressure of 5 atm at a temperature of 400 K. Calculate the number of moles in the tank.

7.4 Dinitrogen oxide, used in dentistry, is an anesthetic also called laughing gas. What is the pressure in atm of 0.35 moles of N_2O at 22°C in a 5L container?

7.5 A 4 moles sample of gas at 400K has a pressure of 10 atm. Calculate the volume of the sample.

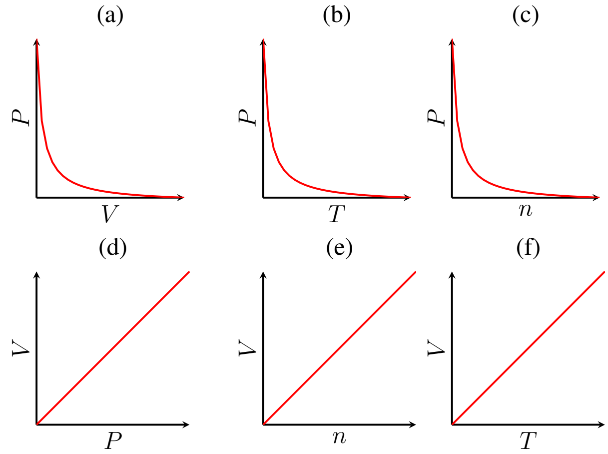
7.6 A 3 grams sample of Ar at 40°C is placed in a 3L container. Calculate the pressure inside the container.

7.7 Eighteen grams of a gas in a 11L container at 400K exert a pressure of 3 atm. Calculate the molar mass of the gas.

7.8 What is the molar mass of a gas if a 3.16 g sample at 0.75 atm and 45°C occupies a volume of 2L.

7.9 Answer the following questions: (a) Calculate the volume of a 4 moles of Ar at STP conditions. (b) Calculate the volume of a 4 moles of Ne at STP conditions. (c) Calculate the moles of gas in 3L of Ar at STP conditions. (d) Calculate the volume of 64 g of O2 gas at STP (273K, 1atm)

7.10 Indicate what plot (or plots) bellow best represent the following gas laws: (a) Boyle's law (b) Charle's law (c) Avogadro's law



Change of Gas Properties

7.11 A sample of a gas at 400K and 12 atm is cooled in the same container to 200K. Calculate the new pressure.

7.12 In a storage area where the temperature has reached 300K, the pressure of oxygen gas in a 15 L steel cylinder is 1 atm. Calculate the volume if the pressure is reduced to 0.5 atm.

7.13 A closed H_2 sample has a volume of 5 L and a pressure of 1 atm. What is the new pressure if the volume is decreased to 2L with no change in temperature and the amount of gas.

7.14 A sample of Ne in a closed, expandable container, has a volume of 3L at 40°C. Calculate the new volume if the container is cooled to 25°C.

7.15 Complete the following statement: if the pressure of a gas increases, at fixed temperature and moles, its volume....

7.16 Complete the following statement: if the temperature of a gas increases, at fixed volume and moles, its pressure....

Mixture of Gases and Gas Stoichiometry

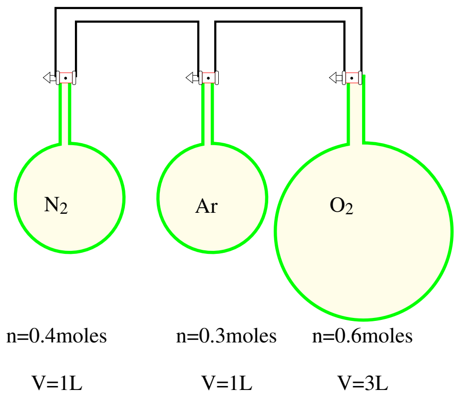
7.17 A tank contains Ne gas at 700 mmHg, Ar at 2 atm, and Kr at 700 torr. What is the total pressure of the mixture in atm?

7.18 An anesthetic consist of a mixture of cyclopropane gas and oxygen gas. If the mixture has a total pressure of 2 atm and the partial pressure of cyclopropane is 0.5atm, what is the partial pressure of O2?

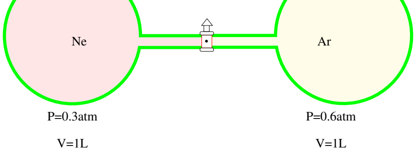
7.19 The atmospheric pressure on a hot day is 780 mmHg. Given that the air is made of 78% of nitrogen and 22% of oxygen, calculate the partial pressure of each gas in the air.

7.20 The atmospheric pressure on a hot day is 790 mmHg. Given that the air is made of 78% of nitrogen and 22% of oxygen, calculate the partial pressure of each gas in the air.

7.21 Consider the set up below with three different gases in three different closed containers at 300K. Assuming that the connecting tubes have zero volume, once the flasks are connected, calculate: (a) The partial pressure of each gas in the mixture (b) The total gas pressure



7.22 Consider the set up presented below, where the connecting tubes have negligible volume. Calculate the partial pressure of each gas after the connection between the flasks is open.



7.23 Phosphorus reacts with oxygen gas to produce tetraphosphorus decaoxide according to the following equation:

$$\underset{}{P_4(s)} + \underset{2L}{5O_2(g)} \longrightarrow \underset{x\ mol}{P_4O_{10}(g)}$$

Calculate the number of moles of phosphorus that react with 2L of oxygen at STP conditions.

7.24 For the following reaction, calculate the unknown x at STP conditions:

$$C_3H_8(g) + \underset{5L}{5O_2(g)} \longrightarrow \underset{x\ L}{3CO_2(g)} + 4\,H_2O_{(l)}$$

7.25 For the following reaction, calculate the unknown x at STP conditions:

$$\underset{3L}{5C_3H_8(g)} + 5O_2(g) \longrightarrow \underset{x\ L}{3CO_2(g)} + 4\,H_2O_{(l)}$$

7.26 For the following reaction, calculate the unknown x at STP conditions:

$$\underset{6moles}{5C_3H_8(g)} + 5O_2(g) \longrightarrow \underset{x\ L}{3CO_2(g)} + 4\,H_2O_{(l)}$$

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Ch. 8. Solutions

THE most common reactions happen in solution and involve cations and anions. Think, for example, when you add sugar to your coffee or how metal rusts when it gets wet. The first example involves a dissolution reaction, whereas the second is a redox reaction in which electrons are exchanged. Overall there are three main different types of reactions happening in solutions. It is not only important to be able to differentiate these reactions, but it is also important to understand the nuances between acid-base reactions, precipitation reactions, and redox reactions. It is also critical to understand the properties of solutions and learn how to quantify and compare the amount of solute in a solution. On one hand, most of you will be surprised to know that water does not conduct electricity. This is because pure water is a weak electrolyte. On the other hand, the importance of electrolytes is well-known among the sports community. If you have ever played a sport, you have probably chugged a sports drink. These are electrolyte solutions. However, few know the specifics of their function. Electrolytes are salts that conduct electricity in water by separating them into positive and negative ions. To understand the properties of solutions it is important to identify and differentiate the different types of electrolytes.

8.1 Solutions and composition

Solutions are homogeneous mixtures of two components. The state of the matter of both components of the mixture or their polarity affects the formation of a solution. For example, a solution will not result from mixing oil and water as they have different polar characteristics and it will form from mixing table salt and water as both are polar chemicals. At the same time, the more solute you add to a solution the more concentrated the solution will be. This section covers polarity and the composition of solutions.

What makes a solution? Solutions are homogeneous mixtures of a solute and a solvent (see Figure 8.1). Homogeneous means that if you look at the mixture you will not be able to differentiate both components and you will only see it as a whole. In a solution, the solute is the component of the mixture in less amount, whereas the solvent is the component in a larger amount. Think about mixing sugar with water. Sugar is sweet and water tasteless. When you mix both, you form a solution of sugar (solute) in water (solvent) and you will not see sugar in the solution as it is dissolved. In this particular example, sugar will be the solute in the solution, as the sugar is in less amount than water. Is important to remember that a solution is a result or mixing a solute and a solvent:

$$\boxed{\text{Solution} = \text{Solute} + \text{Solvent}}$$

Sample Problem 80

Identify the solute and the solvent in the following mixtures: (a) 10g of H_2O(l) mixed with 3g of KCl(s); (b) 10g Cu(s) mixed with 3g Zn(s)

SOLUTION

(a) Potassium chloride is in less amount and hence will be the solute. Water, in larger quantity, will be the solvent. This is an example of an aquose-based solution. (b) Zn is in less amount and hence will be the solute, and Cu the solvent. This is a solid solution.

STUDY CHECK

Identify solute and solvent in the following mixture: (a) 10g of H_2O(l) and 20g of CH_3OH(l); (b) 1g of I_2(s) and 100g of CH_3CH_2OH(l)

`Types of solutions` You can prepare different types of solutions by mixing a solid and a liquid, like when you mix sugar and water, or salt and water. You can create solutions as well by mixing two liquids or two solids. Examples are vinegar–a liquid solution of acetic acid (liquid) in water (liquid)–or steel– a solid solution that contains iron and carbon, both solids.

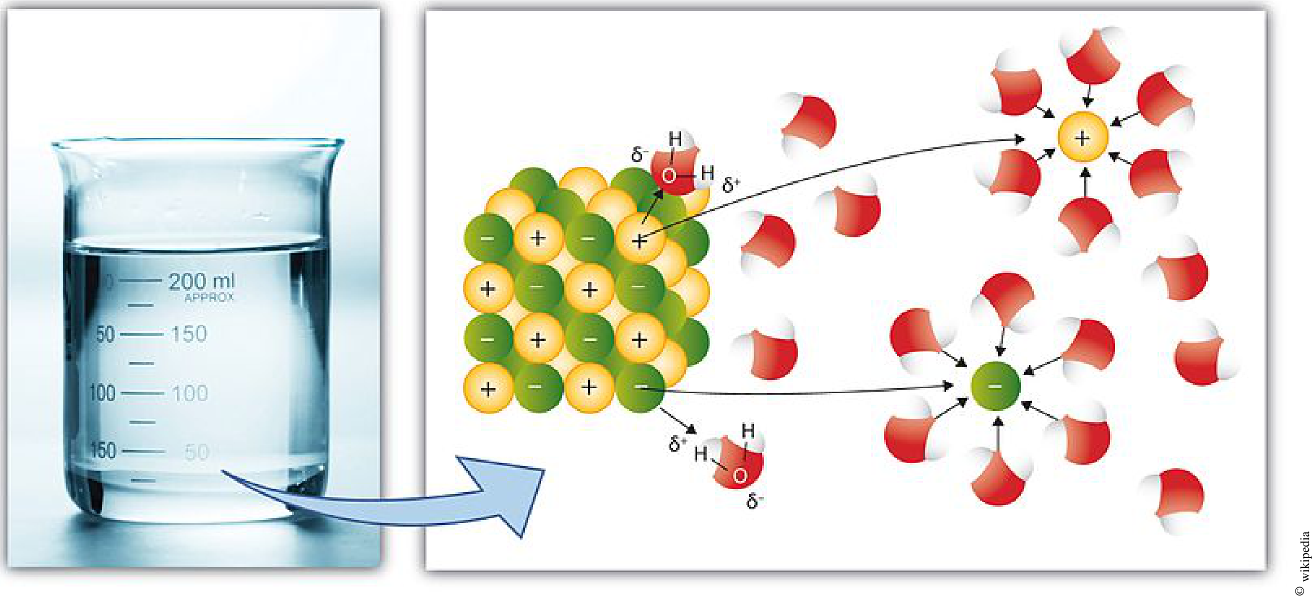


Figure 8.1 A solution results from dissolving a solute into a solvent

`Empirical rules of polarity` The affinity between two chemicals is related to a concept called polarity. Molecules contain electrons and depending on the electron distribution within the molecules, molecules can be polar or non-polar. Molecules with an even electron distribution are non-polar as they have no permanent dipole moment. An example of this is H_2 molecule. Differently, HF is a polar molecule, as F concentrates more on the electron density of the molecule than H. The polar nature of substances–with a permanent dipole moment–is related to miscibility and molecules with similar polar character will mingle and mix creating a single visible phase (see Table **??**). As an example, water (H_2O, polar) and methanol (CH_3OH, polar) will mix together. Differently, water (polar) and oil (non-polar) are immiscible due to their different polar nature and they will not mix. Even if the rules or polarity are based on the nature and structure of the molecule, one can use very simple empirical rules to classify molecules as polar or non-polar. These rules work in general well for the case of diatomic and very large molecules:

1. **Rule one:** Diatomic molecules made of the same element (e.g. H_2) are non-polar.
2. **Rul two:** Diatomic molecules made of different elements (e.g. HI) are polar.
3. **Rul three:**Poliatomic molecules (with more than four atoms) made of C and H (e.g. CH_4) are in general non-polar.
4. **Rul four:** Poliatomic molecules (with more than four atoms) containing C, H, and a different atom (e.g. CH_3F) are in general polar.

Sample Problem 81

Classify the following molecules as polar or nonpolar: H_2, HCl, CH_3CH_3, and CH_3CH_2Cl.

SOLUTION

H_2 is a non-polar molecule, being a diatomic molecule containing two atoms of the same element. Differently HCl is polar. CH_3CH_3 is a non-polar poliatomic molecule made of C and H atoms, whereas CH_3CH_2Cl is polar.

STUDY CHECK

Classify the following molecules as polar or nonpolar: HF, Cl_2, C_2H_4, and C_2H_3Cl.

Mixing and polarity A solution is formed when both the solute and the solvent mix. However, they will only mix if they have the same polarity. As an example, water (H_2O) is a polar molecule, and methanol (CH_3-OH) is too. Hence they will both mix and form a solution. If the elements of a mixture have different polarity they will not mix. An example is benzene (C_6H_6, nonpolar) and water, or for example oil (nonpolar) and water (polar).

Table 8.1 Polarity and mixing

Solvent	Solute	Mixing?
Polar	Polar	Yes
Polar	Nonpolar	No
Nonpolar	polar	No
Nonpolar	Nonpolar	Yes

Sample Problem 82

Use polarity arguments to indicate if the following substances will mix: (a) $H_2O_{(g)}$ and $CH_{4(g)}$; (b) $H_2O_{(g)}$ and $HCl_{(g)}$

SOLUTION

(a) Water and methane (CH_4) will not mix, as water is a polar molecule and CH_4 (methane) is nonpolar. (b) They will mix as HCl is a polar molecule and so is water.

STUDY CHECK

Use polarity arguments to indicate if the following substances will mix: (a) $H_2O_{(l)}$ and $CH_3Cl_{(l)}$; (b) $CH_3Cl_{(l)}$ and $CCl_{4(l)}$

8.2 Concentration of solutions

The concentration of a solution refers to the amount of solute with respect to the amount of solution. The larger concentration the larger the number of solute particles with respect to the particles of solvent. Concentration is one of the most important properties of a solution as it affects the physical properties of a solution such as the freezing and boiling point. There are many different concentration units, such as molarity, mass percent concentration, or volume percent concentration. All these different units overall express the ratio between the particles–mass or volume–of solute and solvent.

`Mass percent concentration` The mass percent (% m/m) is the amount of solute in grams per gram of solution in percent form

$$\boxed{\%m/m = \frac{\text{g of solute}}{\text{g of solution}} \times 100}$$

Sample Problem 83

A NaCl solution is prepared by mixing 4g of NaCl with 50g of H_2O. Calculate the percent (m/m) of the solution.

SOLUTION

We need the grams of solute and the grams of solution. The grams of solute are given (4g of NaCl), whereas the grams of solution result from adding the grams of solvent and solute: 54 g of solution. Using the formula for the percent (m/m), he have:

$$m/m = \frac{\text{4 g of solute}}{\text{54 g of solution}} \times 100 = 7.4\%$$

This means that by mixing 4g of NaCl with 50g of H_2O you prepare a 7.4% (m/m) solution.

STUDY CHECK

A KCl solution is prepared by mixing 5g of KCl with 200g of H_2O. Calculate the percent (m/m) of the solution.

`Volume percent concentration` The volume percent concentration (% v/v) is the volume of solute per volume of solution in percent form

$$\boxed{\%v/v = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100}$$

`Mass/volume percent concentration` The mass/volume percent concentration (% m/v) is the mass of solute per mL of solution in percent form.

$$\boxed{\%m/v = \frac{\text{g of solute}}{\text{mL of solution}} \times 100}$$

`Molarity concentration` The molarity (M) is the moles of solute per L of solution.

$$\boxed{M = \frac{\text{moles of solute}}{\text{L of solution}}}$$

Sample Problem 84

A NaCl solution is prepared by mixing 4g of NaCl (MW=58.4g/mol) with 50 g of water until a final volume of 52mL of solution. Calculate: (a) the mass percent (m/m) concentration; (b) the molarity.

SOLUTION

(a) to calculate the mass percent (m/m) we just need the grams of solute and the grams of solution–that is four plus fifty. Both numbers are already given:

$$m/m = \frac{\text{g of solute}}{\text{g of solution}} \times 100 = \frac{\text{4 g of solute}}{\text{54 g of solution}} \times 100 = 9.2\%$$

(b) To calculate molarity we need the moles of solute and the liters of solution. We have the mL of solution, that can be converted to L: 52mL= $5.2 \times 10^{-2} L$. To calculate the moles of solute, we will use the grams of solute and the molar mass to convert this value into moles: $4g/58.4g/mol$ =0.068moles. Plugging all values into the molarity formula:

$$M = \frac{\text{moles of solute}}{\text{L of solution}} = \frac{\text{0.068 moles of solute}}{5.2 \times 10^{-2}\text{L of solution}} = 1.31M$$

STUDY CHECK

(a) A solution is prepared by mixing 8g of NaCl (MW=74g/mol) with 250mL of H_2O. Calculate the molarity; (b) A KCl solution is prepared by mixing 45g of KCl with 200g of H_2O. Calculate the percent (m/m) of the solution.

Concentration units as conversion factors Each of the different concentration units–molarity, mass percent, volume percent, mass/volume percent–can be used in a conversion factor form. For example, if the molarity of a solution is 3M, this means that in the solution there are 3 moles of solute in every liter of solution.

$$\boxed{\text{3M} \quad \text{or} \quad \frac{\text{3 mol of solute}}{\text{1 L of solution}} \quad \text{or} \quad \frac{\text{1 L of solution}}{\text{3 mol of solute}}}$$

Similarly, if the mass percent of a solution is 5% this means that there are 5 grams of solute for every 100 grams of solution. We often use concentration units as conversion factors when we need to transform between one unit on top (bottom) of the conversion factor and the unit on the bottom (top).

Sample Problem 85

How much volume of a 4M solution do you need to provide 5 moles of solute.

SOLUTION

We will use the conversion factor of Molarity using the volume on top and the moles on the bottom in order to cancel the units:

$$5\,\cancel{\text{moles of solute}} \times \frac{\text{1 L of solution}}{4\,\cancel{\text{moles of solute}}} = \text{1.25L}$$

This means that 1.25L of a 4M solution will provide 5 moles of solute.

STUDY CHECK

How many grams of a 6% (m/m) solution do you need to provide 5 grams of solute.

Molarity calculation involving grams As molarity is just the ratio between the moles of solute and the litters of the solution, we can use this property to calculate the mass of solute contained in a certain volume of solution. We will proceed by first calculating the number of moles in that volume to then convert moles into grams.

8.3 Dilution

Dilution is the process of preparing a diluted solution from a more concentrated solution. Solutions are often stored in a stock room in concentrated form. These stocks should be diluted before use. To dilute a solution we need to take a certain amount of the concentrated solution and add water. When adding water, the number of moles of solute does not change, and the concentration always decreases. We have a concentrated solution (c_1) and we need to prepare a certain volume (V_2) of a more diluted solution (c_2). The question is how much volume of the concentrated solution (V_1) we need to take. To answer this we should use the following formula:

$$\boxed{c_1 \cdot V_1 = c_2 \cdot V_2}$$

Sample Problem 86

How many liters of a 3M NaCl solution are required to prepare 2L of a 1M diluted NaCl solution.

SOLUTION

We have a concentrated solution of 3M molarity and we want to prepare a more dilute solution. In particular 2L of a 1M. Hence: $c_1 = 3$ and $c_2 = 1M$ and $V_2 = 2L$. Using the dilution formula:

$$3M \cdot V_1 = 1M \cdot 2L$$

Solving for V_1 we have a volume of 0.66L.

STUDY CHECK

How many liters of a 5M NaCl solution are required to prepare 3L of a 3M diluted NaCl solution.

8.4 Electrolytes and insoluble compounds

On one hand, electrolytes are compounds that conduct electricity once dissolved in water. Differently, nonelectrolytes are compounds that do not conduct electricity once dissolved in water. On the other hand, insoluble compounds are not soluble in water, whereas soluble compounds can be dissolved in water. This section covers the properties of electrolytes and insoluble (and soluble) compounds. At the end of this section, you should be able to classify a chemical in terms of its electrolyte type and solubility character.

Solubility formula Solubility (s) is the grams of a solute per 100 g of solvent:

$$\boxed{s = \frac{\text{g of solute}}{\text{100 g of solvent}}}$$

A saturated solution can be achieved when you fit the maximum amount of solute in the solvent. If you continue adding solute to a saturated solution it will precipitate and a solid will form.

Soluble and insoluble salts Soluble compounds dissolve in water, whereas insoluble compounds do not. For example, barium chromate ($BaCrO_{4(s)}$) is an insoluble salt. How do we know that? Table 8.2 will help you predict the solubility of a salt. To do this, you need to start by assessing the right ion (the anion, $CrO_4{}^{2-}$) located in the left column of Table 8.2. After that, you need to assess the left ion (the cation, Ba^{2+}) located in the right column. If you follow this, you will see that chromate is insoluble and barium is not part of any exception. Let us predict for example the soluble/insoluble nature of $CaSO_4$, calcium sulfate. We start by looking for $SO_4{}^{2-}$ in the left column to find out is soluble. Next, we continue in the same line as $SO_4{}^{2-}$ and look for the ion in the left Ca^{2+}. In conclusion, even when $SO_4{}^{2-}$ is soluble, when combined with Ca^{2+}, we have that $CaSO_4$ is insoluble, and overall $CaSO_{4(s)}$ is insoluble.

Sample Problem 87

Predict the soluble/insoluble nature of the following compounds: (a) K_2CO_3, (b) $NaNO_3$ and (c) $Ca(OH)_2$.

SOLUTION

(a) $K_2CO_{3(aq)}$ is soluble, as $CO_3{}^{-2}$ is insoluble but when combined with K^+ the salt becomes soluble. (b) All nitrates are soluble without exceptions. (c) $Ca(OH)_{2(aq)}$ is soluble.

STUDY CHECK

Predict the soluble/insoluble nature of the following compounds: (a) Li_3PO_4 (b) Na_2S (c) AgCl

Table 8.2 Soluble and insoluble compounds

Ions that form *soluble* compounds...	... **except** when combined with
Group I ions (Na^+, Li^+, K^+, etc)	no exceptions
Ammonium ($NH_4{}^+$)	no exceptions
Nitrate ($NO_3{}^-$)	no exceptions
Acetate (CH_3COO^-)	no exceptions
Hydrogen carbonate ($HCO_3{}^-$)	no exceptions
Chlorate ($ClO_3{}^-$)	no exceptions
Halide (F^-, Cl^-, Br^-, I^-)	Pb^{2+}, Ag^+ and $Hg_2{}^{2+}$
Sulfate ($SO_4{}^{2-}$)	Ag^+, Ca^{2+}, Sr^{2+}, Ba^{2+}, $Hg_2{}^{2+}$ and Pb^{2+}
Ions that form *insoluble* compounds...	**... except when combined with**
Carbonates ($CO_3{}^{2-}$)	group I ions (Na^+, Li^+, K^+, etc) or ammonium ($NH_4{}^+$)
Chromates ($CrO_4{}^{2-}$)	group I ions (Na^+, Li^+, K^+, etc) or Ca^{2+}, Mg^{2+}or ammonium ($NH_4{}^+$)
Phosphates ($PO_4{}^{3-}$)	group I ions (Na^+, Li^+, K^+, etc) or ammonium ($NH_4{}^+$)
Sulfides (S^{2-})	group I ions (Na^+, Li^+, K^+, etc) or ammonium ($NH_4{}^+$)
Hydroxides (OH^-)	group I ions (Na^+, Li^+, K^+, etc) or Ca^{2+}, Mg^{2+}, Sr^{2+}or ammonium ($NH_4{}^+$)

`Strong electrolytes` Strong electrolytes completely dissociate in water. Hence, in a solution of a strong electrolyte, you will only have ions and never molecules. Strong electrolytes are typically ionic compounds such as $MgCl_2$ or NaCl (table salt). We represent the dissociation of a strong electrolyte with a single arrow, meaning that the reaction proceeds to completion, and for the example below, in the solution, we will only have ions ($Mg^{2+}_{(aq)} + 2\,Cl^-_{(aq)}$) and not molecules ($MgCl_{2(s)}$):

$$MgCl_{2(s)} \xrightarrow{H_2O} Mg^{2+}_{(aq)} + 2\,Cl^-_{(aq)}.$$

`Weak electrolytes` Weak electrolytes partially dissociate in water, and this is indicated using a chemical reaction with a double arrow. Hence in a solution of a weak electrolyte, you will have ions as well as molecules at the same time. Examples of weak electrolytes are hydrofluoric acid, water, ammonia, or acetic acid. The dissociation of hydrochloric acid (HF) proceeds as:

$$HF_{(l)} \underset{}{\overset{H_2O}{\rightleftharpoons}} H^+_{(aq)} + F^-_{(aq)}$$

Acetic acid (CH_3COOH) is an important weak electrolyte and its dissociation proceeds somehow in a peculiar way:

$$CH_3COOH_{(l)} \overset{H_2O}{\rightleftharpoons} CH_3COO^-_{(aq)} + H^+_{(aq)}$$

`Nonelectrolytes` Nonelectrolytes do not dissociate in water. Hence a solution of a nonelectrolyte will only contain molecules and not ions. Examples of nonelectrolytes are carbon-based chemicals such as methanol, ethanol, urea, or sucrose. The dissociation of urea for example CH_4N_2O proceeds as:

$$CH_4N_2O_{(s)} \xrightarrow{H_2O} CH_4N_2O_{(aq)}$$

`Identify the electrolyte character of a chemical` You can use Table 8.3 to identify the electrolyte character of a chemical. Ionic compounds are in general strong electrolytes, and most acids are as well. There are four important weak electrolytes: water, acetic acid, ammonia, and hydrofluoric acid. Covalent compounds are in general nonelectrolytes. Organic compounds, compounds based on carbon atoms (e.g. $C_{12}H_{22}O_{11}$) are in general nonelectrolytes.

Table 8.3 Different types of electrolytess

Electrolyte Type	Dissociation	Particles in solution	Examples
Strong	Fully	Mostly ions	Ionic Compounds and most acids and bases (hydroxides): NaCl, NaOH, HCl, $MgCl_2$, H_2SO_4, etc
Weak	Partially	Ions & molecules	NH_3, CH_3COOH (acetic acid), HF, H_2O
Nonelectrolytes	No	molecules	Most covalent compounds: CH_3OH (methanol), CH_3CH_2OH (ethanol), $C_{12}H_{22}O_{11}$ (sucrose), CH_4NO_2 (urea)

Sample Problem 88

For the following chemicals indicate whether you will have in the solution (a) only ions, (b) ions and some molecules, or (c) molecules: NH_3, KOH, and $C_{12}H_{22}O_{11}$.

SOLUTION

Ammonia (NH_3) is a weak electrolyte and a solution of ammonia will contain ions and well as ammonia molecules. Differently KOH is a strong electrolyte and in solution you would find only ions (K^+ and OH^-). Sucrose ($C_{12}H_{22}O_{11}$) is a nonelectrolyte and in solution you will find molecules.

STUDY CHECK

For the following chemicals indicate whether you will have in the solution only ions, ions and some molecules, or molecules: (a) H_2SO_4, HNO_3 and CH_3OH.

`Breaking down electrolytes into ions` Electrolytes–in particular strong electrolytes–dissociate producing ions. This way, a solution of for example NaCl does not contain NaCl molecules but $Na^+_{(aq)}$ cations and $Cl^-_{(aq)}$ anions. Hence it is important to correctly break down electrolytes into ions. To do this, you need to revert the combination of ions that produce a given chemical while making sure the charges are balanced. For example, let us beak magnesium chloride $MgCl_{2(aq)}$ into ions. This is a strong electrolyte formed by magnesium cations and chloride anions. The valence of magnesium is +II and the valence of chlorine is -I. The $MgCl_2$ formula also tells us we have one magnesium and two chlorines. The overall process is:

$$MgCl_{2(aq)} \longrightarrow Mg^{2+}_{(aq)} + 2\,Cl^{-}_{(aq)}$$

Another example, magnesium nitrate $Mg(NO_3)_2$. This strong electrolyte–as this is an ionic salt–is made of magnesium with valence +II and nitrate with valence -I. The formula indicated we have one $Mg^{2+}_{(aq)}$ and two ${NO_3}^{-}{}_{(aq)}$. Hence:

$$Mg(NO_3)_{2(aq)} \longrightarrow Mg^{2+}_{(aq)} + 2\,{NO_3}^{-}{}_{(aq)}$$

Sample Problem 89

Break down the following chemicals into ions, if possible:

Chemical	Particles in solution
$K_2CrO_{4(aq)}$	
$Ba(NO_3)_{2(aq)}$	
$BaCrO_{4(s)}$	
$KNO_{3(aq)}$	

SOLUTION

We can only break down into ions ionic compounds and oxosalts that are not solid. From the list of chemicals in the example, we will not be able to break down $BaCrO_{4(s)}$ into ions as it is a solid. From the other chemicals, $K_2CrO_{4(aq)}$ is named potassium chromate and contains $2K^{+}_{(aq)}$ and ${CrO_4}^{2-}{}_{(aq)}$ ions. Barium nitrate–$Ba(NO_3)_{2(aq)}$–will produce $Ba^{2+}_{(aq)}$ and $2{NO_3}^{-}{}_{(aq)}$. Finally, potassium nitrate– $KNO_{3(aq)}$–will produce $K^{+}_{(aq)}$ and ${NO_3}^{-}{}_{(aq)}$. In the table:

Chemical	Particles in solution
$K_2CrO_{4(aq)}$	$2K^{+}_{(aq)} + {CrO_4}^{2-}{}_{(aq)}$
$Ba(NO_3)_{2(aq)}$	$Ba^{2+}_{(aq)} + 2{NO_3}^{-}{}_{(aq)}$
$BaCrO_{4(s)}$	$BaCrO_{4(s)}$
$KNO_{3(aq)}$	$K^{+}_{(aq)} + {NO_3}^{-}{}_{(aq)}$

STUDY CHECK

Break down the following chemicals into ions, if possible: $H_2O_{(l)}$, $NH_{3(l)}$, $AgNO_{3(aq)}$.

8.5 An introduction to reactions in solution

There are three different reactions in solution: acid-base reactions, precipitation reactions, and redox reactions. The key to identifying acid-base reactions is in the reactants, as an acid-base reaction results from the reaction between an acid and a base. Precipitation reactions are reactions that produce a precipitate. Hence, the key to identifying a precipitation reaction is in the products. Precipitation reactions always contain a solid as a product. Redox reactions contain two elements with different redox numbers in the reactants and products. The key to identifying redox reactions is to be able to spot elements with different oxidation states, for example: Cu and Cu^{2+} or H^+ and H_2. In the following, we will describe more about the three different types of reactions in solution. The goal of this section is for you to be able to identify each type.

`Acid-base reactions` Acid-base reactions result from the reaction of an acid with a base. Both produce water and another chemical. An example is:

$$HBr_{(aq)} + KOH_{(aq)} \longrightarrow KBr_{(aq)} + H_2O_{(l)} \text{ (acid-base reaction)}$$

Hydrobromic acid (HBr) is an acid and potassium hydroxide (KOH) is a base. The result of an acid-base reaction is always water and an ionic compound, in this case KBr.

`Precipitation reactions` Precipitation reactions result in an insoluble chemical, that is, results in a solid chemical. An example would be:

$$K_2CrO_{4(aq)} + Ba(NO_3)_{2(aq)} \longrightarrow BaCrO_{4(s)}\downarrow + 2\,KNO_{3(aq)} \text{ (precipitation reaction)}$$

The chemical $BaCrO_{4(s)}$ is a solid that precipitates in the solution, hence the name of the type of reaction. The symbol on $BaCrO_{4(s)}\downarrow$ represents the precipitation process. The solubility of a given solute such as $BaCrO_{4(s)}$ is the amount of solute (in grams) that can be dissolved in a given mass of solvent (in particular 100 g of solvent). A solute with a low solubility will be hard to dissolve. Think about cacao and water. The solubility of cacao is low and hence by simply adding cacao powder to water, you will not be able to make a solution. However, solubility depends on the solute and solvent combination, but also on the temperature, and by warming up a solvent you can increase solubility and fit more solute in the same amount of solvent. This section covers different aspects of solubility.

`Redox reactions` Redox reactions are different than acid-base or precipitation reactions. They contain the same chemical element in two different states resulting from the loss or win of electrons. Look for the example:

$$2\,Al_{(s)} + 3\,Cu^{+2}_{(aq)} \longrightarrow 2\,Al^{+3}_{(aq)} + 3\,Cu_{(s)} \text{(redox reaction)}$$

We have that neither $Al_{(s)}$ or ${Cu_{(aq)}}^{+2}$ are an acid or a base, therefore this is not an acid-base reaction. Also, there is no insoluble product, hence this is not a precipitation reaction. Indeed, this is a redox reaction, as we have Al in two different states: as metallic $Al_{(s)}$ and as ionic ${Al_{(aq)}}^{+3}$, which result from the loss of three electrons. Therefore in a redox reaction, there are always elements in the chemicals that lose electrons. In redox reactions, there is also an element that wins electrons. For example, $Cu_{(s)}$ and ${Cu_{(aq)}}^{+2}$ have different redox number. In particular, ${Cu_{(aq)}}^{+2}$ is the result of removing three electrons from $Cu_{(s)}$. At this point, we have that this reaction is redox as it contains an element that gains electrons and an element that loses electrons. Sometimes, the redox state of the elements is not that obvious. Look at this example:

$$Fe_{(s)} + CuSO_{4(aq)} \longrightarrow FeSO_{4(aq)} + Cu_{(s)} \text{(redox reaction)}$$

This is a redox reaction as you can find iron and copper in two states, metallic and also ionic. Therefore, these two metals have two different redox numbers in the reaction.

Sample Problem 90

Classify the following reactions as acid-base, redox or precipitation.

(a) $Fe_{(s)} + Cu^{+2}_{(aq)} \longrightarrow Fe^{+2}_{(aq)} + Cu_{(s)}$

(b) $AgNO_{3(aq)} + NaCl_{(aq)} \longrightarrow AgCl_{(s)} + NaNO_{3(aq)}$

(c) $2\,HCl_{(aq)} + Ca(OH)_{2(aq)} \longrightarrow CaCl_{2(aq)} + 2\,H_2O_{(l)}$

SOLUTION

The first reaction is a redox reaction. This is because we can find two different oxidation states for Cu and also for Fe. That means one of these elements lost electrons and the other won electrons. The second reaction is a precipitation reaction as it produces a solid. The last reaction is an acid base, as the reactants are an acid and a base.

STUDY CHECK

Classify the following reactions as acid-base, redox or precipitation.

(a) $HNO_{2(aq)} + NaOH_{(aq)} \longrightarrow NaNO_{2(aq)} + H_2O_{(l)}$

(b) $2\,Na_{(s)} + Cl_{2(g)} \longrightarrow 2\,NaCl_{(s)}$

(c) $MgCl_{2(aq)} + 2\,AgNO_{3(aq)} \longrightarrow 2\,AgCl_{(s)} + Mg(NO_3)_{2(aq)}$

8.6 Redox reactions

Redox reactions are different than acid-base or precipitation reactions. They contain the same chemical element in two different states resulting from the loss or win of electrons. For example:

$$2\,Al_{(s)} + 3\,Cu^{2+}_{(aq)} \longrightarrow 2\,Al^{3+}_{(aq)} + 3\,Cu_{(s)}\ \text{(redox reaction)}$$

We have that neither $Al_{(s)}$ or $Cu^{2+}_{(aq)}$ are an acid or a base. Also, there is no product precipitate. Hence, this reaction is not an acid-base reaction or a precipitation reaction. This is a redox reaction, as we have Al in two different states: as metallic $Al_{(s)}$ and as ionic $Al^{3+}_{(aq)}$, which result from the loss of an electron. In redox reaction, there is always elements in the chemicals that lose electrons and chemicals winning electrons.

`Oxidation state or redox number` The redox number also called oxidation number or oxidation state helps compare elements that have lost (or gained) different charges. When comparing two elements with different redox numbers, the larger this number the more electrons the element has lost. Similarly, the smaller this number the more electrons the element has won.

`Rules to calculate redox numbers` We indicate redox numbers with a roman number on top of the element. For example, the redox number of manganese in this compound is $+7$: $\underline{Mn^{VII}}O_4{}^-$. The redox number can be noninteger. There are five rules to identify the redox number of an element.

- **Rule 1** Single atoms or elements have zero redox number. Examples are Na or H_2, both with redox zero.
- **Rule 2** Monoatomic ions have redox number equal to their charge. Examples are Na^+ or Cl^- with redox $+1$ and -1, respectively.
- **Rule 3** The redox number of fluorine is -1
- **Rule 4** The redox number of hydrogen on its covalent (e.g. H_2O) compounds is $+1$.

Rule 5 The redox number of oxygen in normal oxides (e.g. MgO) is normally -2, with the exception of peroxides (e.g. H_2O_2) in which is -1.

`Calculating the redox number` How do we calculate the redox number for example of manganese in this chemical: $\underline{Mn}O_4{}^-$, permanganate? To do this, we need to set up a formula so that the redox numbers of all elements in the molecule–taking into account the number of atoms in the molecule–are equal to the charge. In the case of permanganate, let us call x to the redox number of manganese. We know the redox of oxygen is -2 and we have four oxygens in the molecule. We also know the charge of the ion is -1. Therefore we have:

$$x + 4 \cdot (-2) = -1$$

If we solve for x we obtain the redox number of manganese of VII.

Sample Problem 91

Calculate the redox number of the elements underlined in the following molecules: (a) $K_2\underline{C}O_3$ and (b) $H_2\underline{C}O$.

SOLUTION

Let us set up the redox equation for the first compound, knowing that the redox of oxygen is -2 and potassium $+1$. The unknown variable x represents the redox number of the underlined element. We have:

$$2 \cdot (+1) + x + 3 \cdot (-2) = 0$$

Mind we have two potassium and three oxygens hence we need to time the redox of K by two and the redox of O by three. If we solve for x we obtain a redox number for carbon of IV. The redox equation for the second example is:

$$2 \cdot (+1) + x + (-2) = 0$$

Mind that according to the redox rules, the redox number of oxygen is $+1$. Solving for x we have a redox number of zero.

STUDY CHECK

Calculate the redox number of the elements underlined in the following molecules: (a) $\underline{Cr_2}O_7{}^{2-}$ and (b) $\underline{Cr_2}O_3$

`Redox means oxidation and reduction` By comparing the redox number of the same element in two different compounds we can figure out in what compound the element has lost or gained electrons. Look for example the case of $\underline{Cr^{VI}}_2O_7{}^{2-}$ and $\underline{Cr^{III}}_2O_3$. The same element in two different molecules has two different redox numbers. In the case of dichromate, the redox of Cr is VI, whereas in the case of chromium(III) oxide the redox of Cr is III. The larger the redox number the more oxidized is the element, and that means the element has lost electrons. The smaller the redox number the more reduced the element and that means it has gained electrons. If we compare both cases, we have that Cr in dichromate is oxidized–it lost electrons–and Cr in chromium(III) oxide is reduced–it gained electrons.

`Redox numbers in chemical reactions` The goal is to identify the element that undergoes oxidation and reduction in a chemical reaction. We can reach this goal by using the half-reaction method. Every redox reaction is composed of two processes, reduction, and oxidation. These two processes can be separated into two half-reactions so that the combination of both half-reactions leads to a balanced redox. Let us work on a simple unbalanced redox reaction:

$$Al_{(s)} + {Cu_{(aq)}}^{+2} \longrightarrow {Al_{(aq)}}^{+3} + Cu_{(s)}$$

Solid $Al_{(s)}$ on the reactant side has zero redox number, whereas ionic ${Al_{(aq)}}^{+3}$ on the product side has a redox number equal to III. Al has undergone oxidation as its redox number increases from zero to three. Al has lost three electrons. We can write the oxidation half-reaction:

$$Al_{(s)} \longrightarrow {Al_{(aq)}}^{+3} + 3\,e^- \text{ (oxidation)}$$

Mind that electrons have a negative charge and we add electrons to compensate for the charge of ${Al_{(aq)}}^{+3}$. Now let us compare the redox number of Cu. In the reactant side we have ${Cu_{(aq)}}^{+2}$ with redox of II. On the product side, we have $Cu_{(s)}$ with zero redox number. Cu has undergone reduction as its redox number has decreased. This means it has gained electrons, in particular, two electrons:

$${Cu_{(aq)}}^{+2} + 2\,e^- \longrightarrow Cu_{(s)} \text{ (reduction)}$$

`Balancing simple redox reactions` The goal here is to balance a redox chemical reaction by combining two half-reactions. In the example above the oxidation and reaction involving a different number of electrons. Hence to be able to add both redox we need to time each half-reaction by a number so that the number of electrons cancels out. As the first reaction involved three electrons and the second two, we will do:

$$
\begin{array}{lr}
2 \cdot \left(Al_{(s)} \longrightarrow {Al_{(aq)}}^{+3} + 3\,e^- \right) & \text{(oxidation)} \\
3 \cdot \left({Cu_{(aq)}}^{+2} + 2\,e^- \longrightarrow Cu_{(s)} \right) & \text{(reduction)} \\
\hline
 & + \\
2\,Al_{(s)} + 3\,{Cu_{(aq)}}^{+2} + \cancel{6\,e^-} \longrightarrow 2\,{Al_{(aq)}}^{+3} + 3\,Cu_{(s)} + \cancel{6\,e^-} & \text{(redox)}
\end{array}
$$

The overall balanced redox equation is:

Oxidation: $-3e^-$

$$2\ \overset{0}{Al}(s) + 3\ \overset{II}{Cu}^{2+}(aq) \longrightarrow 2\ \overset{III}{Al}^{3+}(aq) + 3\ \overset{0}{Cu}(s)$$

Reduction: $+2e^-$

CHAPTER 8

Solutions

8.1 A solution is prepared by mixing 4 g of $C_6H_{6(l)}$ and 10 g of $CCl_{4(l)}$. Indicate the true statement: (a) C_6H_6 is the solute (b) CCl_4 is the solute (c) Both chemicals do not mix (d) The mixture is not a solution

8.2 A solution is prepared by mixing 5 g of $Au_{(s)}$ and 2 g of $Cu_{(s)}$. Indicate the true statement: (a) Au is the solute (b) Cu is the solute (c) Both elements do not mix (d) The mixture is not a solution

8.3 Classify the following molecules as polar or non-polar: (a) $C_6H_{14(l)}$ (b) $CH_3CH_2CH_2OH_{(l)}$ (c) $C_5H_{10(l)}$ (d) $C_6H_5CH_{3(l)}$ (e) $CH_3CH_2OH_{(l)}$ (f) $C_6H_5NH_{2(l)}$

8.4 Classify the following molecules as polar or nonpolar: (a) $H_2O_{(l)}$ (b) $C_5H_{12(l)}$ (c) $CH_3COOH_{(l)}$ (d) $CH_3OH_{(l)}$

8.5 Which chemicals from the following list will mix with $H_2O_{(l)}$: (a) $NH_{3(l)}$ (b) $C_5H_{12(l)}$ (c) $C_6H_{14(l)}$

8.6 Which chemicals from the following list will mix with $C_5H_{12(l)}$: (a) $H_2O_{(l)}$ (b) $C_6H_{14(l)}$ (c) $CH_3COOH_{(l)}$

Electrolytes and insoluble compounds

8.7 Indicate whether solutions of the following chemicals will have ions (I), ions and molecules (I+M), or just molecules (M):

Chemical	I	I+M	M
NaCl			
HCl			
$CaCl_2$			

8.8 Indicate whether solutions of the following chemicals will have ions (I), ions and molecules (I+M), or just molecules (M):

Chemical	I	I+M	M
H_2O			
NO_2			
CO_2			

8.9 Indicate the soluble/insoluble character of the following compounds:

Chemical	Soluble	Insoluble
$AgNO_3$		
AgBr		
$CaCO_3$		
Na_2CO_3		

8.10 Indicate the soluble/insoluble character of the following compounds:

Chemical	Soluble	Insoluble
$NaCH_3COO$		
$NaHCO_3$		
Ag_2SO_4		
$NaCrO_4$		
CaS		

Concentration Units

8.11 Order the following solution from more to less concentrated: (a) A vanillin solution (1.3×10^{-1}M) (b) An adrenaline solution (1.1×10^{3}M) (c) Vinegar, an acetic acid solution (1.2×10^{0}M)

8.12 Order the following solution from more to less concentrated: (a) Solution A (4.3×10^{-5}M) (b) Solution B (6.1×10^{-3}M) (c) Solution C (6.4×10^{-4}M)

8.13 Sodium hydroxide NaOH, a very strong base, is a chemical used in drain cleaners. A drain cleaning solution is prepared by mixing 25g of NaOH in 250g of water. Calculate the mass percent of solute.

8.14 A solution is prepared by mixing 15g of KCl in 50g of water. Calculate the mass percent of solute.

8.15 Alcohol-hydroxide is a mixture of a base with an organic alcohol, employed to clean glass. An alcohol-hydroxide mixture is prepared by mixing 60g of NaOH with 500g of ethanol. Calculate: (a) the mass percent of solvent. (b) the mass percent of solute.

8.16 Vinegar is not a pure chemical, it is indeed a (m/m) 5% acetic acid solution. How many grams of acetic acid are there in 2g of vinegar.

8.17 Vanilla extract is a solution vanillin in ethanol. A vanilla solution is made by mixing 15 mL of pure vanillin and 50mL of ethanol. Calculate the Percent Volume/Volume.

8.18 A solution is made by mixing 2 mL of a solute and 100mL of solvent. Calculate the Percent Volume/Volume.

8.19 An HCl solution is prepared by mixing 4 moles of HCl with water until reaching a volume of 250mL. Calculate the molarity of the solution.

8.20 How many mL of a 3M KCl solution contains 0.06 moles of KCl.

8.21 How many mL of a 4M NaCl (MW=58g/mol) solution contains 5 grams of NaCl.

8.22 How many mL of a 1M NaCl (MW=58g/mol) solution contains 4 grams of NaCl.

8.23 How many grams of solute are there in 100mL of a 0.01M HNO_3 (MW=63g/mol) solution.

8.24 How many mL of a 0.001M $Ca(OH)_2$ (MW=74g/mol) solution can be prepared from 5 mg of $Ca(OH)_2$.

8.25 What is the final volume when 50mL of a 2M NaCl solution is diluted to a 1M.

8.26 What is the concentration of a solution prepared when 100mL a 4% HCl solution is diluted to a final volume of 500mL.

8.27 Describe how to prepare 50mL of a 0.5M H_2SO_4 solution, starting with a 1M stock H_2SO_4 solution.

8.28 Describe how to prepare 100mL of a 0.1M H_2SO_4 solution, starting with a 2M stock H_2SO_4 solution.

Precipitation and acid-base reactions

8.29 Classify the following reaction as acid-base, precipitation, or redox:

(a) $NaCl_{(aq)} + AgNO_{3\,(aq)} \longrightarrow NaNO_{3(aq)} + AgCl_{(s)}\downarrow$

(b) $H_2SO_{4(aq)} + 2\,NaOH_{(aq)} \longrightarrow 2\,H_2O_{(l)} + Na_2SO_{4(aq)}$

(c) $2\,Na_3PO_{4(aq)} + 3\,CaCl_{2(aq)} \longrightarrow 6\,NaCl_{(aq)} + Ca_3(PO_4)_{2(s)}\downarrow$

8.30 Classify the following reaction as acid-base, precipitation, or redox:

(a) $H_2SO_{4(aq)} + Ba(OH)_{2(aq)} \longrightarrow BaSO_{4(aq)} + 2\,H_2O_{(l)}$

(b) $Na_2CO_{3(aq)} + BaCl_{2(aq)} \longrightarrow BaCO_{3(s)} + 2\,NaCl_{(aq)}$

(c) $4\,Fe + 3\,O_2 \longrightarrow 2\,Fe_2O_3$

Redox

8.31 Balance the following redox reactions:

(a) $Fe_{(s)} + O_{2(g)} \longrightarrow Fe^{3+}_{(aq)} + O^{2-}_{(aq)}$

(b) $Cu_{(s)} + Ag^{+}_{(aq)} \longrightarrow Ag_{(s)} + Cu^{2+}_{(aq)}$

8.32 Balance the following redox reactions:

(a) $Cu^{+}_{(aq)} + Fe_{(s)} \longrightarrow Fe^{3+}_{(aq)} + Cu_{(s)}$

(b) $Ag_{(s)} + Zn^{2+}_{(aq)} \longrightarrow Ag^{+}_{(aq)} + Zn_{(s)}$

Ch. 9. Chemical equilibrium

REACTIONS rarely occur to completion. Once the products are formed they can follow a backward reaction to regenerate the reactants and at the same time, reactants regenerate produce products. Overall chemical reactions proceed until an equilibrium is reached. The equilibrium conditions determine how profitable a reaction can be as it describes how much of the products stay as products. This chapter covers the idea and basic principles of chemical equilibrium. We will describe the role of an equilibrium constant which gives insight into the mixtures of reactants and products in equilibrium. Also, we will discuss the different types of equilibrium constants in terms of molarity and pressure and the role of a reaction quotient giving insight into the direction in which a chemical reaction proceeds towards the equilibrium. Finally, we will cover Le Châtelier principle that describes once an equilibrium is altered, in which direction does a chemical reaction proceed to recover the equilibrium state.

9.1 Chemical Equilibrium

This section covers basic ideas about chemical equilibrium. First, we will introduce the concept of the forward and reverse reactions. Next, we will define the idea of equilibrium, based on the speed of the forward and reverse reactions.

Forward and reverse reactions In chemical reactions, reactants form products. We call this the *forward reaction*:

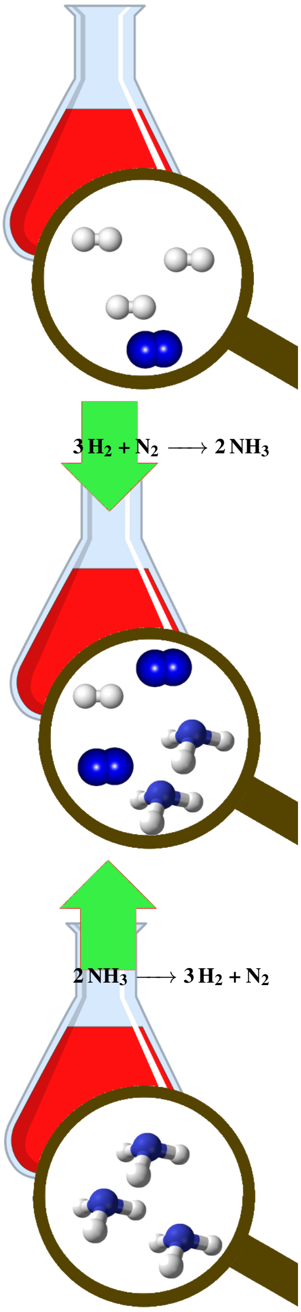
$$\text{reactants} \longrightarrow \text{products}$$

However, once the products form they can also generate reactants through the *reverse reaction*:

$$\text{reactants} \longleftarrow \text{products}$$

Equilibrium In chemical reactions, both the forward and the reverse reactions happen synchronously, so that when products form they also generate reactants. However, both the reverse and forward reactions have different speeds. In the beginning, the forward reaction proceeds at a faster pace than the reverse so that the reaction advances. Once products form, the reverse reaction will start speeding. A reaction reaches *chemical equilibrium* when both the forward and reverse reactions proceed at the same speed. Chemical reactions normally are written down with a double arrow that indicates equilibrium:

$$\text{reactants} \rightleftharpoons \text{products}$$



▼The forward reaction goes from reactants to products whereas the reverse reaction goes from products to reactants.

For example, the chemical process that forms ammonia from nitrogen and hydrogen can proceed using the forward and reverse reaction

$$3\,H_{2(g)} + N_{2(g)} \underset{\text{Reverse Reaction}}{\overset{\text{Forward Reaction}}{\rightleftharpoons}} 2\,NH_{3(g)}$$

`Equilibrium and concentration` It is just a matter of time before a reaction reaches equilibrium. Initially, the forward process normally proceeds at high speed as products start to form. This is because initially there is an abundance of reactants. Once the products start to accumulate the forward process will start happening. This is because at that point there is an increasing amount of product molecules that can go back to reactants. Eventually, the forward and reverse rates become equal as the reaction reaches equilibrium. When the equilibrium has been reached, reactants and products have very specific concentrations that depend on temperature. At the same time, a reaction in equilibrium contains a mixture of reactants and products.

Sample Problem 92

Write down the forward and reverse reactions for:

$$CH_4(g) + 2\,O_2(g) \rightleftharpoons CO_2(g) + 2\,H_2O(g)$$

SOLUTION

(a) The forward reaction is $CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(g)$ and the reverse $CH_4(g) + 2\,O_2(g) \longleftarrow CO_2(g) + 2\,H_2O(g)$

STUDY CHECK

Write down the forward and reverse reactions for the reaction:

$$CO(g) + O_2(g) \rightleftharpoons CO_2(g)$$

9.2 The equilibrium constant

Reactions in equilibrium are characterized by an equilibrium constant. This next section will show how to interpret and calculate the value of the equilibrium constant of a reaction.

`Equilibrium mixtures` Imagine you start a chemical reaction. Initially, products will form while reactants disappear. At equilibrium, you will have a mixture of reactants and products with the forward and reverse processes happening at the same rate. Would an equilibrium mixture contain more reactants or more products? The equilibrium constant helps predict just that.

Table 9.1 Different K_c values at 298K

Reaction	K_c	Equilibrium mixture
$2\,NH_{3(g)} \rightleftharpoons N_{2(g)} + 3\,H_{2(g)}$	17	Products > Reactants
$H_{2(g)} + I_{2(g)} \rightleftharpoons 2\,HI_{(g)}$	50	Products > Reactants
$2\,SO_{3(g)} \rightleftharpoons 2\,SO_{2(g)} + O_{2(g)}$	0.3	Products < Reactants
$H_2O_{(l)} \rightleftharpoons H_2O_{(g)}$	0.2	Products < Reactants

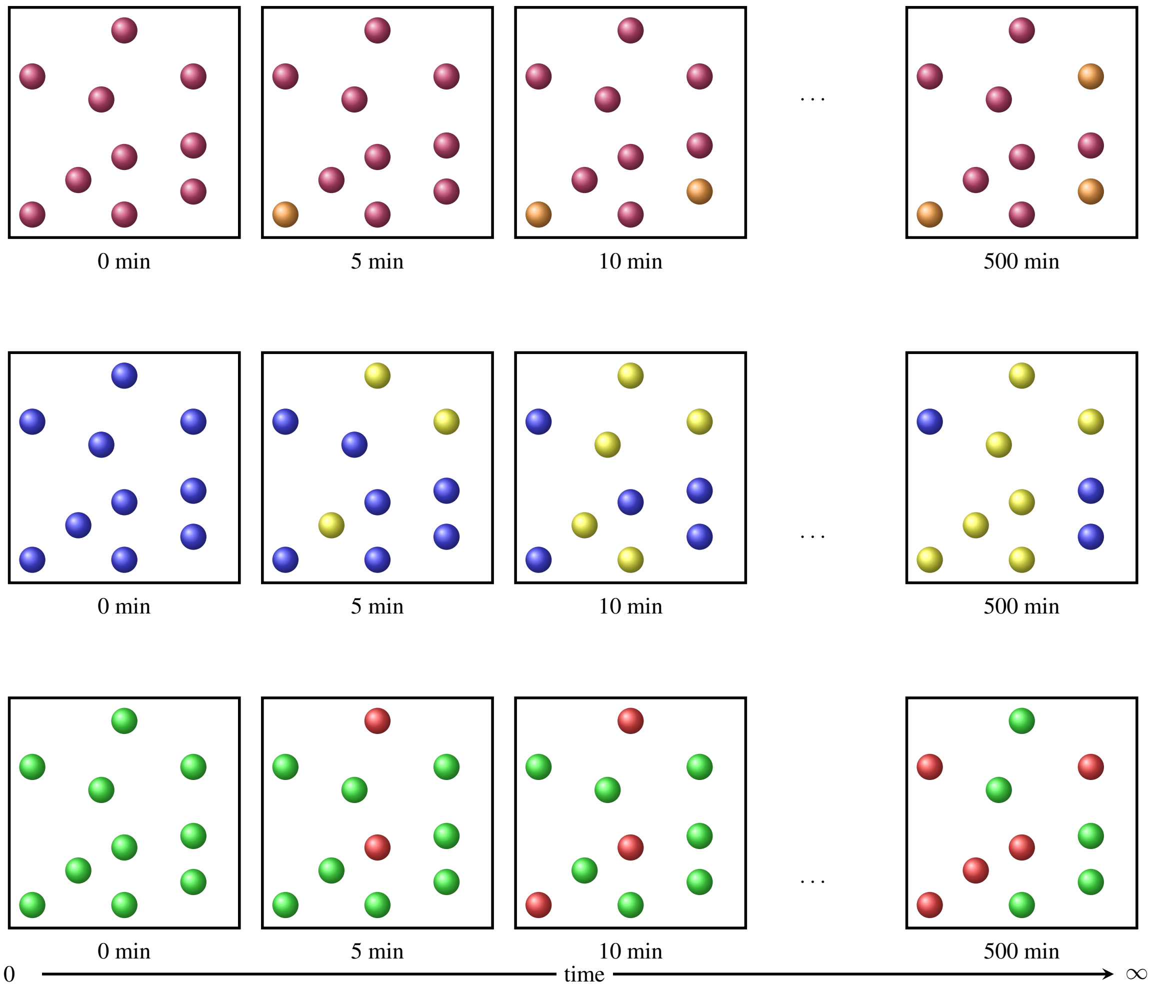


Figure 9.1 Three equilibrium situations

`The equilibrium constant of a reaction` The equilibrium constant associated with a reaction K_c indicates whether reactants or products are abundant at equilibrium. Each reaction has one K_c value that depends only on temperature and the subscript c represents concentration. On one hand, if K_c is larger than one there will be a larger concentration of products than reactants in the equilibrium mixture (see Figure 9.1). On the other hand, if K_c is smaller than one there will be a larger concentration of reactants than products in the equilibrium mixture. If K_c is close to one then both reactants and products will have the same concentration in the equilibrium mixture (see Table 9.1 for examples).

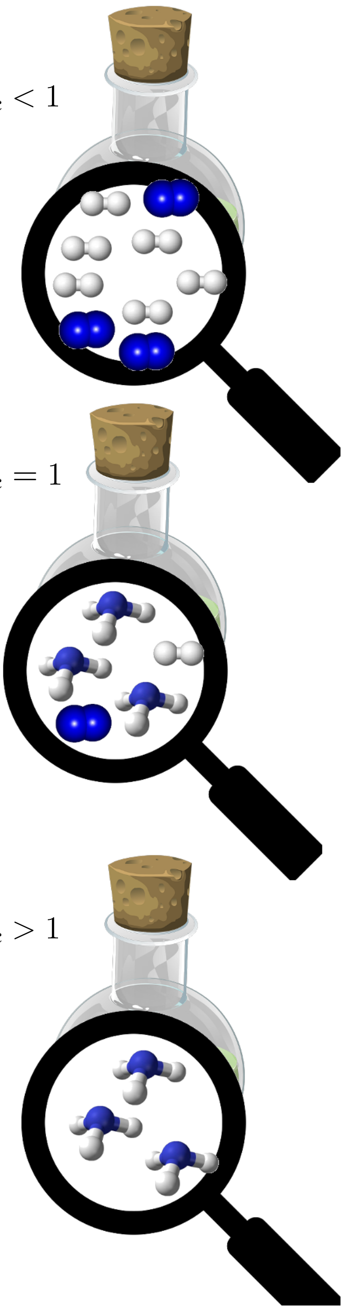
$K_c < 1$ $\quad$ $K_c \simeq 1$ $\quad$ $K_c > 1$

In terms of units, equilibrium constants are unitless numbers–a number without units. In other words, they have no unit and they are simply expressed as a number.

$K_c < 1$

$K_c = 1$

$K_c > 1$



▼The magnitude of the equilibrium constant indicates whether there is more products or reactants in an equilibrium mixture.

Sample Problem 93

The value of K_c for the following reaction at 300K is 4×10^{31}. Indicate whether the equilibrium mixture will contain mostly reactants, mostly products or both.

$$Cl_{2(g)} + H_{2(g)} \rightleftharpoons 2\,HCl_{(g)}$$

SOLUTION

As K_c is larger than one, an equilibrium mixture of Cl_2, H_2 and HCl will contain mainly products, that is will be mainly made of HCl.

STUDY CHECK

The value of K_c for $F_2(g) \rightleftharpoons 2\,F(g)$ at 500K is 7×10^{-13}. Indicate whether the equilibrium mixture will contain mostly reactants, mostly products or both.

`Equilibrium constant expression` Let's consider a general equilibrium reaction in which A and B react to form C and D

$$\underbrace{aA + bB}_{\text{Reactants}} \rightleftharpoons \underbrace{cC + dD}_{\text{Products}}$$

The stoichiometric coefficients of the reaction are a, b, c, and d. The expression for the equilibrium constant depends only on the concentration of the reactants and products. To refer to concentration we will use square brackets. For example, $[A]$ means the concentration of A. Hence, the expression of the equilibrium constant will be:

$$\boxed{K_c = \frac{[\text{Products}]}{[\text{Reactants}]} = \frac{[C]^c \cdot [D]^d}{[A]^a \cdot [B]^b}} \quad \text{Equilibrium constant} \tag{9.1}$$

Let us break down the expression of K_c. On top of the fraction, we have the equilibrium concentration of the products to the power of its coefficients. For example $[C]^c$ means the equilibrium concentration of C to the power of the coefficient c. On the bottom of the fraction, we have the concentration of the reactants to the power of its coefficients. All concentrations in K_c are timed.

product concentration

$$K_c = \frac{[C]^c \cdot [D]^d}{[A]^a \cdot [B]^b}$$

reactant coefficient

Let's focus on the reaction below:

$$CH_{4(g)} + 2\,O_{2(g)} \rightleftharpoons CO_{2(g)} + 2\,H_2O_{(g)}$$

The expression of the equilibrium constant would be:

$$K_c = \frac{[CO_2] \cdot [H_2O]^2}{[CH_4] \cdot [O_2]^2}$$

`Equilibrium involving solids, liquids and solutions` Let us analyze an example of a reaction involving solids or liquids:

$$CaCO_3 \cdot H_2O_{(s)} \rightleftharpoons CaO_{(s)} + CO_{2(g)} + H_2O_{(l)}$$

Solids and liquids have no concentration and hence they should not be included in the expression of K_c. For the example above:

$$K_c = [CO_2]$$

Differently for a reaction involving aqueous solutions, for example:
$HCl_{(Aq)} \rightleftharpoons H^+_{(Aq)} + Cl^-_{(Aq)}$ $\qquad K_c = \frac{[H^+]\cdot[Cl^-]}{[HCl]}$
Overall, remember that in the expression of K_c, you can only include gases (g) or aqueous solutions (aq) as you can ignore solids and liquids without a well-defined concentration.

Sample Problem 94

Write down the expression of K_c for the following reactions:
(a) $2\,NO_{2(g)} \rightleftharpoons N_2O_{4(g)}$
(b) $C_3H_{8(g)} + 5\,O_{2(g)} \rightleftharpoons 4\,H_2O_{(g)} + 3\,CO_{2(g)}$
(c) $Zn_{(s)} + 2\,HCl_{(Aq)} \rightleftharpoons ZnCl_{2(Aq)} + H_{2(g)}$

SOLUTION

Remember you can only include gas and aqueous solution in the expression of K_c. For the first example

$$K_c = \frac{[N_2O_4]}{[NO_2]^2}$$

For the second example,

$$K_c = \frac{[H_2O]^4 \cdot [CO_2]^3}{[C_3H_8] \cdot [O_2]^5}$$

For the last example:

$$K_c = \frac{[ZnCl_2] \cdot [H_2]}{[HCl]^2}$$

STUDY CHECK

Write down the expression of K_c for the following reaction:
$Pb^{2+}_{(Aq)} + 2\,I^-_{(Aq)} \rightleftharpoons PbI_2(s)$.

9.3 Using equilibrium constants

We saw that the equilibrium constant of a reaction tells you whether there are more reactants or products in an equilibrium mixture. At the same time, one can use K_c to quantitatively calculate the value of the equilibrium concentration of reactants and products. This section will explain how to do this.

`Solving from` K_c Let's analyze the reaction of hydrogen (H_2) and iodine (I_2) to produce hydrogen iodide (HI):

$$H_2(g) + I_2(g) \rightleftharpoons 2\,HI(g)$$

The equilibrium constant at 300K is $3 \cdot 10^{-1}$. Analyzing a reaction mixture we find that the concentration of H_2 is 1M and the concentration of I_2 is 2M. We want to calculate how much HI we have in the mixture. As the concentrations of reactants and products are linked together through K_c we can certainly solve for $[\mathrm{HI}]$. The expression for K_c is:

$$K_c = \frac{[\mathrm{HI}]^2}{[\mathrm{H_2}] \cdot [\mathrm{I_2}]}$$

and we know that $[\mathrm{H_2}] = 1M$ and $[\mathrm{I_2}] = 2M$. Plugging the values in the expression of K_c, and given the numerical value of K_c we have:

$$3 \cdot 10^{-1} = \frac{[\mathrm{HI}]^2}{1 \cdot 2}$$

Solving for $[\mathrm{HI}]$ we have

$$[\mathrm{HI}]^2 = 0.6$$

To solve for $[\mathrm{HI}]$ we have to use an square root:

$$[\mathrm{HI}] = \sqrt[2]{0.6} = 0.77M$$

Sample Problem 95

The value of the equilibrium constant for the reaction

$$3\,\mathrm{H_2(g)} + \mathrm{N_2(g)} \rightleftharpoons 2\,\mathrm{NH_3(g)}$$

is $3 \cdot 10^8$ at 300K. An analysis of an equilibrium mixture gave a concentration of nitrogen and ammonia of 2M, respectively. Calculate the equilibrium concentration of hydrogen at 300K.

SOLUTION

The value of the equilibrium constant for the formation of ammonia is:

$$K_c = \frac{[\mathrm{NH_3}]^2}{[\mathrm{H_2}]^3 \cdot [\mathrm{N_2}]}$$

We know $[\mathrm{NH_3}]$ and $[\mathrm{N_2}]$ and both values are 2M, and we also know $K_c = 3 \cdot 10^8$. Plugging these values into the previous equations we obtain:

$$3 \cdot 10^8 = \frac{2^2}{[\mathrm{H_2}]^3 \cdot 2}$$

we can solve for $[\mathrm{H_2}]$:

$$[\mathrm{H_2}]^3 = 6.6 \cdot 10^{-9}$$

In order to obtain $[\mathrm{H_2}]$ we need a a cubic root:

$$[\mathrm{H_2}] = \sqrt[3]{6.6 \cdot 10^{-9}} = 1.9 \cdot 10^{-3}M$$

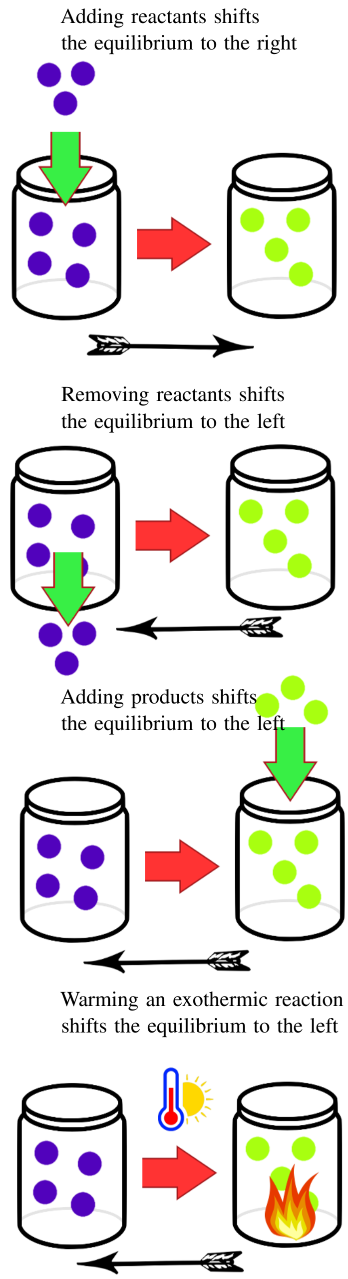
STUDY CHECK

The value of the equilibrium constant for the reaction

$$\mathrm{C_3H_{8(g)}} + 5\,\mathrm{O_{2(g)}} \rightleftharpoons 4\,\mathrm{H_2O_{(g)}} + 3\,\mathrm{CO_{2(g)}}$$

is 500 at a given temperature. An analysis of an equilibrium mixture gave a concentration of water, carbon dioxide and C_3H_8 of 1M. Calculate the equilibrium concentration of oxygen at that temperature.

9.4 Le Châtelier principle



▼Le Châtelier principle helps predicts the outcome of altering an equilibrium mixture.

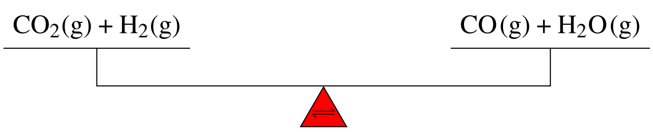
At this point, we have covered the idea of equilibrium and we have seen that the forward and reverse reactions have the same rate at equilibrium. Now, what happens if you alter this equilibrium? Le Châtelier principle claims a reaction will go back to its original equilibrium state by shifting left or right.

`Le Châtelier principle` When a reaction is in equilibrium the forward and reverse reactions proceed at the same speed. Also in an equilibrium state, the concentrations of reactants and products have very specific values. Imagine that you create stress conditions by adding reactants or products or even changing the temperature. This stress will have an impact on the equilibrium and the reaction eventually will reach a new state of equilibrium by somehow counteracting this stress. Le Châtelier principle says that when stress is placed in a reaction (adding or removing reactant or products, increasing or decreasing temperature) the equilibrium will be shifted in the direction that relieves that stress. Table 9.2 displays different aspects regarding Le Châtelier's principle in terms of parameter change and consequence.

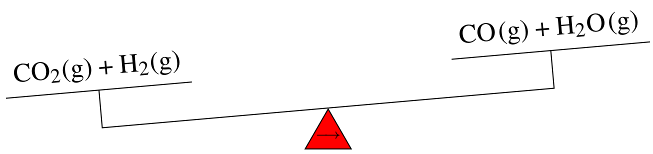
`Change in concentration` Let us consider the following equilibrium:

$$CO_2(g) + H_2(g) \rightleftharpoons CO(g) + H_2O(g)$$

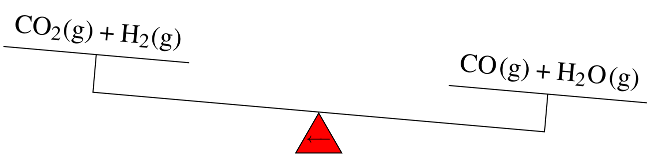
K_c for this reaction equals one at 1200K. This means that the concentration of reactants and products are the same. We can represent this using this balance or seesaw



If we add some CO_2 the equilibrium will be affected. To counteract this stress, the reaction will restore the equilibrium by decreasing the amount of CO_2. This can only be achieved by displacing the equilibrium to the right so that CO_2 is removed. Mind that CO_2 is consumed if the reaction moves from reactants $\longrightarrow$ to products and it is produced when going from products $\longleftarrow$ to reactants. We can represent this with the following seesaw.



As we added CO_2 the reactants now weigh more and hence the reaction has to proceed to the right $\longrightarrow$. Now imagine we remove some $CO_2(g)$. Again, the equilibrium will be affected and the reaction will restore its equilibrium state by doing the opposite, that is producing $CO_2(g)$ as the reaction proceeds from reactants $\longleftarrow$ to products. Again using the seesaw:

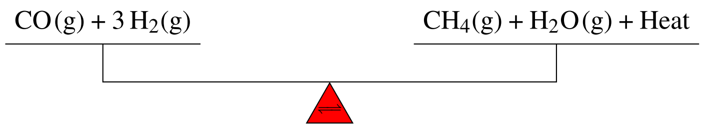


We can also add a different chemical that is not involved in the equilibrium. In this case, the equilibrium will not be affected by this change.

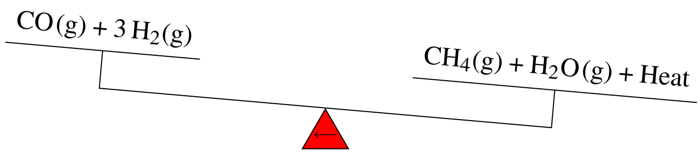
`Temperature change` Let us consider the following equilibrium that produces heat–remember we describe these types of reactions as exothermic:

$$\mathrm{CO\,(g) + 3\,H_2\,(g) \rightleftharpoons CH_4\,(g) + H_2O\,(g) + Heat}$$

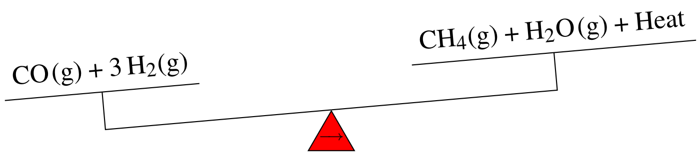
Again, this reaction is in equilibrium so we can use the same seesaw analogy.



If we increase the temperature of the system, the equilibrium will be affected. To go back to an equilibrium state the reaction will decrease the temperature of the container. As the reaction produces heat, a way to decrease the system temperature is to generate reactants ($\longleftarrow$). Again, using the scale that means:



Differently now if we decrease the temperature, the reaction will increase the temperature by going back to its equilibrium state going from reactants $\longrightarrow$ to products. This is because heat is produced as a product



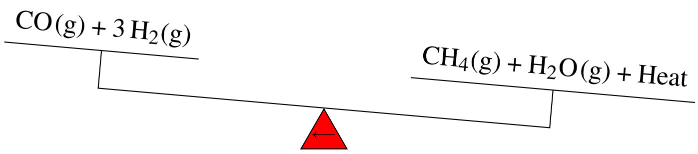
`Volume change` We can also think about increasing or decreasing the volume in which the reaction takes place. This change will have an impact on the reaction equilibrium as the concentrations of reactants and products will be altered by this change. Changes in volume will shift the reaction towards the left or right depending on the overall stoichiometric change of the reaction, that is on whether the reaction produces or consumes molecules. For reactions that generate matter, that is, in the case that $\Delta n > 0$, increasing the volume will follow the increase of the number of moles. In other words, by increasing the volume, the equilibrium will shift towards the products, that is towards the right. For reactions that consume matter ($\Delta n < 0$), increasing the volume will shift the equilibrium towards the reactants, that is towards the left. For example, the reaction below consumes molecules:

$$\mathrm{CO\,(g) + 3\,H_2(g) \rightleftharpoons CH_4(g) + H_2O\,(g)}$$

Table 9.2 Le Châtelier principle

Parameter	Change	Effect
Concentration	Add reactants	$\longrightarrow$
	Remove reactants	$\longleftarrow$
	Add products	$\longleftarrow$
	Remove products	$\longrightarrow$
Volume	Increase volume	More moles
	Decrease volume	Less moles
Catalyst	Add a catalyst	No effect
Temperature		
Endothermic reaction	Increase temperature	$\longrightarrow$
	Decrease temperature	$\longleftarrow$
Exothermic reaction	Increase temperature	$\longleftarrow$
	Decrease temperature	$\longrightarrow$

if we increase the volume of the container in which the reaction takes place, the equilibrium will shift toward the left:



The opposite shift will follow a volume decrease as the reaction shift towards the right.

Sample Problem 96

For the next endothermic reaction indicate whether the reaction will shift right ($\longrightarrow$) or left ($\longleftarrow$) after the following changes:

$$\mathrm{N_2(g) + O_2(g) + Heat \rightleftharpoons 2\,NO\,(g)}$$

(a) adding reactants (b) adding products (c) decreasing the temperature.

SOLUTION

(a) Adding reactants always displaces the equilirbium so that reactants are consumed, hence the reaction will proceed $\longrightarrow$. (b) After adding products the reaction will tend to reduce the amount of products, and hence it will go $\longleftarrow$. (c) The reaction is endothermic that means that it consumes heat. If we decrease the temperature it will tend to increase the temperature and hence heat needs to be formed. This will only happen if the reaction proceeds ($\longleftarrow$).

STUDY CHECK

For next exothermic reaction indicate whether the reaction will shift right ($\longrightarrow$) or left ($\longleftarrow$) after the following changes:

$$\mathrm{C\,(g) + O_2(g) \rightleftharpoons CO_2(g) + Heat}$$

(a) removing reactants (b) removing products (c) decreasing the temperature.

9.5 Collision theory

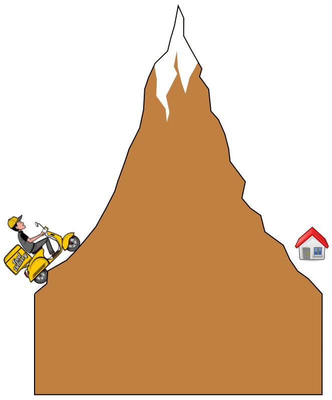
The rate of reaction depends on temperature and reactions proceed at a higher speed at higher temperatures. This is the reason we use refrigerators and freezers to keep food for a longer time. The collision theory provides a microscopic understanding of chemical reactions. A pivotal element of this theory is the transition state also called the activated complex. This is an ephemeral state between reactants and products so that if the reaction proceeds until this point, reactants will certainly evolve into products. The energy difference between the reactants and the transition state of a reaction is called the activation energy. Energy diagrams are useful representations of the energy changes involved in a reaction. By analyzing an energy diagram we can predict whether a reaction is endo of exothermic and estimate the energy involved in the activation of a reaction. This section covers the fundamental principles of collision theory.

`Collision theory` Chemical reactions proceed at a molecular level by means of the collision of reactant molecules. The frequency of collisions is directly related to the reaction rate so that the more collisions per second the faster the reaction would proceed as the higher would be the chance for a reaction to occur. The collision theory of the chemical reactions associates the rate to the frequency of collisions. By increasing the concentration of reactants the number of collisions would increase and that is the reason why there is a concentration dependency on the reaction rate. However, not all collisions are equally effective and only effective collisions will contribute to chemical reactions. Two main factors control the effectiveness of the molecular collisions. First, the orientation of the collisions is critical and only certain orientations will produce successful collisions. The second factor is the energy of the molecules colliding. Only molecules with energy beyond a threshold will be able to form products. In other words: a successful collision should have the right orientation while having enough energy. Let us elaborate on the role of the orientation for the reaction:

$$\mathrm{Cl + NOCl \longrightarrow Cl_2 + NO}$$

We have that a Cl atom reacts with a NOCl molecule to produce Cl_2, as represented in the image below. Only with the Cl atoms hit the NOCl molecule through the Cl atom, the collision will proceed with the right orientation in order to produce the Cl_2 molecule. At the same time, if the orientation is correct and the molecules have enough energy a transition state will form. We will discuss more this state during the next sections.

`Transition state and activation energy` We have discussed that the energy of reactants is a key parameter during a reaction. The minimum energy needed to activate the reactants so that they can produce products is called activation energy. The higher the activation energy the more energy is needed for a reaction to happen and hence the lower the reaction rate will be. When reactants get together they increase their energy from reactants until a state of maximum energy before producing the products. This state of maximum energy is called an activated complex or transition state. Transition states are very ephemeral states that exist for a very short amount of time in comparison to reactants and products. An example of activation happens during the ignition of a gas burner. Cooking gas and oxygen can react together to produce water and carbon dioxide only with the help of a spark. A spark is needed to activate the reactants so that their molecules can reach the transition state and hence produce products. Transition states are normally indicated with a ‡ sign to indicate they have a very different nature than normal molecules. The figure below represents two reaction pathways with different activation energy.



`Effect of temperature` As temperature increases the kinetic energy of the molecules reacting increases as well and with this the energy involved in the collisions also increases. Molecules move faster at higher temperature and hence the frequency of collisions between reactants and molecules increases with temperature. Overall, the reaction rate increases with temperature. In other words, we can speed up chemical reactions by increasing temperature.

`Concentration of reactants` Another way to favor the collision between the molecules of reactants is to increase its concentration. The larger the concentration of reactants the higher the probability of their molecules to collide with each other potentially producing products.

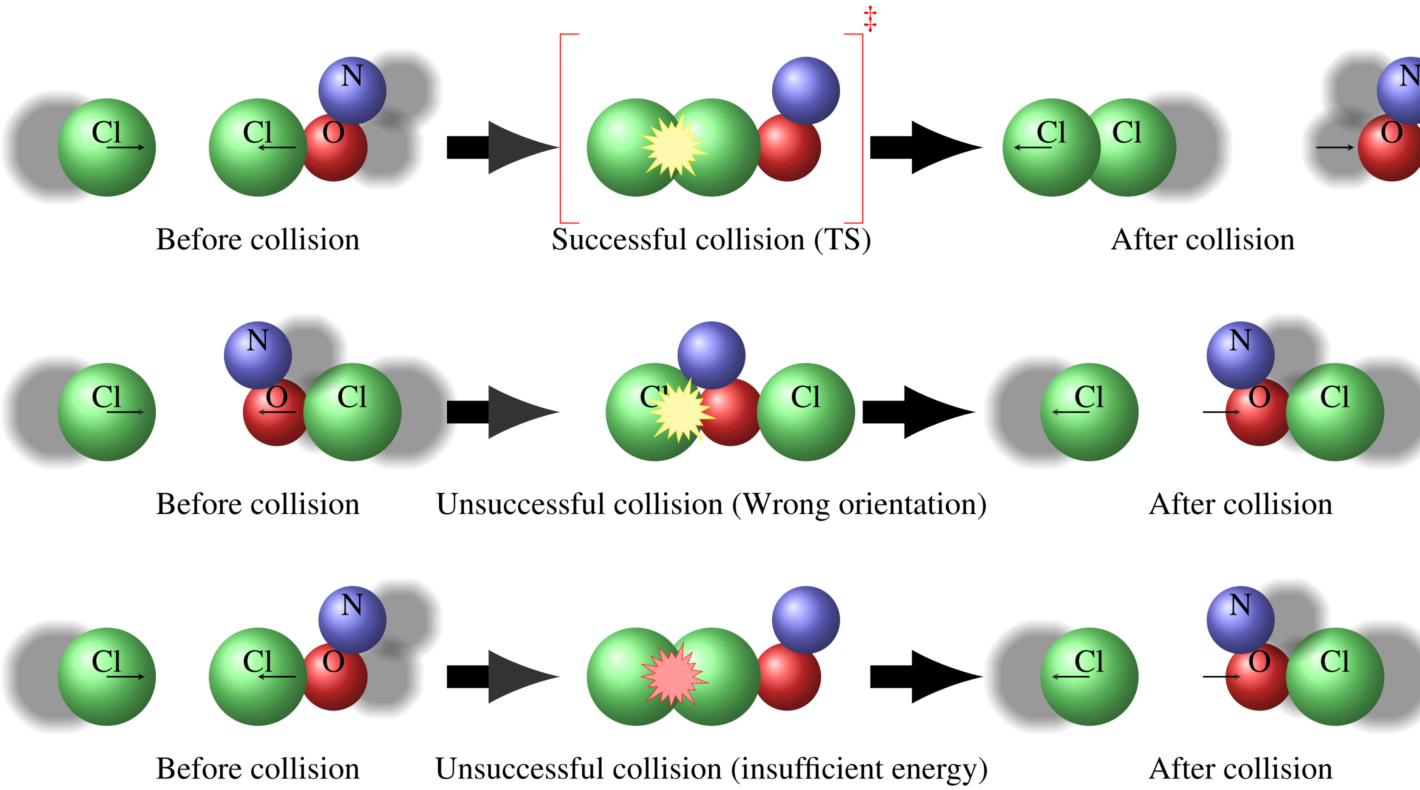


Figure 9.2 Example of three collisions, a successful collision that leads to a transition state and the products and two unsuccessful collisions due to an ineffective molecular arrangement and insufficient energy.

Sample Problem 97

Methane (CH_4) reacts with oxygen (O_2) to produce carbon dioxide (CO_2) and water (H_2O) according to the following reaction

$$CH_4(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

(a) What happens to the number of collision between CH_4 and O_2 when you add extra oxygen?; (b) How would increasing temperature impacts the rate of the reaction?

SOLUTION

(a) The more reactants the mode collisions between their molecules. (b) Increasing temperature would increase the rate of the combustion reaction, as increasing temperature increases collisions.

STUDY CHECK

How would the following changes affect the rate of this reaction:

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(g)$$

(a) Removing oxygen; (b) Decreasing temperature

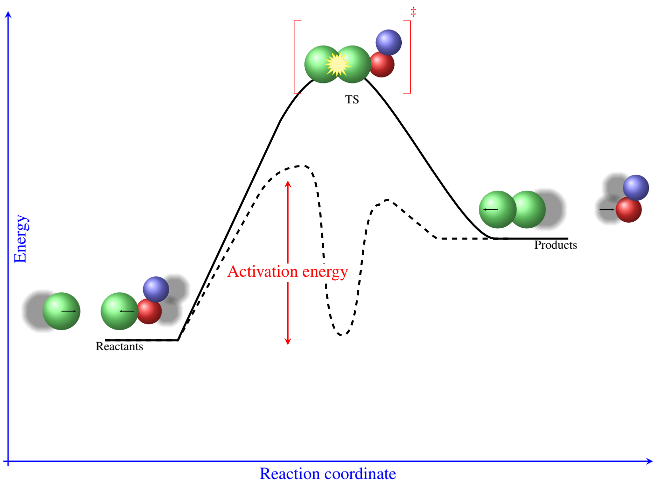


Figure 9.3 Energy diagram showing the catalyzed and non-catalyzed pathways

`Catalysts` Catalysts are chemicals that are not directly involved in a reaction–they are neither reactants nor products and as such, they are not consumed during a reaction. Still, they have the ability to increase the rate of reactions. Each catalyst is specific to a given reaction. As such, a catalyst can speed up a reaction without affecting other reactions. Enzymes are biological catalysts that support chemical reactions happening in the body. Most of the biological reactions involved in the chemistry of life are very slow. It is only with the help of biological catalysts that these reactions can indeed happen at reasonable rates. Every single metabolic reaction has associated an enzyme responsible for speeding up the process. Similarly, catalysts are well-known compounds in the chemical industry. The production of chemicals in the industry is also supported by the help of catalysts which make the chemical processes more economically beneficial. For example, catalysts are used in the manufacturing of margarine from vegetable oils. In the same line, the synthesis of ammonia is catalyzed by iron, the manufacturing of sulfuric acid uses catalysts based on nitrogen(II) oxide and platinum, and the oxidation of hydrocarbons in automobile exhausts are catalyzed by platinum. Enzymes are also added to detergents in order to speed up the breaking of protein molecules in stains. How do catalysts work? Catalyst function by providing alternative reaction pathways with more favorable–normally lower–activation energy. This way, in the presence of a catalyst, reactions proceed by means of two different reaction pathways, the catalyzed and non-catalyzed pathways. Still, as the catalyzed pathway is more favorable the overall speed of the reaction is higher. Catalysts can be classified as homogeneous, heterogeneous, and acid catalysts.

`Heterogeneous catalysis` A heterogeneous catalyst is a material with catalytic properties which has a different state of matter than the main reaction. For example, the synthesis of ammonia is carried in the chemical industry by means of reacting has hydrogen and nitrogen. The catalyst for this process is iron, which is a solid. The reaction proceeds in a gas phase, whereas the catalyst employed is solid. This is an example of heterogeneous catalysis. Another example of the hydrogenation of alkenes by means of molecular hydrogen. This process is used for example in the production of saturated fats or shortenings such as Crisco. As hydrogen hardly dissociates in gas, this reaction proceeds at a very low rate without the help of a catalyst. Differently, in the presence of platinum, the dissociation energy of hydrogen is lower what makes the process function at a faster rate. The necessary steps for a reaction to happen on a heterogeneous catalyst are: (a) the adsorption of the reactants on the surface (the words adsorption refers to the attachment of a gas onto a solid) (b) the activation of the reactants in which its molecular geometry changes due to the adsorption (c) the migration of the adsorbed reactants (d) the surface reaction (e) the desorption of the products

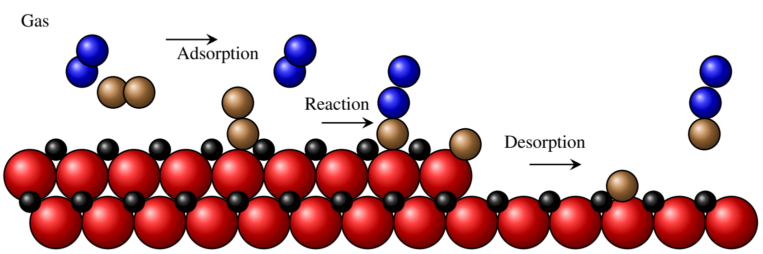


Figure 9.4 Steps in a heterogeneously catalyzed reaction

CHAPTER 9

Chemical Equilibrium

9.1 True or false:(a) At equilibrium, the rate of the reverser reaction is twice the rate of the forward reaction (b) At equilibrium, the concentration of products do not change (c) At equilibrium, the concentration of reactants do not change (d) At equilibrium, the concentration of reactants and products do not change

9.2 True or false:(a) At equilibrium, the rate of the reverse reaction do not change (b) At equilibrium, the rate of the forward reaction do not change (c) At equilibrium, the rate of the reverse reaction equals the rate of the forward reaction (d) At equilibrium, the concentration of reactants and products are not constant

Equilibrium constants

9.3 Write down the forward and reverse reactions for the following reactions in equilibrium:

(a) $CH_{4(g)} + O_{2(g)} \rightleftharpoons CO_{2(g)} + H_2O_{(g)}$

(b) $2\,Mg_{(s)} + O_{2(g)} \rightleftharpoons 2\,MgO_{(s)}$

9.4 Write down the forward and reverse reactions for the following reactions in equilibrium:

(a) $BaCO_{3(s)} \rightleftharpoons Ba^{2+}_{(aq)} + CO_3{}^{2-}{}_{(aq)}$

(b) $2\,H_{2(g)} + O_{2(g)} \rightleftharpoons 2\,H_2O_{(l)}$

9.5 For the reactions below and given the value of the equilibrium constant indicate whether the equilibrium mixture will have: (a) More reactants than products (b) More products than reactants (c) Same amount of products and reactants

(a) $CO_{2(g)} + H_2O_{(g)} \rightleftharpoons CH_{4(g)} + O_{2(g)}$ $\quad K_c = 0.001$

(b) $N_{2(g)} + O_{2(g)} \rightleftharpoons 2\,NO_{(g)}$ $\quad K_c = 2 \cdot 10^{25}$

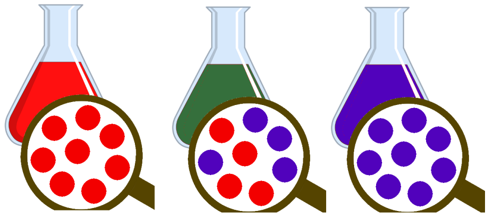
(c) $2\,NO_{(g)} + O_{2(g)} \rightleftharpoons 2\,NO_{2(g)}$ $\quad K_c = 6.4 \cdot 10^{9}$

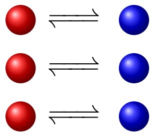
9.6 For the reactions below and given the value of the equilibrium constant indicate whether the equilibrium mixture will have: (a) More reactants than products (b) More products than reactants (c) Same amount of products and reactants

(a) $N_{2(g)} + 3\,H_{2(g)} \rightleftharpoons 2\,NH_{3(g)}$ $\quad K_c = 1$

(b) $2\,NO_{(g)} + Cl_{2(g)} \rightleftharpoons 2\,NOCl_{(g)}$ $\quad K_c = 6.5 \cdot 10^{4}$

9.7 Indicate which of the following diagrams represent better the system at equilibrium:



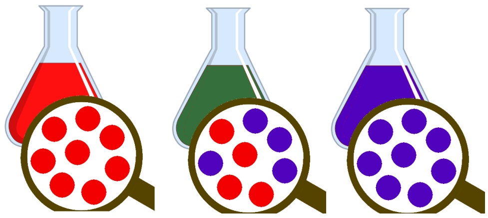


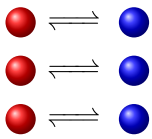
(a) $K_c = 10$

(b) $K_c = 0.1$

(c) $K_c = 1$

9.8 Indicate which of the following diagrams represent better the system at equilibrium:





(a) $K_c = 1$

(b) $K_c = 0.04$

(c) $K_c = 200$

9.9 Write down the expression of K_c for the following reaction:

(a) $2\,SO_{2(g)} + O_{2(g)} \rightleftharpoons 2\,SO_{3(g)}$

(b) $CO_{(g)} + 2\,H_{2(g)} \rightleftharpoons CH_3OH_{(g)}$

(c) $C_2H_{6(g)} + Cl_{2(g)} \rightleftharpoons C_2H_2Cl_{(s)} + HCl_{(g)}$

9.10 Write down the expression of K_c for the following reaction:

(a) $BaCO_{3(s)} \rightleftharpoons Ba^{2+}_{(aq)} + CO_3{}^{2-}{}_{(aq)}$

(b) $2\,H_{2(g)} + O_{2(g)} \rightleftharpoons 2\,H_2O_{(l)}$

Using equilibrium constants

9.11 The reaction of carbon monoxode with hydrogen to produce methanol has a equilibrium constant in terms of concentration that at a certain temperature is larger than one

$$CO(g) + 2\,H_2(g) \rightleftharpoons CH_3OH(g) \qquad K_c = 14.5$$

Calculate: (a) the equilibrium concentration of hydrogen (H_2) given that the equilibrium concentration of methanol (CH_3OH) and carbon monoxice (CO) for the reaction is 2M, respectively. (b) the equilibrium concentration of hydrogen (H_2) given that the equilibrium concentration of methanol (CH_3OH) and carbon monoxide (CO) for the reaction are 3M and 1M, respectively.

9.12 Consider the following reaction:

$$2\,SO_2(g) + O_2(g) \rightleftharpoons 2\,SO_3(g)$$

(a) Write down the expression of K_c. (b) Calculate the numerical value of K_c for the reaction if the concentrations at equilibrium at 1000K are 2M for SO_3, 0.3M for O_2 and 1M for SO_2. (c) indicate whether an equilibrium mixture will contain mostly products, mostly reactants or equal amounts of reactants and products.

Le Châtelier principle

9.13 Using the Le Châtelier principle indicate whether the reaction below

$$2\,SO_2(g) + O_2(g) \rightleftharpoons 2\,SO_3(g)$$

will shift in the direction of products ($\longrightarrow$) or reactants ($\longleftarrow$) after the following actions: (a) add SO_2 (b) add SO_3(c) remove O_2

9.14 Using the Le Châtelier principle indicate whether the reaction below

$$2\,NH_{3(g)} \rightleftharpoons N_2 + 3\,H_2$$

will shift in the direction of products ($\longrightarrow$) or reactants ($\longleftarrow$) after the following actions: (a) add NH_3 (b) add N_2(c) remove H_2

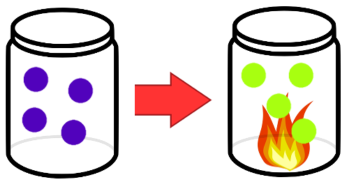
9.15 According to Le Châtelier principle indicate whether the reaction will shift in the direction of products ($\longrightarrow$) or reactants ($\longleftarrow$) after we increase temperature:

$$CO(g) + 2\,H_2(g) \rightleftharpoons CH_3OH(g) + \text{Heat}$$

9.16 According to Le Châtelier principle indicate whether the reaction will shift in the direction of products ($\longrightarrow$) or reactants ($\longleftarrow$) after we increase temperature:

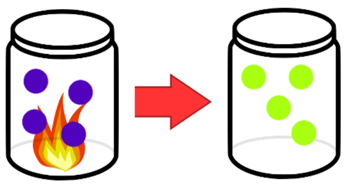
$$2\,B(s) + 3\,H_2(g) + \text{Heat} \rightleftharpoons B_2H_6(g)$$

9.17 According to Le Châtelier principle indicate whether the following reaction will shift in the direction of products ($\longrightarrow$) or reactants ($\longleftarrow$) after the following changes:



(a) adding reactants (b) increasing temperature (c) decreasing temperature

9.18 According to Le Châtelier principle indicate whether the following reaction will shift in the direction of products ($\longrightarrow$) or reactants ($\longleftarrow$) after the following changes:



(a) adding products (b) removing products(c) increasing temperature

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Ch. 10. Acids & Bases

Acids and bases are very important chemicals in our everyday life. Think about vinegar or Sour Patch Kids. On one hand, vinegar tastes sour as it contains acetic acid. Sour Patch Kids, on the other hand, are coated in a combination of sugar and acids. Acids help us digest food and help bacteria produce yogurt or cottage cheese. Bases on the other hand are used in drain openers, oven cleaners, or the production of soap. This chapter covers the properties of acids and bases qualitatively and quantitatively. You will learn how to identify each of these chemicals and categorize them according to their strength. Yes! acids and bases are strong, and some of them can seriously hurt you. More importantly, this chapter introduces the idea of PH, which quantifies the acidity of a solution. The PH of an acid or base depends on its strength and here we will cover how to compute the PH of solutions of strong and weak acids and bases. Balancing PH is crucial for health. Finally, we will briefly cover the idea of a buffer that helps regulate the PH of solutions and titrations used to elucidate the molarity of an unknown acid or base.

10.1 The nature of acids and Bases

Acids have very different properties than bases. Acids are acidic, have a sour taste, and can sting to the touch. Bases are basic, have a bitter–chalky–taste, and feel soapy–slippery–to the touch. On one hand, acids are extensively used in the food and perfume industry. For example, vinegar–a liquid solution of acetic acid–is used in pickles and food preparations. On the other hand, lemon and orange juice, containing citric acid, is used in the preparation of effervescent salts and as food preservatives. Acids are also used in the production of batteries. For example, car batteries contain corrosive sulphuric acid. On the other hand, bases are extensively used in manufacturing. As a first example, sodium hydroxide is used in the manufacture of soap, medicines, and even paper. As a second example, calcium hydroxide–also known as slaked lime–is used to neutralize the acid in water supplies or as an antidote for food poisoning. This hydroxide is also used in the construction industry, mixed with sand and water to make mortar. As a third example, potassium hydroxide (KOH), is used in alkali batteries. Finally, ammonia is an extensively used cleaning product, also used to remove ink spots from clothes or grease from window panes.

Strong and weak acids and bases Strong acids are strong electrolytes that completely dissociate to produce protons. Similarly, strong bases completely dissociate in solution generating hydroxyls. On the contrary, weak acids and bases are just weak electrolytes that dissociate partially in solution generating only a small amount of protons and hydroxyls. The dissociation of strong acids and bases is represented using a single arrow, whereas the dissociation of weak acids and bases is represented using a double harpoon. For example, hydrochloric acid is a strong electrolyte and its dissociation is represented by

▼Hydrofluoric acid is a weak acid used to dissolve glass.

$$HCl_{(g)} \longrightarrow Cl^-_{(aq)} + H^+_{(aq)} \qquad \text{(strong acid)}$$

whereas hydrofluoric acid is a weak electrolyte and its dissociation is represented by

$$HF_{(g)} \rightleftharpoons F^-_{(aq)} + H^+_{(aq)} \qquad \text{(weak acid)}$$

`Arrhenius acid-base model` In general terms, we can identify some acids and bases by inspecting its formula. Svante Arrhenius claimed around 1884 that acids are acidic because contain hydrogen in their structure and when dissolved in water they produce *protons*: H^+, also called hydronium ion written as H_3O^+. Let us consider these chemicals: HF, H_2SO_4 and HNO_3. All these chemicals, named hydrofluoric acid, sulfuric acid, and nitric acid, are Arrhenius acidic. The reaction below described the process of dissociation of hydrogen chloride to produce chloride and a proton:

▼Citrus such as lemons or oranges are acidic.

$$HCl_{(g)} \longrightarrow Cl^-_{(aq)} + H^+_{(aq)} \qquad \text{(an acid)}$$

Based on the acid dissociation reaction above, we say chloride is the result of deprotonating hydrochloric acid. In other words, chloride is deprotonated. Based on this dissociation reaction that happens in water, we can say hydrogen chloride also known as hydrochloric acid is an Arrhenius acid, as it produces protons in water. We have that acid are classified as monoprotic, diprotic, and polyprotic. Monoprotic acids have only one acidic H on their molecule (e.g. HNO_3), whereas diprotic acids have two (e.g. H_2NO_4) and polyprotic acids have more than two (e.g. H_3PO_4). Differently, bases are basic because when dissolved in water they produce *hydroxyls*: OH^-. Hydroxides are Arrhenius bases and for example, NaOH and $Ca(OH)_2$, named sodium hydroxide and calcium hydroxide, are well-known bases. Let us address now the dissociation of sodium hydroxide in water:

▼Pickles contain vinegar that is a solution of acetic acid in water.

$$NaOH_{(s)} \longrightarrow Na^+_{(aq)} + OH^-_{(aq)} \qquad \text{(a base)}$$

This chemical is an Arrhenius base as it produces hydroxyls. Therefore, Arrhenius acids and bases produce protons and hydroxyls in water, respectively. However, the Arrhenius model does not explain why chemicals unsolved in water can also be acidic or even why chemicals such as NH_3–a molecule without OH on its structure–can be basic.

`Brönsted-Lowry acid-base model` In 1923 two different chemists Thomas Lowry and Johannes Brönsted proposed what is now known as the Brönsted-Lowry model of acids and bases. The Brönsted-Lowry model is a more advanced acid-base model. This model claims acids are chemicals that give away protons (H^+) whereas bases receive protons. This way, Brönsted-Lowry extends the Arrhenius model to other solvents different than water, as the solvent is not part of the definition. Based on this model, we can understand how ${HCO_3}^-{}_{(aq)}$ can act as an acid giving away protons to water and as a base receiving protons from water. When acting as an acid:

$$ {HCO_3}^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons {CO_3}^{2-}{}_{(aq)} + H_3O^+_{(aq)} $$

When acting as a base:

$$HCO_3{}^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons H_2CO_{3(aq)} + OH^-_{(aq)}$$

As you can see, an acid-base reaction is essentially a proton transfer reaction, in which a proton H^+ transfers from an acid into a base. Compounds that can act as acids or bases are referred to as *amphiprotic* as they can act as a proton donor or acceptor. Other examples are: H_2O, HS^- or $HSO_4{}^-$. Still, the Brönsted-Lowry model does not explain why chemicals such as ammonia are a base and carbon dioxide an acid. In particular, the model does not justify what structural particularity makes ammonia behave as a base and carbon dioxide (with no hydrogen on its structure) as an acid.

Sample Problem 98

Identify the following chemicals as Arrhenius acids or bases and give their names: HCl, KOH, H_3PO_4 and CH_3COOH.

SOLUTION

The acids are: HCl, H_3PO_4 and CH_3COOH. Their names are: hydrochloric acid, phosphoric acid and acetic acid, the later is a common name. KOH is a base called potassium hydroxide.

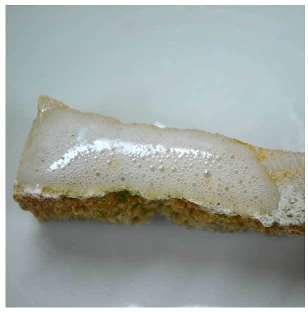
STUDY CHECK

Identify the following chemicals as Arrhenius acids or bases and give their names: NaOH and H_2CO_3.

▼Bath bombs are made of acidic and basic ingredients that combine in water to make a fizzy bath time experience.

▼Ashes are basic.

▼Limestone reactant with hydrochloric acid to give carbon dioxide bubbles.



10.2 Dissociation of acids and bases

This second section will cover the acid and base dissolution in water. Water plays a key role in the acid-base character of a chemical as these chemicals ultimately react with water. When acids and bases solve in water, they dissociate producing a byproduct called the conjugate base and conjugate acid. We will describe how to set up the dissociation equilibrium and how to identify conjugate acid-base pairs.

`Conjugate acids and bases` A conjugate acid-base pair are molecules or ions related by the loss of one H^+. For example: hydroiodic acid HI and iodate I^- or water H_2O and protons H_3O^-. The product of the dissociation of acids is a conjugate base. For example:

$$\underset{\text{Acid}}{HI_{(l)}} + H_2O_{(l)} \longrightarrow \underset{\text{Conjugate Base}}{I^-_{(aq)}} + H_3O^+_{(aq)}$$

Similarly, bases produce a conjugate acid. In the example below, water acts as a base and a proton is the conjugate acid:

$$HI_{(l)} + \underset{\text{Base}}{H_2O_{(l)}} \longrightarrow I^-_{(aq)} + \underset{\text{Conjugate Acid}}{H_3O^+_{(aq)}}$$

At the same time acids react with bases as they have opposite characters. Following the previous example:

$$\underset{\text{Acid}}{HI_{(l)}} + \underset{\text{Base}}{H_2O_{(l)}} \longrightarrow \underset{\text{Conjugate Base}}{I^-_{(aq)}} + \underset{\text{Conjugate Acid}}{H_3O^+_{(aq)}}$$

Hence, we have that acid reactants with a base to produce a conjugate base and a conjugate acid. We can use the diagram below to identify the acid-conjugate base pairs:

H^+ loss

Acid + Base ⟶ Conjugate Base + Conjugate Acid

H^+ gain

Sample Problem 99

Identify the acid, the base, the conjugate acid and the conjugate base in the reaction:

$$NH_{3(g)} + H_2O_{(l)} \rightleftharpoons NH_4{}^+{}_{(aq)} + OH^-_{(aq)}$$

SOLUTION

The trick is first identifying the acid and the base starting for the left side of the formula. In this case $NH_{3(g)}$ is the base and hence water is the acid. Now connect the acid and the base with the other side of the arrow, and use conjugate with the opposite term. For example: NH_3 is a base and should be related with $NH_4{}^+$ that is the conjugate acid. Similarly, H_2O is a acid, being related with OH^- that is the conjugate base. In summary:

$$\underset{\text{Base}}{NH_{3(g)}} + \underset{\text{Acid}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{Conjugate Acid}}{NH_4{}^+{}_{(aq)}} + \underset{\text{Conjugate Base}}{OH^-_{(aq)}}$$

STUDY CHECK

Identify the acid, the base, the conjugate acid and the conjugate base in the reaction:

$$CO_3{}^{2-}{}_{(aq)} + H_2O_{(l)} \rightleftharpoons HCO_3{}^-{}_{(aq)} + OH^-_{(aq)}$$

`Writing down acid-base equilibria` Now let us address how to write down acid-base equilibria from scratch, starting with the dissociation of ammonia (NH_3). Dissociation reactions are the reaction of an acid or base and water:

$$H_2O_{(l)} + NH_{3(g)} \rightleftharpoons$$

We have that ammonia is a base and hence water will act as an acid. Bases receive protons whereas acids give protons away. In the equilibrium, we will remove one proton from water and add it to ammonia, producing respectively a hydroxyl and an ammonium ion:

H^+ loss

$$\underset{\text{Acid}}{H_2O_{(l)}} + \underset{\text{Base}}{NH_{3(g)}} \rightleftharpoons \underset{\text{Conjugate Base}}{OH^-_{(l)}} + \underset{\text{Conjugate Acid}}{NH_4{}^+{}_{(aq)}}$$

H^+ gain

And that will give the dissociation equilibrium of ammonia.

$$\underset{\text{Acid}}{H_2O_{(l)}} + \underset{\text{Base}}{NH_{3(g)}} \rightleftharpoons \underset{\text{Conjugate Base}}{OH^-_{(aq)}} + \underset{\text{Conjugate Base}}{NH_4{}^+{}_{(l)}}$$

We can now address the reaction between an acid and a base for example, between sodium hydroxide and hydroiodic acid:

$$\underset{\text{Acid}}{HI_{(g)}} + \underset{\text{Base}}{NaOH_{(aq)}} \longrightarrow$$

In these types of acid-base reactions, we have that the acid will generate a proton and the base will generate a hydroxyl, which will both combine to produce water and a salt, sodium chloride:

H^+ loss

$$\underset{\text{Acid}}{HI_{(g)}} + \underset{\text{Base}}{NaOH_{(s)}} \longrightarrow NaCl_{(s)} + H_2O_{(l)}$$

OH^- loss

This way, dissociation and acid-base reaction function in a very similar manner.

`Including water in the dissociation` Let us consider an acid such as HCl. We know acids produce protons (H^+ or H_3O^+) so we can express the acid dissociation as:

$$HCl_{(g)} \xrightarrow{H_2O} H^+_{(aq)} + Cl^-_{(aq)}$$

We indicate water on top of the arrow to represent that the dissociation process happens in water. There is an alternative way to represent this process by including explicitly water:

$$HCl_{(g)} + H_2O_{(l)} \longrightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$$

This way we explicitly represent the role of water as a proton receiver. Ultimately, both ways are correct and for some chemicals (e.g. NH_3) it is more convenient to use the role of water whereas for others it makes no difference. In the following, we present more examples. For the case of HNO_3, nitric acid, we have:

$$HNO_3(l) \xrightarrow{H_2O} H^+_{(aq)} + NO_3{}^-{}_{(aq)}$$

that means once dissolved in water, nitric acid gives a proton and forms a nitrate $NO_3{}^-$ cation. Some other acids, such as H_2SO_4 are diprotic, as they have two protons and hence they can lose both while reacting with water:

$$H_2SO_4(l) \xrightarrow{H_2O} 2\,H^+_{(aq)} + SO_4{}^{2-}{}_{(aq)}$$

Bases produce hydroxyls (OH^-) and an example of a base dissociation would be:

$$NaOH(s) \xrightarrow{H_2O} Na^+_{(aq)} + OH^-_{(aq)}$$

Sample Problem 100

Write down the dissociation reaction using double arrows for the following chemicals: $H_3PO_{4(l)}$ and $NH_{3(g)}$.

SOLUTION

Phosphoric acid is a triprotic acid with three possible protons that can be given away:

$$H_3PO_{4(l)} \overset{H_2O}{\rightleftharpoons} 3\,H^+_{(aq)} + PO_4{}^{3-}{}_{(aq)}$$

As the molecules contains protons there is no need to explicitly include water in the equilibrium. Ammonia is a base and needs is the only case in which you need to explicitly use water to help dissociate the base. This is because ammonia does not contain hydroxyls.

$$NH_{3(g)} + H_2O_{(l)} \rightleftharpoons NH_4{}^+{}_{(aq)} + OH^-_{(aq)}$$

STUDY CHECK

Write down the dissociation reaction using double arrows for the following chemicals: $HI_{(g)}$ and $HClO_{2(l)}$.

10.3 Strength of acids and bases

At this point, we are familiar with acids and bases. Acids have a sour taste and produce protons in water. Differently, bases feel soupy to the touch and produce hydroxyls. Both acids and bases react together giving conjugate species. This section gains further insight into the strength of acids and bases. Some acids are weaker while others are stronger. The same idea can be applied to bases. Here we will also learn how to quantify the strength of an acid or base and how to compare the acidic or basic character of a chemical.

`Review of acid-base strength` Acids and bases are indeed electrolytes. Remember electrolytes can be weak or strong depending on the degree to which they dissociate. Strong acids are strong electrolytes that dissociate completely in water producing large quantities of protons H^+. Strong bases are strong electrolytes that dissociate completely to produce this time large quantities of hydroxyls OH^-. Weak acids or weak bases dissociate only partially and hence they produce fewer protons or hydroxyls. Examples of strong electrolytes are: HCl, H_2SO_4 or HNO_3. As they dissociate completely we use a single arrow to indicate they are strong electrolytes:

$$HCl_{(l)} \xrightarrow{H_2O} H^+_{(aq)} + Cl^-_{(aq)}$$ One arrow indicates strong acids or bases

Weak acids or bases that only dissociate partially are represented by a double arrow as the reaction is indeed an equilibrium. Examples of weak acids or bases are NH_3 or HF.

$$HF_{(l)} \xrightleftharpoons{H_2O} H^+_{(aq)} + F^-_{(aq)} \qquad \text{A double arrow indicates weak acids or bases}$$

`Strength of acids and bases` How do we quantify the strength of an acid or base? Weak acids dissociate partially in the water. And hence, the dissociation process is in equilibrium. This means we will have at the same time in the same container the molecular form of the acid and its ionic-conjugate form. For this reason, we can employ equilibrium constant K_a–in this chapter they are called acidity constant equivalent to K_c–to characterize the degree of dissociation. The larger K_a the stronger the acid and hence the more protons will the acid produce in the solution. Let's consider the case of HF. We have that the expression of the acidity constant will be:

$$HF_{(aq)} \xrightleftharpoons{H_2O} H^+_{(aq)} + F^-_{(aq)} \qquad K_a = \frac{[H^+]\cdot[F^-]}{[HF]}$$

If we compare the acidity constant of two different acids, for example, hydrofluoric acid and hypochlorous acid

$$HF_{(aq)} \xrightleftharpoons{H_2O} H^+_{(aq)} + F^-_{(aq)} \qquad K_a = 6.30 \times 10^{-4}$$

$$HClO_{(l)} \xrightleftharpoons{H_2O} H^+_{(aq)} + ClO^-_{(aq)} \qquad K_a = 4.00 \times 10^{-8}$$

we can conclude that HF is stronger than HClO as its acidity constant is larger. Table 10.1 reports acidity constants for several acids and bases.

`Basicity constant` In a very similar way, bases also have what we call base dissociation constants: K_b. The bigger this value the stronger the base and the more hydroxyls will be produced. In the case of ammonia:

$$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH_4{}^+{}_{(aq)} + OH^-_{(aq)} \qquad K_b = \frac{[NH_4{}^+]\cdot[OH^-]}{[NH_3]}$$

remember liquid water can not be part of an equilibrium constant. If we compare the basicity constant of two different bases, for example, ammonia and hydrazine

$$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH_4{}^+{}_{(aq)} + OH^-_{(aq)} \qquad K_b = 1.80 \times 10^{-5}$$

$$N_2H_{4(aq)} + H_2O_{(l)} \rightleftharpoons N_2H_5{}^+{}_{(aq)} + OH^-_{(aq)} \qquad K_b = 1.30 \times 10^{-6}$$

we can conclude that NH_3 is stronger than N_2H_4 as its basicity constant is larger.

K_a and K_b Acidity and basicity are related concepts. As such, there is a relationship between the constant of acidity and the constant of basicity of an acid or base given by:

$$\boxed{K_a \cdot K_b = 1.0 \cdot 10^{-14}} \tag{10.1}$$

This relationship implies that all acids and bases have a constant of acidity as well as basicity. For example, if the constant of acidity of a given acid is 1.5×10^{-5}, the basicity constant of the same species would be 6.6×10^{-10}. If K_a is large and the acid is strong K_b must be small and the basic character of a strong acid is very weak. Further in the chapter, you will find a table listing some acidity and basicity constants.

In general, some simple rules predict the acid-base character of an acid-conjugate base pair. Strong acids in general produce weak conjugate bases. For example, HCl is a strong acid and its conjugate base (Cl^-) is a weak base. Differently, HF is a weak acid and its conjugate base (F^-) is a moderately strong base. Also, if pK_a is smaller than 7 we can consider that the chemical is more acidic than basic, and we call the chemical an acid. If pK_a is larger than 7 (and therefore pK_b is smaller than 7) we can consider that the chemical is more basic than acidic, and we call the chemical a base.

Table 10.1 Acidity and basicity constants at 25°C

Increasing acidic strength ↓ (both column groups)

Name	Formula	K_a	Name	Formula	K_a
Hypoiodous acid	HIO	3.20×10^{-11}		⋮	
Phenol	C_6H_5OH	1.00×10^{-10}	Fluoroacetic acid	CH_2FCO_2H	2.60×10^{-3}
Hydrocyanic acid	HCN	6.20×10^{-10}	Formic acid	CH_2O_2	1.80×10^{-4}
Hypobromous acid	$HBrO$	2.80×10^{-9}	Bromoacetic acid	CH_2BrCO_2H	1.30×10^{-3}
Hypochlorous acid	$HClO$	4.00×10^{-8}	Chloroacetic acid	CH_2ClCO_2H	1.30×10^{-3}
Benzoic acid	$C_6H_5CO_2H$	6.25×10^{-5}	Dichloroacetic acid	$CHCl_2CO_2H$	4.50×10^{-2}
Hydrazoic acid	HN_3	2.50×10^{-5}	Periodic acid	HIO_4	2.30×10^{-2}
Acetic acid	CH_3COOH	1.75×10^{-5}	Chlorous acid	$HClO_2$	1.10×10^{-2}
Iodoacetic acid	CH_2ICO_2H	6.60×10^{-4}	Trichloroacetic acid	CCl_3CO_2H	2.20×10^{-1}
Hydrofluoric acid	HF	6.30×10^{-4}	Trifluoroacetic acid	CF_3CO_2H	3.00×10^{-1}
Nitrous acid	HNO_2	5.60×10^{-4}	Iodic acid	HIO_3	1.70×10^{-1}
Cyanic acid	$HCNO$	3.50×10^{-4}	Chromic acid	$HCrO_4$	1.80×10^{-1}
	⋮				

Increasing basicity strength ↓ (both column groups)

Name	Formula	K_b	Name	Formula	K_b
Aniline	$C_6H_5NH_2$	7.40×10^{-10}		⋮	
Pyridine	C_5H_5N	1.70×10^{-9}	n-Butylamine	$C_4H_9NH_2$	4.00×10^{-4}
Hydroxylamine	NH_2OH	8.70×10^{-9}	Ethylamine	$C_2H_5NH_2$	4.50×10^{-4}
Hydrazine	N_2H_4	1.30×10^{-6}	Methylamine	CH_3NH_2	4.60×10^{-4}
Ammonia	NH_3	1.80×10^{-5}	tert-Butylamine	$(CH_3)_3CNH_2$	4.80×10^{-4}
Propylamine	$C_3H_7NH_2$	3.50×10^{-4}	Dimethylamine	$(CH_3)_2NH$	5.40×10^{-4}
	⋮				

Sample Problem 101

Indicate the strongest acid from:

$HF_{(aq)} \xrightleftharpoons{H_2O} H^+_{(aq)} + F^-_{(aq)}$ $\qquad K_a = 7.2 \cdot 10^{-4}$

$HNO_{2(aq)} \xrightleftharpoons{H_2O} H^+_{(aq)} + NO_2{}^-{}_{(aq)}$ $\qquad K_a = 4.5 \cdot 10^{-4}$

SOLUTION

The acid dissociation constant tells how strong is the acid, hence the larger K_a the stronger the acid. Comparing both values: K_a(HF)= $7.2 \cdot 10^{-4}$ and $K_a(HNO_2)$= $4.5 \cdot 10^{-4}$, HF is the stronger of both.

STUDY CHECK

Indicate the strongest of the following acids:

$$H_2PH_4^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + HPH_4^{2-}{}_{(aq)} \qquad K_a = 6.2 \cdot 10^{-8}$$

$$H^+_{(aq)} + HCO_3^{2-}{}_{(aq)} \rightleftharpoons H_2CO_{3(aq)} + H_2O_{(l)} \qquad K_b = 2.3 \cdot 10^{-8}$$

pK_a and pK_b As acidity and basicity constants tend to be very small numbers there is a convenient form to deal with these values. By using logarithms one can transform a very tinny value into a larger digit. Specifically, if we use the negative value of a logarithm then we can convert a negative power of ten into a larger positive value. For example,

$$-log(1 \times 10^{-5}) = 5$$

Using this mathematical trick, we define the pK_a and pK_b:

$$\boxed{PK_a = -log(K_a)} \text{ and } \boxed{PK_b = -log(K_b)} \tag{10.2}$$

For example, as the acidity constant of acetic acid is 1.75×10^{-5} its PK_a would be 4.74. Similarly, the basicity constant of aniline is 7.40×10^{-10}, and PK_b would be 9.13. We also use the following expression to convert pK_a and pK_b into K_a and K_b

$$\boxed{K_a = 10^{-PK_a}} \text{ and } \boxed{K_b = 10^{-PK_b}} \tag{10.3}$$

For example, as the PK_a of nitrous acid is 3.37, the acidity constant of this acid is 4.26×10^{-4}.

The conjugate seesaw The strength of an acid and its conjugate base are not independent. Strong acids produce conjugate bases that are moderately weak. Similarly, weak acids produce conjugate bases that are moderately strong. The same reasoning can be applied to bases. This idea is called the conjugate seesaw and helps predict why the conjugate salt of a weak base such as ammonia (NH_4Cl) gives a moderately strong acidic solution. This idea is summarized in the following relationship and we have that pK_a and pK_a are also related:

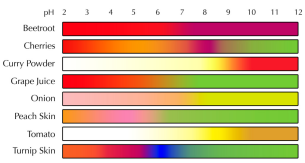
$$\boxed{PK_a + PK_b = 14} \tag{10.4}$$

▼A PH meter is used to measure the PH of solutions.

10.4 The PH scale

This section describes the PH scale that simply transforms a concentration value–oftentimes a very small number–into a simple round value. In short, the PH value tells you how many protons are there in a solution so that the larger PH the fewer protons are there in the solution. It also informs about the hydroxyl concentration, as protons and hydroxyls are connected using the dissociation equilibrium of water.

▼Color for different natural indicators depending on the pH.



Autoprotolysis of water and K_w Water is a weak electrolyte. The dissociation equilibrium of water is listed below:

$$\underset{\text{Acid}}{H_2O_{(l)}} + \underset{\text{Base}}{H_2O_{(l)}} \rightleftharpoons \underset{\text{Conjugate Base}}{OH^-_{(aq)}} + \underset{\text{Conjugate Acid}}{H_3O^+_{(aq)}} \qquad K_w$$

▼Over breathing causes alkalosis and the PH of blood increases from 7.4. The CO_2 level in the blood raises after breathing from a plastic bag. CO_2 is an acid and decreases the PH of blood.

This chemical equilibrium representing the dissociation of water is called the autoprotolysis of water. We have that water, as well as many other chemicals, that can act as a base or an acid, hence giving and receiving protons. We call these types of chemicals *amphiprotic*. Other examples of amphiprotic chemicals are $HCO_3{}^-$ or $HSO_4{}^-$. The dissociation constant of water is called the ion-product of water K_w:

$$H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)} \qquad \left[H^+\right] \cdot \left[OH^-\right] = 1.0 \cdot 10^{-14}$$

K_w is a constant that only depends on the temperature of the water being $1.0 \cdot 10^{-14}$ at 25°C. The ion product of water established a relationship between protons and hydroxyls in solution:

$$\boxed{\left[H^+\right] \cdot \left[OH^-\right] = 1.0 \cdot 10^{-14}} \tag{10.5}$$

`Protons and Hydroxyls` Acids and bases exist in solution with water. That means that as they produce protons or hydroxyls water receives these ions as it ionizes as well. Hence, the concentration of protons and hydroxyls in the solution is not independent. Indeed, the ion-product of water relates the concentration of protons ($\left[H^+\right]$) and the concentration of hydroxyls ($\left[OH^-\right]$):

$$\left[H^+\right] \cdot \left[OH^-\right] = 1.0 \times 10^{-14}$$

Water is neutral, which means that the concentration of protons ($\left[H^+\right]$) and the concentration of hydroxyls ($\left[OH^-\right]$) and both equal to $1.0 \cdot 10^{-7}$M. When we dissolve an acid or a base into the water, $\left[OH^-\right]$ and $\left[H^+\right]$ change drastically. When dissolving an acid, $\left[H^+\right]$ increases as acids produce protons, while $\left[OH^-\right]$ decreases. Differently, when dissolving a base, $\left[OH^-\right]$ increases, as bases produce hydroxyls, while $\left[H^+\right]$ decreases.

Sample Problem 102

The proton concentration in an acid solution is $7.0 \cdot 10^{-5}$M. Calculate $\left[OH^-\right]$.

SOLUTION

We will use Equation 10.5. The value given is $\left[H^+\right] = 7.0 \cdot 10^{-5}$M and the problem ask $\left[OH^-\right]$. Solving for $\left[OH^-\right]$ we have:

$$7.0 \cdot 10^{-5} \cdot \left[OH^-\right] = 1.0 \cdot 10^{-14}$$

Hence $\left[OH^-\right] = 1.4 \cdot 10^{-10}$M.

STUDY CHECK

The hydroxyl concentration in a basic solution is $2.3 \cdot 10^{-6}$M. Calculate the concentration of protons.

`The PH scale` The proton concentrations in aqueous solutions tend to be rather small. For example, the proton concentration in normal vinegar is $2 \cdot 10^{-3}$M. As it is hard to work with these small concentrations, scientists developed the PH scale that transforms $\left[H^+\right]$ into a larger number (see Figure 10.1). The formula for the PH is:

$$\boxed{PH = -log\left[H^+\right]} \tag{10.6}$$

The PH scale normally ranges from 0 to 14. PH values lower than 7 correspond to acidic solutions, whereas PH values larger than 7 correspond to basic solutions. Solutions with a PH of 7 are neutral. For example, the PH for vinegar is $-log\left(2 \cdot 10^{-3}\right)$ that is 2.69. However, it exists PH values out of the scale for very concentrated solutions. Examples of PH values and common chemicals are given in the figure below. An equivalent scale is also defined for the concentration of hydroxyls. The POH values are defined as:

$$\boxed{POH = -log\left[\mathrm{OH}^{-}\right]} \tag{10.7}$$

The POH scale also ranges from 0 to 14. POH values lower than 7 correspond to this time to basic solutions, whereas POH values larger than 7 correspond to acidic solutions. Solutions with a POH of 7 are neutral. The values of PH and POH are hence related by the following equation:

$$\boxed{PH + POH = 14} \tag{10.8}$$

For example, if the PH of a solution is 4 therefore the POH will be 10. Both indications suggest that the solution would be acidic.

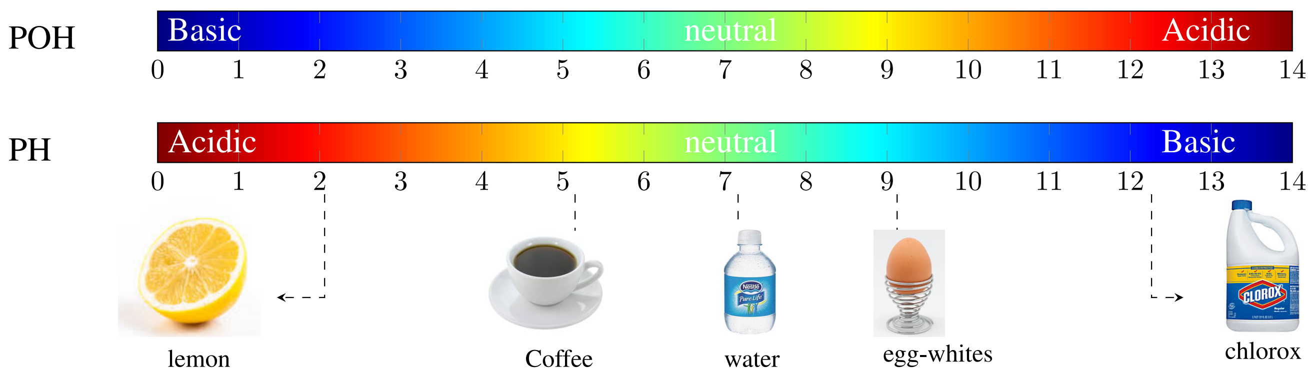


Figure 10.1 The PH scale

Sample Problem 103

Calculate the PH for: (a) an acid solution with proton concentration of $7.0 \cdot 10^{-5}$M (b) a basic solution with a hydroxyl concentration of $7.0 \cdot 10^{-5}$M.

SOLUTION

(a) We will use Equation 10.6. Given is $\left[\mathrm{H}^{+}\right] = 7.0 \cdot 10^{-5}$M and the problem ask for the PH. Solving for PH we have:

$$PH = -log\left(7.0 \cdot 10^{-5}\right)$$

and the results is 4.15. This is an acidic PH. (b) We will also use Equation 10.6. However, before doing that, we need to compute the concentration of protons. In order to do this we will use Equation 10.5 given $\left[\mathrm{OH}^{-}\right] = 8.0 \cdot 10^{-2}$M

$$\left[\mathrm{H}^{+}\right] \cdot 8.0 \cdot 10^{-2} = 1.0 \cdot 10^{-14}$$

We have $\left[\mathrm{H}^{+}\right] = 1.25 \cdot 10^{-13}$M. Now we can compute the PH. Solving for PH we have:

$$PH = -log\left(1.25 \cdot 10^{-13}\right)$$

and the results is 12.90. This is a basic PH.

STUDY CHECK

Calculate the PH for: (a) a basic solution with proton concentration of $3.0 \cdot 10^{-8}$M (b) a basic solution with a hydroxyl concentration of $2.0 \cdot 10^{-9}$M.

`From PH to proton concentration` At this point, we know that the PH quantifies the proton concentration of a solution. So given $[\mathrm{H}^+]$ we can calculate PH using the logarithm with the opposite sign. But what if we know the PH and we want to calculate the corresponding proton concentration? We can do this by using the formula:

$$\boxed{[\mathrm{H}^+] = 10^{-PH}} \tag{10.9}$$

To use the previous formula you need to use the power key in your calculator. For example, if the PH is 3.3 and we need to calculate the proton concentration you will need to type: $10\boxed{\wedge}\boxed{-}3.3$, and the result is $5.0 \cdot 10^{-4}$M. Mind that: (a) in some calculators, sometimes the power key looks like $\boxed{10^x}$; (b) you need to use the negative key and not the minus key. The minus sign is used for substations, the negative key is used for numbers. An equivalent relation exists between the concentration of hydroxyls and the POH:

$$\boxed{[\mathrm{OH}^-] = 10^{-POH}} \tag{10.10}$$

The diagram in Figure 10.2 displays some of the most important equations involved in this section:

Sample Problem 104

The PH of a solution is 4.5. Calculate the proton concentration of that solution.

SOLUTION

We will use Equation 10.9, given PH and asking $[\mathrm{H}^+]$.

$$[\mathrm{H}^+] = 10^{-PH} = 10^{-4.5}$$

and the results is $3.16 \cdot 10^{-5}$M.

STUDY CHECK

The PH of a solution is 9.5. Calculate the proton concentration of that solution.

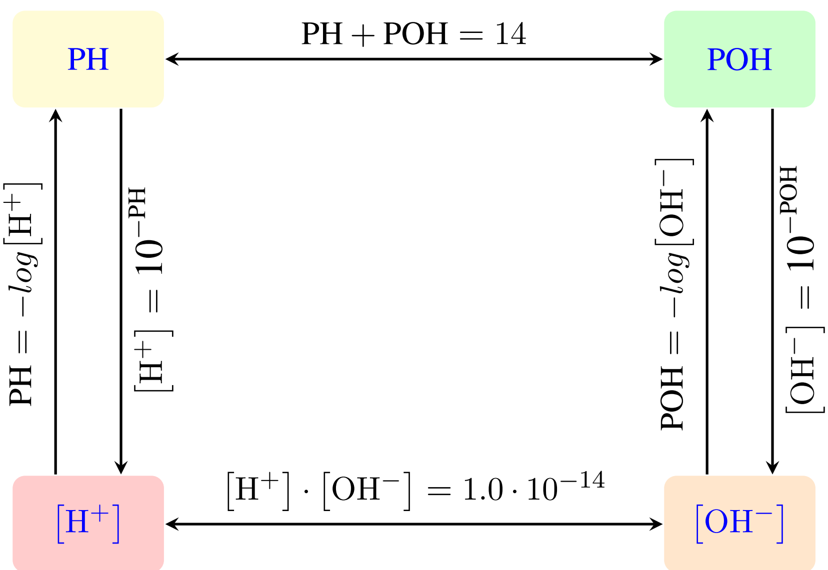


Figure 10.2 Diagram with formulas relating the PH, POH and the concentration of protons and hydroxyls in water.

10.5 Buffer solutions

We have previously addressed the properties of acids and bases. Buffers are specific solutions able to accommodate acids or bases without changing their PH. Buffers play a key role for example in our blood where a buffer system absorbs small quantities of acids and bases produced during biological reactions while keeping its PH constant. This section covers the properties of buffers. You will learn what are buffer, what are they made of. You will also learn how to compute the PH of a buffer system and the PH of a buffer after an acid or a base it is been added.

`Buffers` Buffers are solutions of an acid and a base. But not any kind of acid or base. Buffers are solutions of a weak acid with its conjugate base, or weak bases and its conjugate acid. For example, a mixture of 0.1M NH_3 and 0.1M NH_4Cl is a buffer. You can find acidic or basic buffers. For example, the previous example was a basic buffer, whereas a mixture of 0.1M CH_3COOH and 0.1M $NaCH_3COO$ is an acidic buffer. Acidic buffers contain a mixture of a weak acid and its conjugate base. Basic buffers contain a mixture of a weak base and its conjugate acid. Buffer function thanks to the equilibrium that links the acid and base so that when small quantities of acid or base are added the conjugate species contra rest this external action keeping the PH constant. Still, buffers have a limit of action, and if large quantities of external acid or bases are added the buffer equilibrium can be broken and they lose their capacity to modulate the PH.

`PH of a Buffer solution` A buffer solution consists of a solution containing both a weak electrolyte and its conjugate counterpart in the same or different concentration. For example, the PH of a 0.01M CH_3COOH/0.1M $NaCH_3COO$ ($K_a = 1.75 \times 10^{-5}$) acidic buffer can be computed using the following formula:

$$\boxed{PH = PK_a + \log\left(\frac{c_b}{c_a}\right)} \tag{10.11}$$

where:

PK_a is the PK of the acid in the buffer

c_a is the acid concentration in the buffer

c_b is the base concentration in the buffer

This formula is called the Henderson-Hasselbalch equation. Using the date above, we have that: $PH = 4.76 + \log\left(\frac{0.1}{0.01}\right) = 5.75$. The following example will further demonstrate how to calculate the PH of buffer solutions.

Sample Problem 105

Calculate the PH of 20mL of a 0.1M NH_4Cl/0.2M NH_3 ($K_b = 1.80 \times 10^{-5}$).

SOLUTION

This is a basic buffer and the main equilibrium involves ammonia, a weak base. In order to calculate the PH we need the concentration of the acid and base counter parts. The buffer volume is not important as it will be cancel out in the the Henderson-Hasselbalch equation. We would also need K_a, as we have K_b we can easily compute K_a, giving 5.5×10^{-10}. The final PH will be:

$$PH = 9.25 + \log\left(\frac{20 \cdot 0.2}{20 \cdot 0.1}\right) = 9.56$$

STUDY CHECK

Calculate the PH of a 0.2M HF/0.3M KF ($K_a = 6.30 \times 10^{-4}$).

10.6 Titrations

Titration is a chemical technique used to calculate the unknown molarity of an acid or base. It is based on the principle that acids neutralize bases and we can figure out the molarity of the unknown chemical (the titrate) by knowing the reacting amounts. A titration uses chemical equipment: a burette, Erlenmeyer, and an indicator (see Figure **??**). The unknown chemical is called the titrate and the known chemical is called the titrant. The goal of a titration is to calculate the volume of titrant needed to neutralize the titrate. We reach the endpoint of a titration when the titrant and titrate completely neutralize. At the end point, the mixture of titrant and titrate has a specific PH. Even though the chemical procedure in the lab is similar when titrating strong or weak acids or bases, the calculations needed to calculate the PH at the endpoint differ. This section will cover the principles and calculations involved in titrations.

`Neutralization Reactions` Titrations involve a neuralization reaction in which an acid neutralizes a base. Acids produce protons H^+ and bases hydroxyls OH^- that neutralize forming water, H_2O. More importantly, they react in very specific ratios. Let us take a look at the reaction of hydrochloric acid with sodium hydroxide to produce water and sodium chloride:

$$HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)} \qquad \text{Neutralization Reaction}$$

In this reaction, one mole of HCl reacts with one mole of NaOH. The fact that one more reacts with one more can be used as a principle for acid-base titration. We will have to use the stoichiometry of the reaction to calculate the volume of titrant needed to neutralize the titrate. Imagine you have an unknown sample of HCl and you need to know the amount of acid in the solution. If you know that this sample reacts with a specific amount of NaOH as you know that they react in a one-2-one ratio then you would know the acidic content. This is the idea behind titration: a laboratory procedure in which an unknown sample is neutralized with a known solution. A chemical *indicator*, which changes color depending on the acidity of the medium, is used to visually reveal the moment in which the acid and the base are completely neutralized. The point at which the indicator changes color is called the *equivalency point* or the *endpoint*. At the endpoint, the acid and the base are neutralized.

`Endpoint formula` At the *equivalence point*, also called the *stoichiometric point*, the moles of acid and the moles of the base are the same. A simple formula is extensively used to calculate the unknown acid concentration in a titration:

$$\boxed{n_H \cdot c_a \cdot V_a = n_{OH} \cdot c_b \cdot V_b} \qquad (10.12)$$

where:

$n_H \cdot c_a \cdot V_a$ and $n_{OH} \cdot c_b \cdot V_b$ is moles of protons and hydroxyls, respectively

c_a and V_a is acid concentration and volume respectively

c_b and V_b is base concentration and volume respectively

n_H and n_{OH} is the number of protons of the acid and hydroxyls of the base

Regarding the units in this formula, the units in V_a and V_b can either be L or mL. They just need to be the same units. This formula can be used for example when we titrate a given acid amount with a known base and we arrive at the volume of base needed to the endpoint to calculate the molarity of the acid. This formula can also be used when we titrate a known acid with a known base and we need to calculate the volume of titrant needed to reach the endpoint.

Equation 10.12 can also be used to identify if we already passed the endpoint in a titration. For example, we titrate 2mL of 3M H_2SO_4 (titrant) with 2mL of 1M NaOH (titrate). The question would be: are be before, after, or at the endpoint? We have that to neutralize completely the titrant (H_2SO_4), and using Equation 10.12 we would need:

$$2 \cdot 3M \cdot 2mL = 1 \cdot 1M \cdot V_b$$

that is we would need 12 mL of the base. Therefore, as we only used 2mL we would be before the endpoint and we would have not reached the endpoint.

Sample Problem 106

A 50mL sample of an unknown acid is neutralized with 25 mL of a NaOH 3M solution. Calculate the molarity of the unknown acid.

SOLUTION

We will use Equation 10.12, given: $c_b = 3$M, $V_b = 25mL$ and $V_a = 50mL$.

$$c_a \cdot 50mL = 3M \cdot 25mL$$

and the results is 1.5M.

STUDY CHECK

A 15mL sample of an unknown acid is neutralized with 45 mL of a NaOH 1M solution. Calculate the molarity of the unknown acid.

10.7 Blood as a buffer

Carbon dioxide acts as a buffer in the blood, hence regulating its PH. Here, we will discuss some chemical aspects that impact its role as buffer.

`Carbon dioxide is an acid` Carbon dioxide is an acid gas. When it dissolved reversibly in water it produces carbonic acid:

$$CO_{2(g)} + H_2O_{(l)} \rightleftharpoons H_2CO_{3(aq)}$$

At the same time H_2CO_3 partially dissociates in water as a weak acid producing bicarbonate–also known as hydrogen carbonate:

$$H_2CO_{3(aq)} + H_2O_{(l)} \rightleftharpoons HCO_3{}^-{}_{(aq)} + H_3O^+_{(aq)}$$

As a result, the dissolution of carbon dioxide in the blood produced a buffer containing carbonic acid and bicarbonate–the conjugate base of carbonic acid. This buffer system keeps the PH of blood constant to a value of 7.4 which is slightly above the neutral value of 7.

`The dangerous change in blood PH` Most biological molecules are very sensitive to PH changes. Small deviations in PH can be dangerous or even fatal. If the blood PH falls below 7.34 it becomes more acidic causing a clinical condition called *acidosis*. On the other hand, if the PH rises above 7.45, the blood becomes more basic causing a condition called *alkalosis*. Changes in PH below 6.8 or above 8.0 can cause death. Figure 10.3 displays the PH range that leads to acidosis and alkalosis.

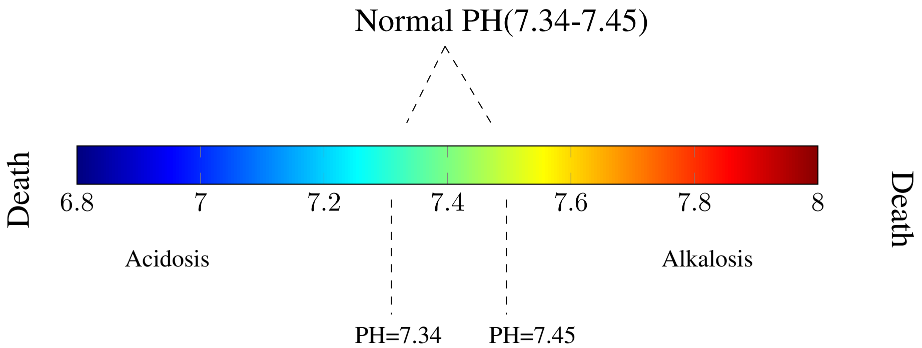


Figure 10.3 PH range that leads to acidosis and alkalosis

`Alkalosis and carbon dioxide` Our body needs oxygen to function properly. When you inhale, you introduce oxygen into the lungs. When you exhale, you release carbon dioxide, which is a waste product. Normally, the respiratory system keeps these two gases in balance. Respiratory alkalosis occurs when you breathe too fast or too deep. This hyperventilation–also known as over-breathing–causes carbon dioxide levels to drop too low. This means there is not enough CO_2 in the blood and that causes the PH to increase becoming too alkaline. Treating the condition is a matter of raising carbon dioxide levels in the blood. You can do this by breathing from a plastic bag. Respiratory acidosis occurs when too much CO_2 builds up in the body. Normally, the lungs remove CO_2 while you breathe. However, sometimes your body is unable to get rid of enough CO_2. Treatments for this condition are usually designed to help your lungs. For example, you may be given drugs to dilate your airway. You might also be given oxygen or a continuous positive airway pressure (CPAP) device. The CPAP device can help you breathe if you have an obstructed airway or muscle weakness.

Sample Problem 107

Explain why CO_2 decreases the PH of blood.

SOLUTION

CO_2 is an acid and when dissolved in water it generates protons, H^+. Increasing the amount of protons decreases PH as PH is a measure the concentration of protons in solution.

STUDY CHECK

The PH of a patient's blood sample is 7.3. Indicate whether the patient suffers alkalosis, acidosis, or perhaps none of the affections. Can the PH level cause death?

CHAPTER 10

The nature of acids and bases

10.1 Classify the following species as Arrhenius acids, or bases: (a) H_2SO_4 (b) NaOH (c) HNO_3 (d) HCl (e) $Ca(OH)_2$

10.2 Classify the following species as Arrhenius acids, or bases: (a) KOH (b) H_3PO_4 (c) HI (d) $Mg(OH)_2$

10.3 Classify the following species as Brönsted-Lowry acids, bases or both: (a) H_2O (b) OH^- (c) NH_3 (d) $NO_2{}^-$

10.4 Classify the following species as Brönsted-Lowry acids, bases or both: (a) $HCO_3{}^-$ (b) HI (c) HCN (d) $HSO_4{}^-$ (e) HCOONa

Dissociations of acids & bases

10.5 From the following pairs, select the strongest acid: (a) HIO_3 ($K_a = 1.6 \cdot 10^{-1}$) or H_2SO_3 ($K_a = 1.5 \cdot 10^{-2}$) (b) HN_3 ($K_a = 1.9 \cdot 10^{-5}$) or H_2CO_3 ($K_a = 4.3 \cdot 10^{-7}$)

10.6 From the following pairs, select the strongest base: (a) CN^- ($K_a = 6.2 \cdot 10^{-10}$) or H_2O ($K_a = 1.0 \cdot 10^{-14}$) (b) $H_2C_6H_5O^-$ ($K_a = 1.8 \cdot 10^{-5}$) or HCOOH ($K_a = 1.7 \cdot 10^{-4}$)

10.7 Write down the formula of the conjugate bases: (a) H_2O (b) HCl (c) HNO_3 (d) H_2SO_4

10.8 Write down the formula of the conjugate bases: (a) $HSO_4{}^-$ (b) H_2S (c) HCOOH (d) $H_2PO_4{}^-$

10.9 Write down the formula of the conjugate acids: (a) $NH_4{}^+$ (b) CH_3COO^- (c) HS^- (d) CN^-

10.10 Identify the conjugate acid-base pairs: (a) $HCl_{(g)} + H_2O_{(l)} \rightleftharpoons Cl^-_{(aq)} + H_3O^+_{(aq)}$ (b) $CH_3COO^-_{(l)} + HCl_{(g)} \rightleftharpoons Cl^-_{(aq)} + CH_3COOH_{(aq)}$ (c) $CO_3{}^{2-}{}_{(aq)} + HCN_{(g)} \rightleftharpoons CN^-_{(aq)} + HCO_3{}^-{}_{(aq)}$ (d) $HNO_{3(aq)} + OH^-_{(aq)} \rightleftharpoons NO_3{}^-{}_{(aq)} + H_2O_{(aq)}$

10.11 Write down the following dissociation or acid-base reaction involving the exchange of one proton: (a) $HNO_{3(l)} + H_2O_{(l)} \rightleftharpoons$ (b) $H_2SO_{4(l)} + H_2O_{(l)} \rightleftharpoons$ (c) $HCl_{(g)} + H_2O_{(l)} \rightleftharpoons$ (d) $NH_{3(g)} + H_3O^+_{(aq)} \rightleftharpoons$

10.12 Write down the following dissociation or acid-base reaction involving the exchange of one proton: (a) $CO_3{}^{2-}{}_{(aq)} + HCN_{(g)} \rightleftharpoons$ (b) $HCO_3{}^-{}_{(aq)} + HCN_{(g)} \rightleftharpoons$ (c) $HCO_3{}^-{}_{(aq)} + OH^-_{(aq)} \rightleftharpoons$

The PH scale

10.13 Answer the following questions: (a) The proton concentration of a solution is 3×10^{-3}M. Calculate the hydroxyl concentration in the same solution (b) The hydroxyl concentration of a solution is 8×10^{-6}M. Calculate the proton concentration in the same solution.

10.14 Answer the following questions: (a) The proton concentration of a solution is 1×10^{-10}M. Calculate the hydroxyl concentration in the same solution (b) The hydroxyl concentration of a solution is 3×10^{-5}M. Calculate the proton concentration in the same solution.

10.15 Answer the following questions: (a) The PH of a solution is 1.34. Calculate the POH of the same solution (b) The POH of a solution is 12. Calculate the PH of the same solution.

10.16 Answer the following questions: (a) The PH of a solution is 3. Calculate the POH of the same solution (b) The POH of a solution is 13. Calculate the PH of the same solution.

10.17 Answer the following questions: (a) The PH of a solution is 1.56. Calculate the concentration of protons. (b) The POH of a solution is 10.34. Calculate the hydroxyl concentration in the same solution. (c) The PH of a solution is 12.4. Calculate the hydroxyl concentration in the same solution.

10.18 Fill the table below:

$[H^+]$	$[OH^-]$	Acidic/Basic/Neutral?
1.5×10^{-4}	6.6×10^{-11}	
4.9×10^{-12}	2.0×10^{-3}	
1.9×10^{-6}	5.3×10^{-9}	
1.0×10^{-7}	1.0×10^{-7}	

10.19 Fill the table below:

$[H^+]$	$[OH^-]$	**PH**	**POH**
4.5×10^{-3}		–	–
	3.2×10^{-7}	–	–
–	–	5.1	
–	–		6.9

10.20 Fill the table below:

$[H^+]$	$[OH^-]$	**PH**	**POH**
3.5×10^{-1}	–	–	
–	–	–	2

Buffer solutions

10.21 Which of the following mixtures of solutions can act as a buffer: (a) 0.1M-H_2SO_4/0.1M-$Na(SO_4)_2$ (b) 0.01M-NH_3/0.01M-NH_4Cl (c) 0.2M-HNO_2/0.2M-$NaNO_2$

10.22 Which of the following mixtures of solutions can act as a buffer: (a) 0.4M-H_2SO_4/0.1M-$NaHSO_4$ (b) 0.23M-HCl/0.20M-KCl (c) 0.56M-HCN/0.22M-NaCN

10.23 Which of the following mixtures of solutions can act as a buffer: (a) KCl/HCl (b) H_2SO_4/$NaHSO_4$ (c) Na_2SO_4/$NaHSO_4$

10.24 Which of the following mixtures of solutions can act as a buffer: (a) H_3PO_4/NaH_2PO_3 (b) Na_3PO_4/NaH_2PO_3 (c) HNO_2/$NaNO_2$

10.25 Which of the following substances can be mixed together to prepare a buffer: (a) CH_3COOH (b) CH_3COONa (c) NaOH (d) HCl (e) NaCl

10.26 Which of the following substances can be mixed together to prepare a buffer: (a) NH_3 (b) NH_4Cl (c) Na_2SO_4 (d) H_2SO_4 (e) NaCl (f) HCl

10.27 Calculate the PH of the following buffers: (a) 0.15 M-HCN/0.35M-NaCN, $K_a = 6.20 \times 10^{-10}$ (b) 0.15 M-HCN/0.15M-NaCN, $K_a = 6.20 \times 10^{-10}$ (c) 0.25 M-HCN/0.15M-NaCN, $K_a = 6.20 \times 10^{-10}$

10.28 Calculate the PH of the following buffers: (a) 0.25 M-NH_3/0.45M-NH_4Cl, $K_b = 1.80 \times 10^{-5}$ (b) 0.15 M-HNO_2/0.05M-$NaNO_2$, $K_a = 5.60 \times 10^{-4}$

10.29 Calculate the PH of the following buffers: (a) 0.02 M-$C_6H_5NH_2$/0.05M-C_6H_5NHCl, $K_b = 7.40 \times 10^{-10}$ (b) 0.4 M-HCNO/0.5M-NaCNO, $K_a = 3.50 \times 10^{-4}$

10.30 The PH of a $C_6H_5NH_2$/C_6H_5NHCl ($K_b = 7.40 \times 10^{-10}$) buffer is 4.0. Calculate the ratio of $[C_6H_5NH_2]/[C_6H_5NHCl]$.

Titrations

10.31 Solve the following titration scenarios: (a) In a titration experiment, 13.5 mL of 0.34 M HCl neutralize 34.3 mL of KOH. What is the concentration of the KOH solution? (b) In a titration experiment, 20.4 mL of 0.10 M HCl neutralize 12.4 mL of $Ca(OH)_2$. What is the concentration of the base solution?

10.32 Solve the following titration scenarios: (a) In a titration experiment, 10.4 mL of 0.20 M H_2SO_4 neutralize 8.4 mL of $Ca(OH)_2$. What is the concentration of the base solution? (b) In a titration experiment, 12.5 mL of 0.23 M H_3PO_4 neutralize 4.8 mL of $Ca(OH)_2$. What is the concentration of the base solution?

10.33 A 5mL sample of a monoprotic acid neutralizes 10mL of a 0.01M-KOH solution. Calculate the molarity of the acid.

10.34 A 10mL sample of a monoprotic acid neutralizes 5mL of a 0.2M-$Ca(OH)_2$ solution. Calculate the molarity of the acid.

10.35 A 5 grams sample of an acid neutralizes 30mL of a 0.5M-NaOH solution. Calculate the molar mass of the acid.

10.36 You prepare a base solution by dissolving 3-g of the base into 250mL of solution. Five milliliters of this solution neutralizes 5mL of a 0.3MH_2SO_4 solution. Calculate the molar mass of the base.

Ch. 11. Organic Chemistry

ORGANIC chemistry is a vast subject that studies the properties of carbon-based compounds, and organic compounds are chemicals based mostly on carbon as well as hydrogen. On one hand, organic chemistry plays a key role in the understanding of the functioning and composition of living cells. On the other hand, many products of industrial organic chemistry such as plastics, fuels, perfumes, or prescription drugs are an accepted part of our everyday lives. These types of compounds have common organic properties, from their smells to their powerful action. Key organic chemistry concepts are the naming of organic compounds and the properties of functional groups that give unique properties to molecules such as caffeine or even addictive drugs such as cocaine.

11.1 Organic and non-organic compounds

Organic compounds are chemicals mostly made of Carbon and Hydrogen. Here a few examples of organic compounds: CH_4, C_2H_6, or C_6H_6. Differently, inorganic compounds contain other elements different than Carbon and Hydrogen. Examples of inorganic compounds are: NaCl, CO_2, HCl, FeO or NaOH. Mind that carbonates (Na_2CO_3), carbon monoxide (CO) or carbon dioxide (CO_2) are not organic compounds. Alkanes are simple organic compounds made of Carbon and Hydrogen with all carbons connected using simple bonds–these are single lines to represent the connections between atoms. The simplest alkane and the simplest organic compound is methane: CH_4, a fuel and the main constituent of natural gas.

$$\begin{array}{ccccc} & & \mathrm{H} & & \\ & & | & & \\ \mathrm{H} & - & \mathrm{C} & - & \mathrm{H} \\ & & | & & \\ & & \mathrm{H} & & \end{array} \qquad \text{Methane, } CH_4$$

Its structure is very representative of organic compounds in general as it shows that each carbon atom in an organic molecule is connected to four different atoms.

11.2 Alkanes

This first section will introduce organic chemistry, covering the most simple organic compounds: the alkanes. Alkanes–also called hydrocarbons–are simply made of carbon and hydrogen with all C-C bonds being single bonds. First, you will be introduced to a few organic chemicals and you will learn about a series of different organic formulas that can represent the same compound. Then, you will learn the basic naming rule of alkanes,

▼Methane (CH_4) is used as a fuel for ovens, homes, water heaters.

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which extend to other–more complex–organic chemicals.

`Molecular formula for alkanes` The naming of alkanes results from the combination of a prefix and a suffix. On one hand, the suffix is always *ane*. On the other hand, the prefix depends on the number of carbons in the molecule. Table **??** shows a list of the different prefixes. For example, the alkane with a single carbon is called methane (CH_4). Other examples of alkanes are ethane that contains two carbons (C_2H_6) or propane with three carbons (C_3H_8). The molecular formula for an alkane with n carbon atoms is:

$$C_nH_{2n+2} \tag{11.1}$$

Hence, we have that the molecular formula for methane ($n = 1$) is CH_4 and the molecular formula for octane ($n = 8$) is C_8H_{18}. *Molecular formulas* represent only the molecular compositions, showing only the elements in the molecule.

▼Hydrocarbons, made of carbon and hydrogen, tend to be gaseous molecules.

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Sample Problem 108

Write down the molecular formula for decane and pentane.

SOLUTION

Using Equation 11.1 we have that the molecular formula for decane ($n = 10$) would be: $C_{10}H_{22}$. Similarly, the molecular formula for pentane ($n = 5$) would be: C_5H_{12}.

▶Answer: (decane) $C_{10}H_{22}$; (pentane) C_5H_{12}

STUDY CHECK

Name the alkane with formula C_7H_{16} and give the molecular formula for nonane.

▼The octane rating of gas is a standard measure of the performance of engine fuels, originally determined by mixing a gasoline made entirely of heptane and 2,2,4-trimethylpentane (a highly branched octane).

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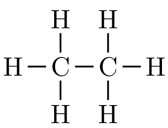
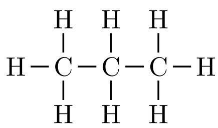
Table 11.1 Prefixed for alkane naming

# Carbons	prefix	# Carbons	prefix
1	Meth	6	Hex
2	Eth	7	Hepta
3	Prop	8	Octa
4	But	9	Nona
5	Pent	10	Deca

`Expanded structural formula for alkanes` At this point have addressed the molecular formulas of three simple alkanes: methane (CH_4), ethane (C_2H_6) and propane (C_3H_8). However, these type of formulas does not inform about the structure of the molecule, that is how the atoms are connected. The structure of ethane and propane is shown below:



Ethane, C_2H_6



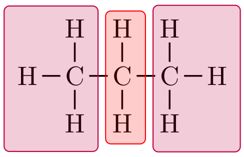
Propane, C_3H_8

These structures show that each carbon is connected to four different atoms and each hydrogen is only connected to one atom. The formulas presented above are called *expanded structural formulas* as all atoms and all bonds connecting these atoms are shown, hence representing the structure of the molecule.

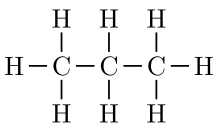
Table 11.2 Alkane names based on the number of carbons in the chain

Number of Carbons	Name	Condensed Structural formula	Molecular formula
1	Methane	CH_4	CH_4
2	Ethane	CH_3-CH_3	C_2H_6
3	Propane	$CH_3-CH_2-CH_3$	C_3H_8
4	Butane	$CH_3-CH_2-CH_2-CH_3$	C_4H_{10}
5	Pentane	$CH_3-CH_2-CH_2-CH_2-CH_3$	C_5H_{12}
6	Hexane	$CH_3-CH_2-CH_2-CH_2-CH_2-CH_3$	C_6H_{14}
7	Heptane	$CH_3-CH_2-CH_2-CH_2-CH_2-CH_2-CH_3$	C_7H_{16}
8	Octane	$CH_3-CH_2-CH_2-CH_2-CH_2-CH_2-CH_2-CH_3$	C_8H_{18}
9	Nonane	$CH_3-CH_2-CH_2-CH_2-CH_2-CH_2-CH_2-CH_2-CH_3$	C_9H_{20}

`Condensed structural formula` Let us analyze the formula of propane shown below. In this molecule, there are two different types of carbons: the end of the chain carbon, in the left and the right of the structure, and a central carbon. The extremes are bounded to three hydrogens (and one carbon), whereas the central atoms are bonded to two hydrogens (and two carbons).



The extremes of the chain are indeed CH_3 units and the central carbon is a CH_2 unit. So, another way to represent propane would be:

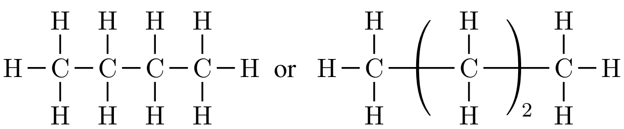


Expanded propane, C_3H_8

$$CH_3-CH_2-CH_3$$

Condensed propane, C_3H_8

This representation is called *condensed structural formula* or simple condensed structure, as the units of carbon and hydrogen are condensed into CH_2 and CH_3 units. Table 11.2 displays some molecular formula as well as condensed structures for different alkanes. For larger alkenes is often convenient to use parenthesis to simplify the structure. For example, the expanded structure of butane can be written as:



Similarly, the condensed structure of butane can be written as:

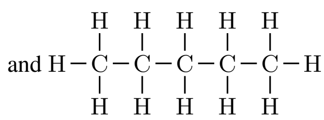
$$CH_3-CH_2-CH_2-CH_3 \quad \text{or} \quad CH_3 \left(CH_2 \right)_2 CH_3$$

Sample Problem 109

Write down the condensed and expanded formulas for pentane.

SOLUTION

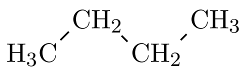
Pentane has five carbons, hence its condensed formula will have two CH_3 units and three CH_2 units. On the other hand, the expanded formula for pentane should display all carbons and hydrogens:

▶ Answer: $CH_3-CH_2-CH_2-CH_2-CH_3$ and 

STUDY CHECK

Write down the condensed and expanded formulas for heptane.

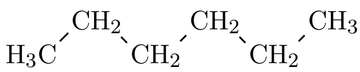
Skeletal structural formula Let us analyze the structure of butane again. This molecule has four carbons in the form of a C-C chain such as C-C-C-C.

 Butane, C_4H_{10}

However, in reality, C-C chains with more than two carbons are not linear, and their structure resembles more a zig-zag rather than a line. So instead of representing butane as a line it should be represented as:

 Butane, C_4H_{10}

This type of structural representation is called *skeletal structural formula* as you only represent the C-C skeleton of the molecule. Another example would be:



condensed formula of hexane

skeletal formula of hexane

The ending of a skeletal formula represents a CH_3 and the points in between represents CH_2.

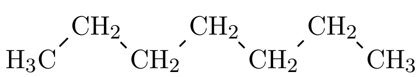
A review of the different structural formulas At this point, you have seen four different ways to represent organic molecules. Using propane as example, here all the formulas:We have that the *molecular formula* (e.g. C_3H_8 for propane) is mainly used to indicate the composition of the molecule in the form of Carbon and Hydrogen atoms. A second way to represent propane is using its *expanded structural formula*, that is by representing all atoms in the molecule and all atomic connections. A third molecular representation is the *condensed structural formula* that uses CH_3 and CH_2 units, only representing the C-C bonds. Finally, the *skeletal formula* is perhaps the most simplistic representation as only the C-C bonds are represented in the form of simple lines. It is important to understand that *all formulas are just different ways to represent the same molecule*.

Sample Problem 110

Write down the condensed and skeletal formulas for heptane.

SOLUTION

Heptane has seven carbons, hence its condensed formula will have two CH_3 units and five CH_2 units:

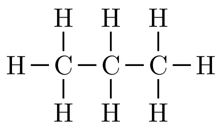


The skeletal formula for would be:

▶Answer: 

STUDY CHECK

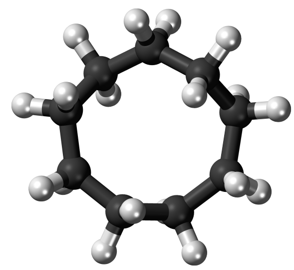
Draw the skeletal formula of decane.

C_3H_8 | $CH_3-CH_2-CH_3$ |  | 

Molecular formula | Condensed structural formula | Expanded structural formula | Skeletal structural formula

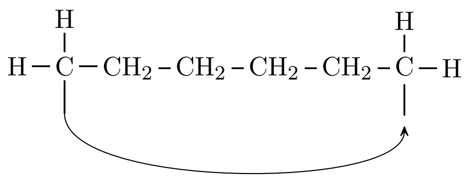
11.3 Cycloalkanes

▼3D representation of Cyclononane

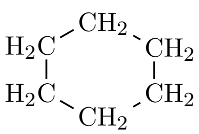
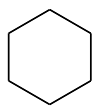


Alkanes are perfect examples of hydrocarbons, with C-C chains and all carbon atoms saturated with hydrogen. Cycloalkanes are simply alkanes with a cyclic structure. We will cover the molecular, condensed, and skeletal formulas for these chemicals.

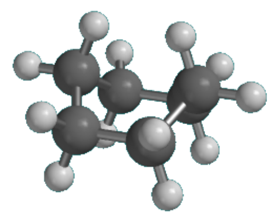
`Cyclic alkanes` Consider the expanded structure of hexane. A cycloalkane results from removing the left and right hydrogen while connecting the molecule in the form of a cycle:



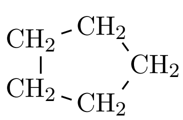
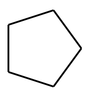
As the most stable structure for six lines is the hexagon, the resulting structure of cyclohexanes would be:

 or  cyclohexane

▼3D representation of Cyclohexane with the shape of a chair.

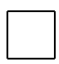


`Naming cycloalkanes` The naming of alkanes and cycloalkanes is very similar. You just need to add the cyclo prefix to the name. For example, the alkane with five carbons is called pentane, whereas the corresponding cycloalkane is called cyclopentane:

 or  Cyclopentane

Sample Problem 111

Write down the condensed structure and name the following cycloalkane:



SOLUTION

The cycloalkane has four carbons and its name is cylcobutane. Its condensed structure is

$$\begin{array}{l} CH_2 - CH_2 \\ \;\;| \qquad\;\; | \\ CH_2 - CH_2 \end{array}$$

▶Answer: $CH_2 - CH_2$ / $CH_2 - CH_2$

STUDY CHECK

Write down the condensed structure and name the following cycloalkane:

Molecular formulas for cycloalkanes The molecular formula for a general cycloalkane with n carbon atoms is:

$$C_nH_{2n} \tag{11.2}$$

As an example, the formula for cyclopropane ($n = 4$) is C_4H_8 and the formula for cycloctane ($n = 8$) is C_8H_{16}.

Sample Problem 112

Write down the molecular formula for cyclodecane and cyclopentane.

SOLUTION

Using Equation 11.2 we have that the molecular formula for cyclodecane ($n = 10$) would be: $C_{10}H_{20}$. Similarly, the molecular formula for cyclopentane ($n = 5$) would be:

▶Answer: C_5H_{10}.

STUDY CHECK

Name the alkane with formula C_7H_{14} and give the formula for cyclononane.

11.4 Alkanes with substituents

Oftentimes alkanes have other groups of atoms called substituents attached to the hydrocarbon chain. This section covers the naming of alkanes with substituents. Here is an example of an alkane and an alkane with a substituent:

$$\begin{array}{l} H \\ | \\ CH_2 - CH_2 - CH_3 \end{array} \qquad\qquad \begin{array}{l} NO_2 \\ | \\ CH_2 - CH_2 - CH_3 \end{array}$$

Propane   Nitropropane

In the substituted molecule, a nitro group has replaced a hydrogen atom.

Substituents There are many different substituents–also called groups–that can be found attached to an alkane chain. Their names are indicated in Table 11.2. The easiest substituents are halogens; atoms of chorine (Cl —), bromine (Br —) or iodine (I —) can replace hydrogen atoms in an alkane. The name of these substituents–chloro, bromo and iodo– resembles the name of the corresponding atom. Other substituents can contain carbon, like a methyl (CH_3 —) or a ethyl (CH_3CH_2 —). There are even more complex substituents such as tert-butyl in which a central carbon atom is connected to three different methyl groups. The name of substituents (methyl) comes from the name of the alkane (methane) by replacing the *-ane* suffix with *-yl*.

`Alkanes with a single substituent` Let us consider the following example. The condensed formula for propane is

$$CH_3-CH_2-CH_3 \qquad \text{Propane}$$

Now, this would be propane with a substituent:

$$\overset{\displaystyle Br}{\overset{|}{CH_2}}-CH_2-CH_3 \qquad \text{Bromopropane}$$

As you can see a bromine atom substitutes one of the hydrogen atoms of the second first of the molecule (starting from the left). The name simply results of the combination of the substituent group and the alkane name.

Table 11.3 Substituent name

Substituents	name	Substituents	name		
CH_3-	Methyl	$F-$	Fluoro	$CH_2=CH-$	Vinyl
CH_3CH_2-	Ethyl	$Cl-$	Chloro	NO_2-	Nitro
$CH_3CH_2CH_2-$	Propyl	$Br-$	Bromo	NH_2-	Amino
$CH_3CH_2CH_2CH_2-$	Butyl	$I-$	Iodo	$H_3C-\overset{\vert}{C}H-CH_2CH_3$	sec-Butyl
$H_3C-\underset{\underset{\displaystyle CH_3}{\vert}}{\overset{\vert}{C}}-CH_3$	t-butyl	$H_3C-\overset{\vert}{C}H-CH_3$	Isopropyl	$CH_3CH_2CH_2-$	Propyl

`Alkanes with two or more equal substituents` In the same way as when you have a single bromo substituent attached to propane, you can also have two $Br-$. In this case, you need to use the prefix *di* to indicate there are two identical bromos. For example, the name of the following molecule would be:

$$\underset{\underset{\displaystyle Br}{|}}{\overset{\overset{\displaystyle Br}{|}}{CH}}-CH_2-CH_3 \qquad \text{Dibromopropane}$$

Similarly, you should use the prefix *tri* and *tetra* for three equal substituents.

Sample Problem 113

Name the following hydrocarbon:

$$CH_3-CH_2-CH_2-\underset{\underset{\displaystyle I}{|}}{\overset{\overset{\displaystyle I}{|}}{CH}}$$

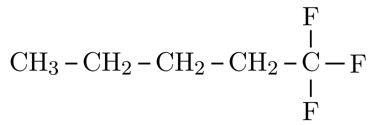
SOLUTION

The carbon chain has four carbons and hence the ending of the name would be: butane. Also there are two iodines (iodo substituents) attached to the carbon chain. As there are two of the same iodo atoms, we need to use the prefix *di*. The full name would be:

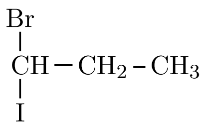
▶Answer: Dioiodobutane.

STUDY CHECK

Name the following hydrocarbon:



Alkanes with different substituents Now imagine you have two different halogens as substituents: Br — and I — like in next example



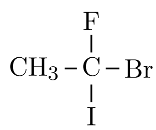
As both substituents have different names you cannot use the prefix *di*. Still, when you indicate the names of the substituents you need to order them alphabetically. So Bromo goes first in the name and Iodo after. You also need to separate the different substituents with a hyphen ('-'). The final name of the hydrocarbon above would be Bromo-Iodopropane.

Naming rules for alkanes Overall, the rules to name alkanes are:

1. **Step one:** The number of carbon atoms in the chain will give the ending name of the molecule (e.g. four carbons would be butane).
2. **Step two:** Number the main chain starting at the end closest to the substituents so that the numbers for the substituents are small.
3. **Step three:** Name the substituents with their position and order them alphabetically (di, tri etc. do not count in the alphabetic order).

Sample Problem 114

Name the following hydrocarbon:



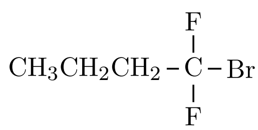
SOLUTION

The carbon chain has two carbons and hence the ending of the name would be: ethane. Also there are three different substituents: iodine (iodo substituents), bromine (bromo substituents) and fluorine (fluoro substituents). We need to order them according to the *abc*, hence the order would be: bromo, then fluoro and finally iodo. The full name of the alkane would be:

▶Answer: Bromo-Fluoro-Iodoethane.

STUDY CHECK

Name the following hydrocarbon:



`Numbering the chain` Substituents are atoms or groups of atoms replace hydrogens in alkane chains. You could envision attaching substituents at different points of the chain. For example:

$$\begin{array}{ccccc} & & Br & & \\ & & | & & \\ CH_3 & - & CH & - & CH_3 \end{array} \quad \text{or} \quad \begin{array}{ccccc} Br & & & & \\ | & & & & \\ CH_2 & - & CH_2 & - & CH_3 \end{array}$$

In the right example, Br is plugged to the left C atom, whereas in the left example C is plugged into the middle carbon. Hence, it is important first, to learn how to number a hydrocarbon chain. Let us use propane–a molecule with three atoms–as an example. In order to number the chain, you start by selecting the extreme that is the closest to the substituent using this carbon as number one. Next carbon would be carbon number two and so on until you arrive at carbon number three.

$$\begin{array}{ccccc} Br & & & & \\ | & & & & \\ CH & - & CH_2 & - & CH_3 \\ 1 & & 2 & & 3 \end{array}$$

As the Br atom in the carbon number one, the name of the molecule would be: 1-bromopropane or simply bromopropane. Differently, when the substituent is in a carbon different than one, you need to indicate that location. For example:

$$\begin{array}{ccccc} & & Br & & \\ & & | & & \\ CH_3 & - & CH & - & CH_3 \end{array} \quad \text{2-Bromopropane}$$

Sample Problem 115

Name the following hydrocarbon:

$$\begin{array}{ccccccc} & & CH_3 & & & & \\ & & | & & & & \\ CH_3 & - & CH & - & CH & - & Br \\ & & & & | & & \\ & & & & I & & \end{array}$$

SOLUTION

First we find the ending of the name: as the molecule has three carbons in the main chain, the ending of the name would be: propane. Then we need to number the chain so that the number one carbon is the closest to the substituents:

$$\begin{array}{ccccccc} & & CH_3 & & & & \\ & & | & & 1 & & \\ CH_3 & - & CH & - & CH & - & Br \\ 3 & & 2 & & | & & \\ & & & & I & & \end{array}$$

A methyl is connected to carbon two, and two halogens, a iodo and a bromo are connected to carbon number one. The substituents are: 2-methyl, 1-bromo, 1-iodo. If we order them: 1-bromo-1-iodo-2-methyl. And the final name would be:

▶Answer: 1-bromo-1-iodo-2-methylpropane.

STUDY CHECK

Name the following hydrocarbon:

$$\begin{array}{ccccccc} & & CH_3 & & & & \\ & & | & & & & \\ CH_3 & - & C & - & CH_2 & - & I \\ & & | & & & & \\ & & F & & & & \end{array}$$

Finding the longest chain Alkanes are complex molecules and often time they contain more than one hydrocarbon chain. Therefore, one can envision several ways to number the chain. However, the rule is to first locate the longest chain. Let us use the following hydrocarbon. How many chains can you find, and which is the longest chain?

$$\begin{array}{lllllll} & & & & \mathrm{CH_2} & - & \mathrm{CH_2} - \mathrm{CH_2} - \mathrm{CH_3}\\ & & & & | & & \\ \mathrm{CH_3} & - & \mathrm{CH_2} & - & \mathrm{CH} & - & \mathrm{CH_3}\end{array}$$

The answer should be three chains. Let me highlight the three different possibilities and the number of atoms in each chain:

$$\begin{array}{lllllll} & & & & \mathrm{CH_2} & - & \mathrm{CH_2} - \mathrm{CH_2} - \mathrm{CH_3}\\ & & & & | & & \\ \mathrm{CH_3} & - & \mathrm{CH_2} & - & \mathrm{CH} & - & \mathrm{CH_3}\end{array}$$

4 Carbons

$$\begin{array}{lllllll} & & & & \mathrm{CH_2} & - & \mathrm{CH_2} - \mathrm{CH_2} - \mathrm{CH_3}\\ & & & & | & & \\ \mathrm{CH_3} & - & \mathrm{CH_2} & - & \mathrm{CH} & - & \mathrm{CH_3}\end{array}$$

7 Carbons

$$\begin{array}{lllllll} & & & & \mathrm{CH_2} & - & \mathrm{CH_2} - \mathrm{CH_2} - \mathrm{CH_3}\\ & & & & | & & \\ \mathrm{CH_3} & - & \mathrm{CH_2} & - & \mathrm{CH} & - & \mathrm{CH_3}\end{array}$$

6 Carbons

As the longest chain has seven carbons, the name of the molecule would be heptane. Still, you need to add the substituents to the name. Hence, after you locate the longest chain you need to number the chain so that the substituents are located the closest to the carbon number one the possible:

$$\begin{array}{lllllllllll} & & & & 4 & & 5 & & 6 & & 7\\ & & & & \mathrm{CH_2} & - & \mathrm{CH_2} & - & \mathrm{CH_2} & - & \mathrm{CH_3}\\ & & & & | & & & & & & \\ \mathrm{CH_3} & - & \mathrm{CH_2} & - & \mathrm{CH} & - & \mathrm{CH_3} & & & & \\ 1 & & 2 & & 3 & & & & & & \end{array}$$

As there is a methyl in carbon number three the final name of the molecule would be: 3-methylheptane.

Naming rules for branched alkanes Overall, the rules to name branched alkanes are:

1. **Step one:** Look for the longest carbon-carbon chain that will give the ending name of the molecule (e.g. four carbons would be butane).
2. **Step two:** Number the main chain starting at the end closest to the substituents so that the numbers for the substituents are small.
3. **Step three:** Name the substituents with their position and order them alphabetically.

Sample Problem 116

Name the following hydrocarbon:

$$\begin{array}{lllllll} & & & & \mathrm{CH_3} & & \\ & & & & | & & \\ \mathrm{CH_3} & - & \mathrm{CH_2} & - & \mathrm{C} & - & \mathrm{CH_3}\\ & & & & | & & \\ & & & & \mathrm{CH_2} & - & \mathrm{CH_2} - \mathrm{CH_3}\end{array}$$

SOLUTION

First we locate the longest chain. We have five possible chains, and the longest one has six carbons. Hence the name of the hydrocarbon would be hexane. Now we need to number the carbons so that we start numbering the closes to the substituents the possible.

$$
\begin{array}{llllll}
 & & CH_3 & & & \\
 & & | & & & \\
CH_3 - & CH_2 - & C - & CH_3 & & \\
1 & 2 & |3 & & & \\
 & & CH_2 - & CH_2 - & CH_3 & \\
 & & 4 & 5 & 6 &
\end{array}
$$

We have two methyl connected to carbon number three. Hence the final name will be:

▶Answer: 3-dimethylhexane.

STUDY CHECK

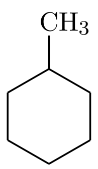
Name the following hydrocarbon:

$$
\begin{array}{lllll}
 & & CH_2 - CH_2 - CH_2 - CH_3 \\
 & & | \\
CH_3 - & CH_2 - & C - CH_3 \\
 & & | \\
 & & CH_2 - CH_2 - CH_3
\end{array}
$$

11.5 Cycloalkanes with substituents

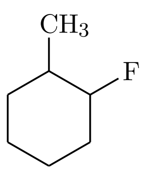
The naming rules for cycloalkanes are the same as the rules for naming linear alkanes. This means first you will find the ending of the name by counting the number of carbons in the cycle. Then you will locate each substituent and number the carbon chain so that the position numbers are small. Finally, all substituents should be ordered alphabetically.

Cycloalkanes with one substituent Let us take a look at the following cycloalkane:



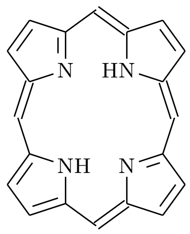
this is a cyclohexane connected to a methyl substituent. As there is only one substituent, there is no need to number the carbon chain. Hence the name would be simply methylcyclohexane.

Cycloalkanes with two substituents Let us take a look at the following cycloalkane:

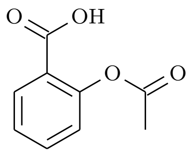


this is a cyclopentane connected to two different substituents: a methyl and a fluoro. In order to name this molecule we need to number the carbons first, and there are two different ways to number the cyclohexane ring:

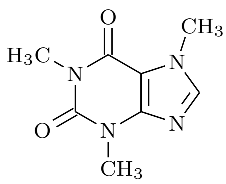
▼Porphyrin has several amine groups.

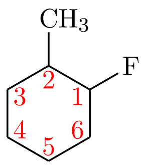


▼Aspirin contains an aromatic cycle, a carboxylic acid and a ester.

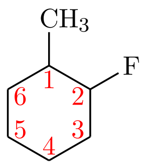


▼Caffeine contains amides and amines





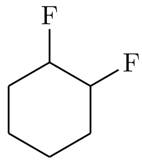
1-Fluoro-2-methylcyclohexane



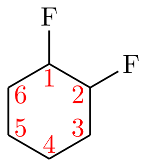
~~Fluoro-1-methylcyclohexane~~

we will choose the name that gives the lowest numbers: 1-Fluoro-2-methylcyclohexane.

`Cycloalkanes with repeated substituents` Let us take a look at the following cycloalkane, which has two repeated fluorine substituents:



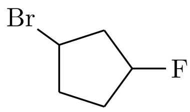
After numbering the chain:



The name would be: 1,2-difluorocyclohexane.

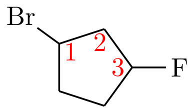
Sample Problem 117

Name the following hydrocarbon:



SOLUTION

The molecule is a cyclopentane with two substituents: fluoro and bromo. I will start numbering in bromo and continue until bromo. This way I will have small numbers and follow the abc rule:

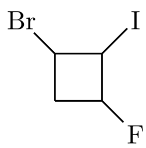


And the name of the molecule would be:

▶Answer: 1-bromo-3-fluorocyclopentane.

STUDY CHECK

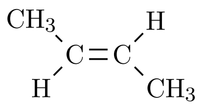
Name the following hydrocarbon:



11.6 Molecular diversity

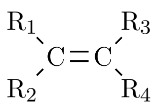
You have certainly taken painkillers for a headache or over-the-counter drugs to get over a cold. Maybe you drink coffee and perhaps you like tea. All these substances contain active organic molecules. These active molecules are hydrocarbon derivatives and differ from plain hydrocarbons, which are simply made of carbon and hydrogen. Active molecules contain functional groups such as alcohol, ethers, carboxylic acids, amines, amides, or aromatic groups. These groups of atoms have a specific function and give activity to the molecule. The goal of this section is simply to identify the different groups.

`Alkene group: double bonds` Alkenes contain at least one double bond between carbons. An example would be:



or $CH_3CH_2-HC=CH-CH_3$

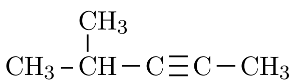
Double bonds are also called unsaturations. Hence, we say alkenes are unsaturated. As both sides of an alkene can be connected to different hydrocarbon chains we normally represent this as:



Alkene

where R and R' represent any hydrocarbon chain. For Example R and R' can be CH_3- or CH_3CH2-. Alkynes contain at least one triple bond between carbons. An example would be:

$CH_3-C\equiv C-C(CH_3)_3$ or

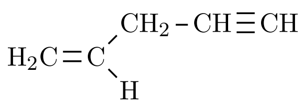


Again, we could use R and R' to represent any hydrocarbon chain:

$R-C\equiv C-R'$ Alkyne

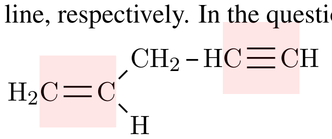
Sample Problem 118

Identify the alkene and alkyne groups in the molecule:



SOLUTION

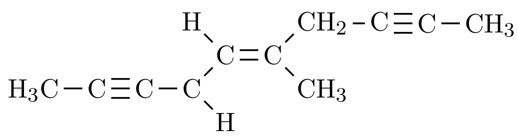
A double bond is carbon atoms sharing two pairs of electrons, whereas a triple bond is a pair of atoms sharing three pairs of electrons. They are represented with a double and triple line, respectively. In the question:



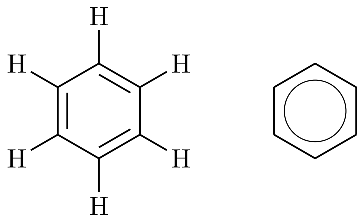
▶Answer: (left) alkene; (right) alkyne

STUDY CHECK

Identify the alkene and alkyne groups in the molecule:



Aromatic group Aromatic groups are based on bezene, a carcigenic cyclohexane-like molecule with a series of alternates double bonds, which is often represented as a circle for simplicity:



Benzene Benzene

Examples of molecules containing aromatic groups are:

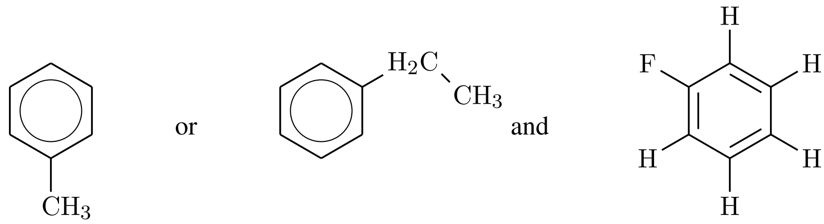
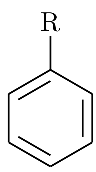


Table 11.4 Names of several functional groups

Functional group	Name	Functional group	Name	Functional group	Name
$R_1R_2C{=}CR_3R_4$	Alkene	$R-C(=O)-R'$	Ketone	$R-OH$	Alcohol
$R-C\equiv C-R'$	Alkyne	$R-C(=O)-H$	Aldehyde	$R-SH$	Thiol
$R-C(=O)-OH$	Carboxylic acid	$R-N(R')-R''$	Amine	$R-O-R'$	Ether
$R-C(=O)-O-R'$	Ester	$R-C(=O)-N(R'')-R'$	Amide		Phenyl

Alcohol, ether and thiol group Alcohols contain an $-$OH group attached to a carbon.

$R-OH$ Alcohol

Whereas ethers have oxygen atoms attached to two carbon atoms:

$R-O-R'$ Ether

Examples of alcohols and ether are:

$$CH_3-CH_2-OH \qquad CH_3-O-CH_2-CH_3$$

alcohol ether

Thiols contain a $-SH$ group attached to a carbon. They are equivalent to alcohols but based in sulfur:

$$R-SH \quad \text{thiol}$$

Examples of thiols are:

$$CH_3-CH_2-SH \qquad CH_3-\overset{\displaystyle SH}{\overset{|}{C}}H-CH_3$$

Thiols are responsible for the characteristic smell of rotten eggs and burned hair.

Sample Problem 119

Classify the following molecules as alcohol or ether.

$$CH_3-CH_2-CH_2-OH \qquad CH_3-O-CH_2-CH_3$$

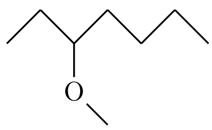
SOLUTION

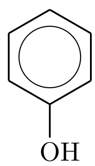
The OH groups is an alcohol, and we find this group in the left molecule. Differently, the right molecule is an ether as it contains the $R-O-R'$ group.

▶Answer: (left) alcohol; (right) ether.

STUDY CHECK

Classify the following molecules as alcohol or ether.





Aldehydes and ketones Ketones and aldehydes both contain a $C=O$ group. Still these are two different groups and ketones have a $C=O$ group bounded to two different carbon atoms

$$R-\overset{\displaystyle O}{\overset{||}{C}}-H \quad \text{Aldehyde}$$

whereas aldehydes have the same $C=O$ group but this time bounded to a carbon and a hydrogen.

$$R-\overset{\displaystyle O}{\overset{||}{C}}-R' \quad \text{Ketone}$$

Examples of aldehydes and ketones are:

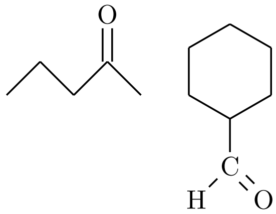
$$CH_3CH_2-\overset{\displaystyle O}{\overset{||}{C}}-H \qquad \text{or} \qquad CH_3-\overset{\displaystyle O}{\overset{||}{C}}-CH_3$$

Aldehyde Ether

Ketones are good solvents and acetone is commonly found in nail polish removes for example. Aldehydes have strong smells and for example, vanillin is an aldehyde responsible for the vanilla smell and cinnamaldehyde is responsible for the cinnamon odor.

Sample Problem 120

Classify the following molecules as an aldehyde or ketone



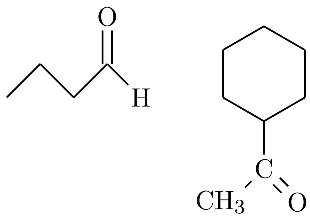
SOLUTION

The left molecule is a ketone as the carbonyl group (C=O) is connected to a CH_3 and a CH_2. Differently, the right molecule is an aldehyde as the carbonyl group is connected to a cycloalkane but also to a hydrogen.

▶Answer: (left) ketone; (right) aldehyde.

STUDY CHECK

Classify the following molecules as an aldehyde or ketone



Carboxylic acids and esters Carboxylic acids contain a carbonyl group (C=O) connected to an hydrocarbon and also an alcohol group:

$$\mathrm{R-\overset{\overset{\large O}{||}}{C}-OH} \qquad \text{Carboxylic acid}$$

Esters have the same C=O group but this time bounded to a carbon (R) and an ether group (–O–R').

$$\mathrm{R-\overset{\overset{\large O}{||}}{C}-O-R'} \qquad \text{Ester}$$

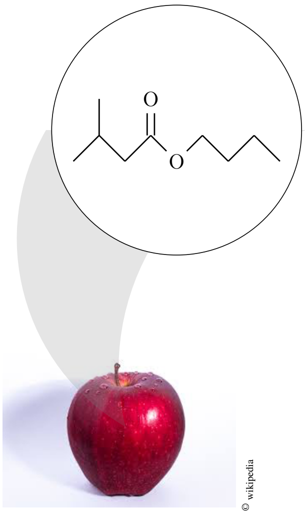
Examples of carboxylic acids and esters are:

$$\mathrm{CH_3-\overset{\overset{\large O}{||}}{C}-OH} \qquad \text{or} \qquad \mathrm{CH_3-\overset{\overset{\large O}{||}}{C}-O-CH_3}$$

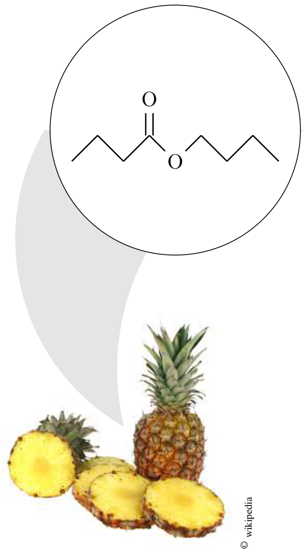
Carboxylic acid Esters

▼Acetone, the smallest ketone, is also known as nail polish remover.

▼Ethyl isovalerate is an ester that flavors apples.

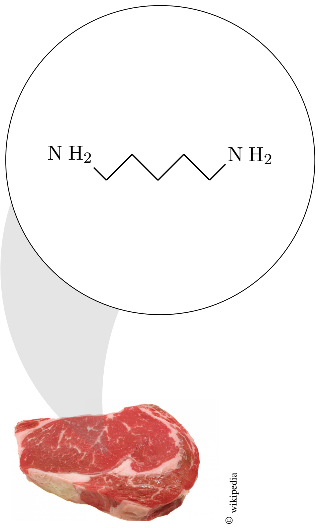


▼Butyl butyrate is an ester that flavors pineapple.

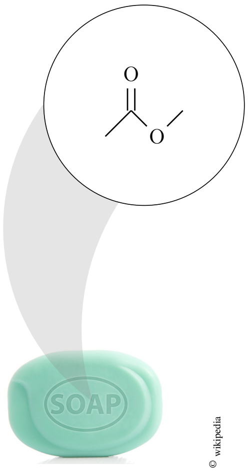


▼Diethyl ether was formerly used as a general anesthetic, until non-flammable drugs were developed.

▼Amines are responsible of the bad smell of decaying meat.



▼Esters are found in soaps



▼Vinegar contains acetic acid, the smallest carboxylic acid.

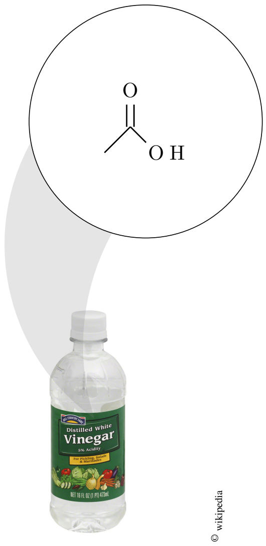
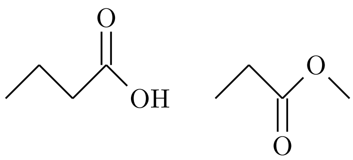


Figure 11.1 Some examples of functional groups with applications to real life.

Sample Problem 121

Classify the following molecules as an carboxylic acid or ester:

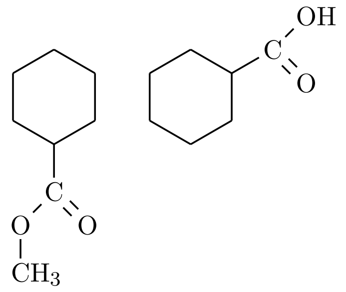


SOLUTION

The molecule in the left is a carboxylic acid as the carbonyl group (C=O) is connected to an alcohol –OH. Differently, the molecule in the right is a ester as the carbonyl group is connected to a –O–CH_3 group.

STUDY CHECK

Classify the following molecules as an carboxylic acid or ester:



`Amines and amides` Amines and amides are groups containing nitrogen. Amines are derivative of ammonia (NH_3) with one or more of the hydrogen atoms being replaced by a hydrocarbon

$$\begin{array}{c} R' \\ | \\ R-N-R'' \end{array} \qquad \text{Amine}$$

All these molecules are amines:

$$\begin{array}{c} CH_3 \\ | \\ CH_3-N-H \end{array} \qquad CH_3-NH_2 \qquad \begin{array}{c} CH_3 \\ | \\ CH_3-N-CH_2CH_3 \end{array}$$

Amines have very strong and unpleasant odors often associated with the smell of decaying flesh:

$$NH_2 - \left(CH_2 \right)_3 - NH_2 \qquad\qquad NH_2 - \left(CH_2 \right)_5 - NH_2$$

putrescine cadaverine

Amides, on the other hand, contain a carbonyl group (C=O) connected to an amine group ($-\overset{|}{N}-$)

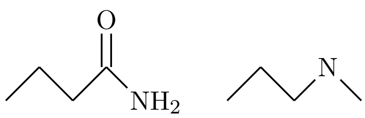
$$\begin{array}{c} OR'' \\ \|| \\ R-C-N-R' \end{array} \qquad \text{Amide}$$

Examples of amides are:

$$\begin{array}{c} O \\ \| \\ CH_3-C-NH_2 \end{array} \qquad \text{or} \qquad \begin{array}{c} OH \\ \|| \\ CH_3-CH_2-C-N-CH_3 \end{array}$$

Sample Problem 122

Classify the following molecules as an amide or amine:



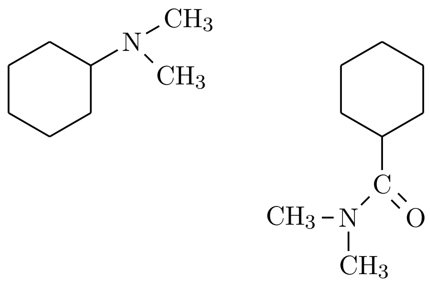
SOLUTION

The molecule in the left is a amide as the carbonyl group (C=O) is connected to a nitrogen atom. Differently, the molecule in the right is an amine as the nitrogen group is not connected to any carbonyl group.

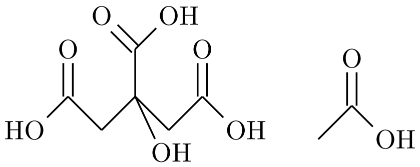
▶Answer: (left) amide; (right) amine.

STUDY CHECK

Classify the following molecules as an amide or amine:

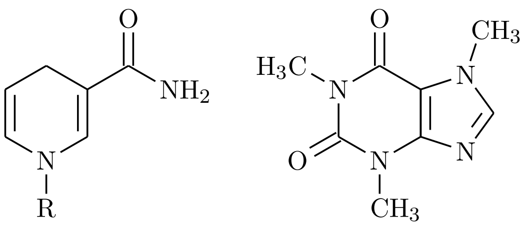


Identifying organic acids and bases Acid organic molecules, in general, contain one or more carboxylic acid groups. An example would be vinegar or citric acid, the acid found in citrus.



Citric acid Acetic Acid

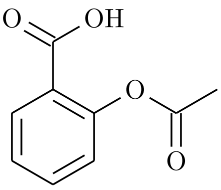
Bases contain one or more amine groups. Examples of bases are NADH, a coenzyme found in all living cells, or caffeine, a base found in coffee.



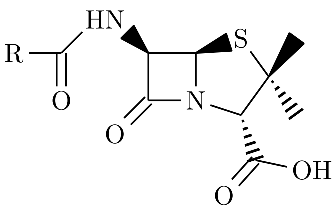
NADH Caffeine

Sample Problem 123

Classify the following molecules as an acid or base:



Aspirin



Penicillin

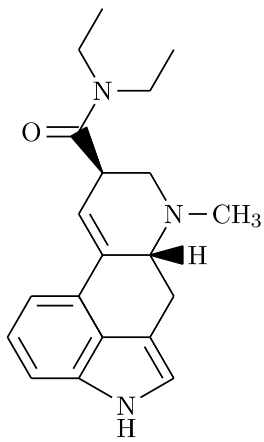
SOLUTION

Aspirin contains an acid group and hence it will be acidic. Penicillin contains both acidic and amine groups and hence will be both acidic and basic.

▶Answer: (left) acid; (right) both.

STUDY CHECK

Classify the following molecules as an carboxylic acid or base:



LDS

CHAPTER 11

Alkanes

11.1 Classify the following chemicals in two different categories. Give a rationale for your classification. For the group with more chemicals, further classify those chemicals in two categories. (a) KCl (b) C_2H_2 (c) C_4H_{10} (d) $FeO_{(s)}$ (e) C_6H_{12} (f) PH_3 (g) H_2O

11.2 Working in groups, select an everyday-life object (e.g a spoon) and guess whether the materials that made this object are mostly organic or inorganic (e.g mostly inorganic). Without revealing your answer, present the material you selected to another team member and ask him or her to give you his or her point or view regarding whether the materials that made this object are mostly organic or inorganic.

11.3 Indicate the molecular formula of the following organic compounds: (a) ethane (b) butane

11.4 Indicate the molecular formula of the following organic compounds: (a) methane (b) decane

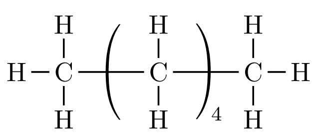
11.5 Name the following alkanes: (a) C_3H_8 (b) C_8H_{18}

11.6 Name the following alkanes: (a) C_5H_{12} (b) C_9H_{20}

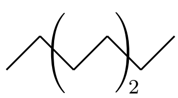
11.7 Write down the condensed formula of the following alkanes: (a) hexane (b) propane

11.8 Write down the expanded formula of the following alkanes: (a) pentane (b) decane

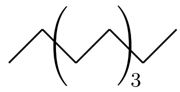
11.9 Write down the molecular formula for the chemical below:



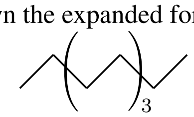
11.10 Write down the molecular formula for the chemical below:



11.11 Write down the molecular formula for the chemical below:



11.12 Write down the expanded formula for hexane.

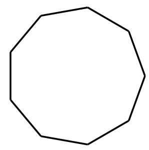


Cycloalkanes

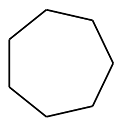
11.13 Indicate the molecular formula of the following cycloalkanes: (a) cyclobutane (b) cyclopentane

11.14 Indicate the molecular formula of the following cycloalkanes: (a) cyclopropane (b) cyclohexane

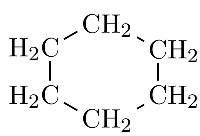
11.15 Name the following cycloalkane:



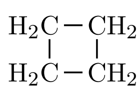
11.16 Name the following cycloalkane:



11.17 Name the following compound:

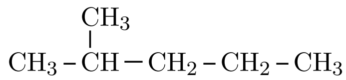


11.18 Name the following compound:

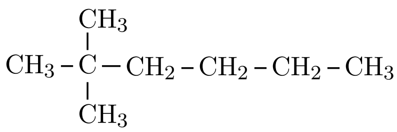


Alkanes with substituents

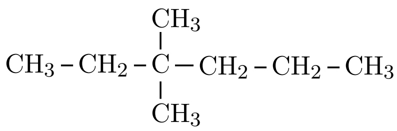
11.19 Name the following compound:



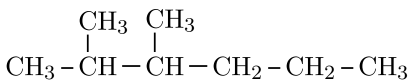
11.20 Name the following compound:



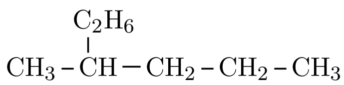
11.21 Name the following compound:



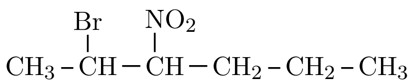
11.22 Give the name for the following compound:



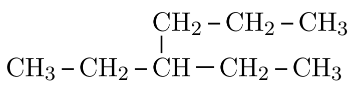
11.23 Give the name for the following compound:



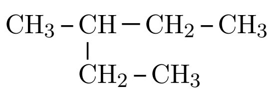
11.24 Give the name for the following compound:



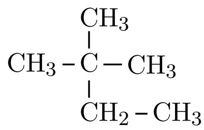
11.25 Give the name for the following compound:



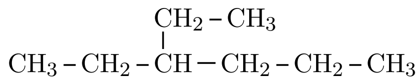
11.26 Give the name for the following compound:



11.27 Give the name for the following compound:

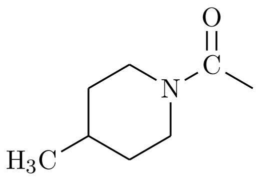


11.28 Give the name for the following compound:

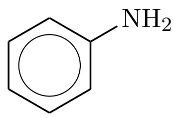


Functional groups

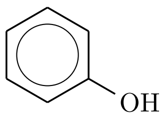
11.29 Identify the functional groups in the following molecule:



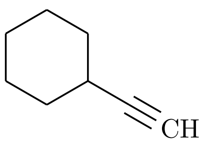
11.30 Identify the functional groups in the following molecule:



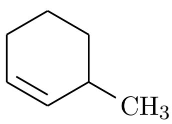
11.31 Identify the functional groups in the following molecule:



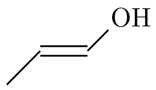
11.32 Identify the functional groups in the following molecule:



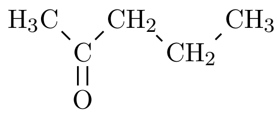
11.33 Identify the functional groups in the following molecule:



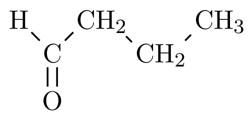
11.34 Identify the functional groups in the following molecule:



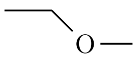
11.35 Identify the functional groups in the following molecule:



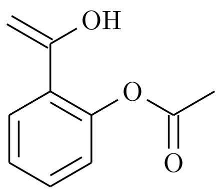
11.36 Identify the functional groups in the following molecule:



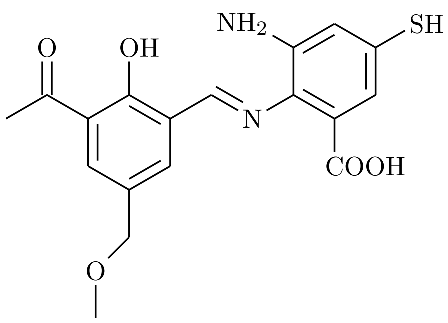
11.37 Identify the functional groups in the following molecule:



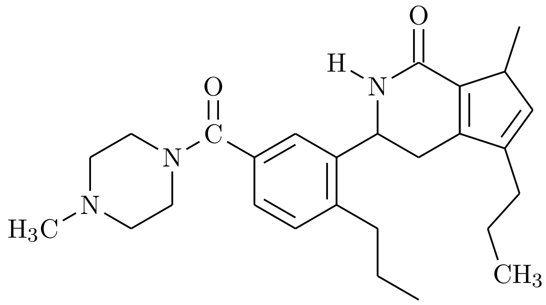
11.38 Identify the functional groups in the following molecule:



11.39 Identify the functional groups in the following molecule:



11.40 Identify the functional groups in the following molecule:

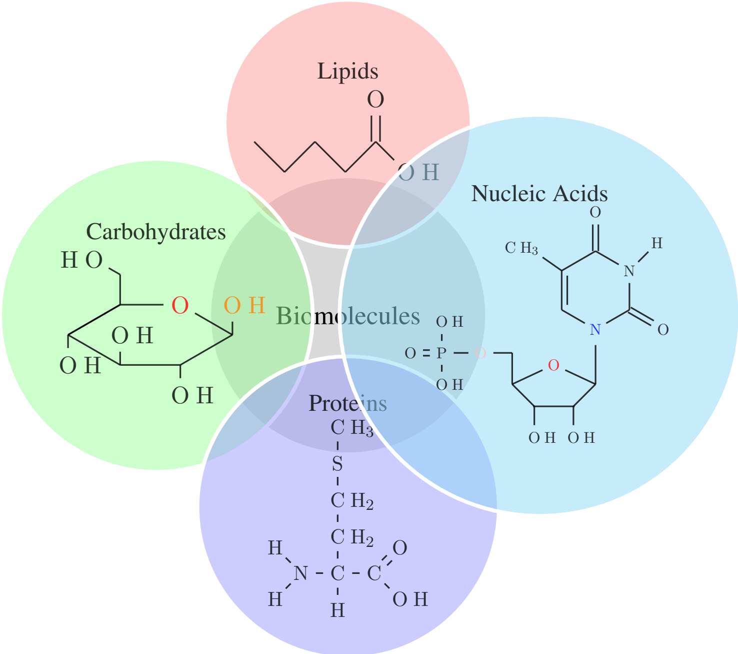


Ch. 12. Biochemistry

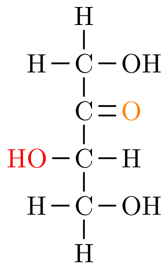
BIOCHEMISTRY applied the principles of chemistry to organic molecules found in living systems. In particular it covers the properties of biomolecules: carbohydrates, lipids, proteins and nucleic acids, perhaps less known biomolecules. You can find carbohydrates in pasta, bread or rice and they give your body quick energy. These molecules are in essence biological alcohols, which contain numerous $-OH$ groups. Lipids on the other hand are often referred as fats or oils. Cholesterol is for example a lipid found in eggs or milk. The purpose of lipids is to store energy and insulate organs. Proteins have numerous functions in the body such as building muscle or transport oxygen in the blood. Perhaps more importantly is to know what are they made off: amino acids. Finally, if you are not familiar with the term nucleic acids, perhaps you recognize the term DNA. Nucleic acids are large molecules that contains the information for cellular growth and reproduction. Over all, this chapter will briefly cover the structure and composition of these important molecules of life.

12.1 Biomolecules

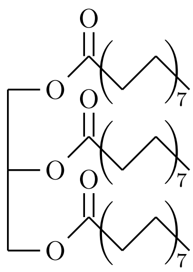
Biomolecules, also called biological molecule, are any of numerous substances that are produced by cells and living organisms. Biomolecules have a wide range of sizes and structures and perform a vast array of functions.



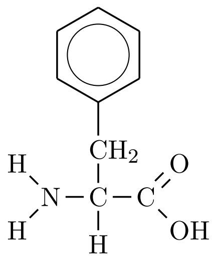
`Different types of biomolecules` The four major types of biomolecules are carbohydrates, lipids, nucleic acids, and proteins. Carbohydrates (sugars or simply carbs) are organic molecules composed of numerous alcohol groups as well as a ketone or ether group. An example of a carbohydrate is presented below. We can find simple carbohydrates such as glucose and complex carbohydrates such as starch.



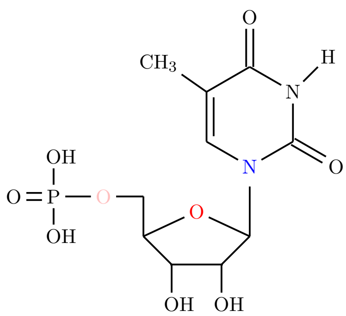
Lipids (fats) tend to contain carboxylic acids or esters. We can find different types of lipids such as fatty acids, waxes or steroids. An example of a lipid is presented below.



Proteins (meat) contain amino or amide groups. Proteins are made of amino acids. We can find complex proteins or simple proteins called peptides. An example of a protein in general is presented below.



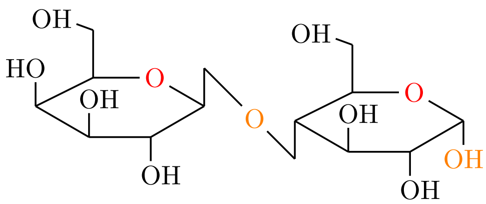
Finally, nucleic acids (DNA or RNA) are made of a sugar connected to an amine, or amide. Sometimes you will find phosphorus atoms on its structure. An example of nucleic acid is presented below.



It is very important to be able to differentiate the different biomolecules as well as to be able to identify each of them.

Sample Problem 124

Classify the following biomolecule as a carbohydrate, a lipid, a protein or a nucleic acid.

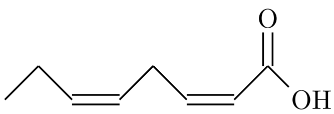


SOLUTION

The biomolecule presented above contains numerous ether groups as well as alcohols. This is a carbohydrate. In particular, it is a complex carbohydrate resulting of the combination of two simple carbohydrates.

STUDY CHECK

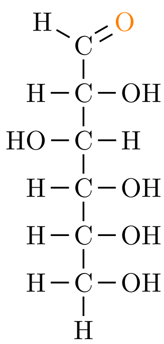
Classify the following biomolecule as a carbohydrate, a lipid, a protein or a nucleic acid.



12.2 Simple Carbohydrates

This first section introduces one of the most important biomolecules: the carbohydrates. If you think about rice, bread, and even fruits, they all contain carbohydrates as their main ingredients. This section cover the rules for their naming of single carbohydrates, called monosaccharides, and will show you how to identify these common molecules in two different forms: a line structure and a cycle, the later called Haworth structure, named after Sir Norman Haworth, a famous english biochemist. Finally carbohydrates can form dimer and polymers and these more complex molecules will also be addressed in this section.

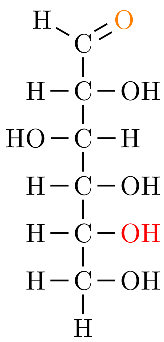
`Carbs are made of C, H and O` Carbohydrates are made of carbon and oxygen and hydrogen, thus their name: carbon hydrates. The simples carbohydrate is glucose with molecular formula $C_6H_{12}O_6$ and with structure



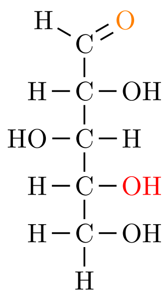
This molecule is produced in plants by means of a process called photosynthesis, in which carbon dioxide and water are added together with the help of sunlight to produce glucose and oxygen. At the same time our bodies burn glucose with oxygen while we breath.

$$6\,CO_2 + 6\,H_2O + \text{☼} \;\underset{\text{Respiration}}{\overset{\text{Photosynthesis}}{\rightleftharpoons}}\; C_6H_{12}O_6 \;+ 6\,O_2$$

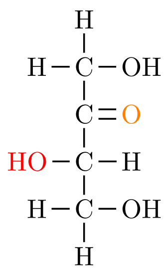
`Monosaccharides: aldo or keto?` The simples sugars are called monosaccharides and their name ends in *ose* like in gluc*ose*. These sugars are made of C, O and H and contain a few functional groups that you should be familiar with. They contain alcohol groups and either and aldehyde ($R-\overset{\overset{\displaystyle O}{||}}{C}-H$) or a ketone ($R-\overset{\overset{\displaystyle O}{||}}{C}-R'$) group in the carbon number one. Depending on the number of carbons, monosaccharides can be classifies in triose, tetrose, pentose, or hexose. Here a few examples:



An aldohexose



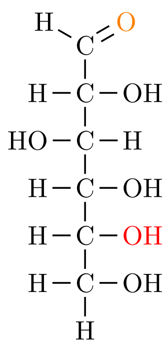
An aldopentose



A ketotetrose

Sample Problem 125

Classify the following monosaccharide as aldo or keto:

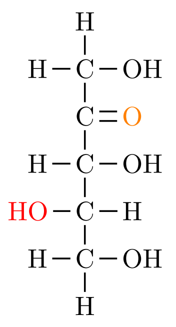


SOLUTION

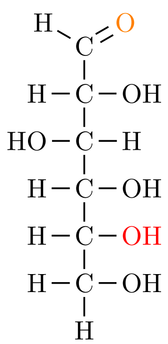
The monosaccharide has a $R-\overset{\overset{\displaystyle O}{||}}{C}-R'$ group on carbon one and therefore it will be aldo. As it also has six carbons, it will be an aldohexose.

STUDY CHECK

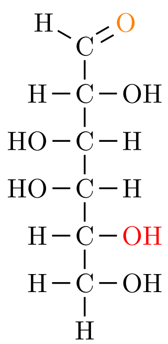
Classify the following monosaccharide as aldo or keto:



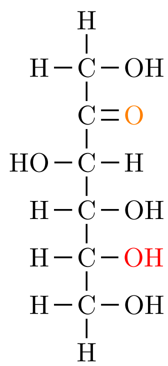
Some important monosaccharides The most important monosaccharides are glucose, fructose and galactose. These are all hexose with six carbon atoms. Glucose and galactose are aldohexose whereas fructose is a ketohexose.



Glucose



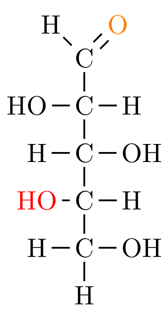
Galactose



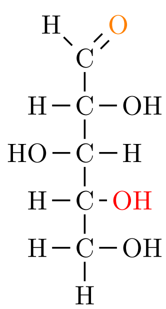
Fructose

Glucose is also known as dextrose or blood sugar, is the most important simple sugar involved in the metabolism of humans. Glucose is found in corn syrup and honey as well as in fruits and vegetables in the form of more complex carbohydrates. Galactose can not be found in free form, but is found in the form of lactose, a disaccharide, in milk. Only when lactose is broken down you can find free galactose. Fructose is the sweetest of all carbs found in fruits–thus its name–and honey.

$d-$ and $l-$ monosaccharides Monosaccharides are classified as $d-$ and $l-$ depending on the position of the second OH group starring from the bottom of the molecule. If that specific OH points to the right we call it $d-$ monosaccharides, whereas $l-$ have the second OH group starring from the bottom pointing left. Here some examples:



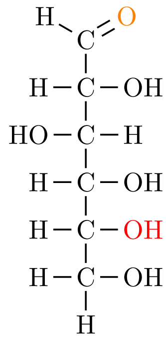
$l-$ monosaccharide



$d-$ monosaccharide

Sample Problem 126

Classify the following monosaccharide as $d-$ or $l-$:

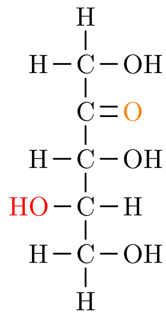


SOLUTION

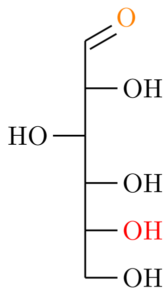
As the second OH group starting from the bottom point to the right, it will be $d-$, in particular it would be a: $d-$aldohexose.

STUDY CHECK

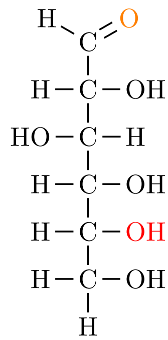
Classify the following monosaccharide as $d-$ or $l-$:



`Skeletal structure of monosaccharides` Often times, the structure of carbohydrates are represented in skeletal form, ignoring the C and H atoms and showing only oxygen and OH groups. Here an example:

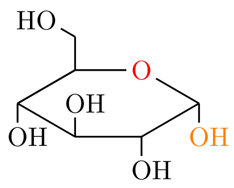


Linear Glucose (Skeletal form)

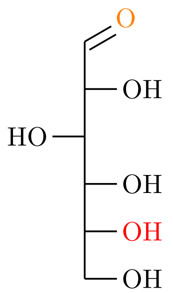


Linear Glucose (Full form)

`Haworth cyclic structure of monosaccharides` Normally monosaccharides exist in liquid solution. In these conditions its structure is not a linear structure such the one presented before. Instead they exist in a cyclic form called Haworth structure. The goal of this section is just to introduce you to this cyclic form. Here an example:

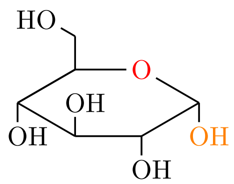


Haworth Glucose

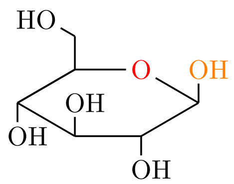


Linear Glucose
(Skeletal form)

α or β form of Cyclic monosaccharides The Haworth form of monosaccharides are classified as α or β depending on the orientation of a single OH group: the OH group of the C atoms directly connected to the O atom in the cycle. If this OH groups points down we call this α and if its points up we call the structure as β . Here some examples:



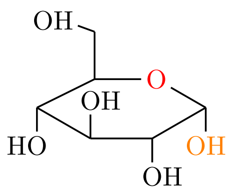
α-Glucose



β-Glucose

Sample Problem 127

Classify the following haworth monosaccharide as α or β :

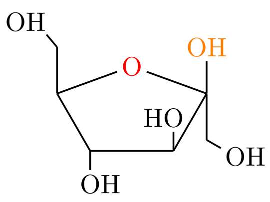


SOLUTION

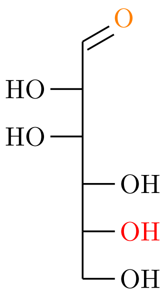
As the second OH group after the -O- in the cycle points down, it will be α.

STUDY CHECK

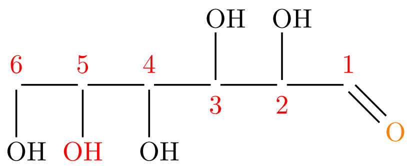
Classify the following haworth monosaccharide as α or β :



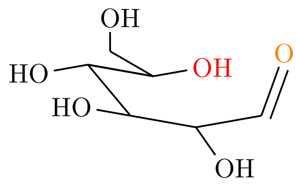
Obtaining Haworth cyclic structure of monosaccharides Linear monosaccharides cycle forming Haworth cycles. In order to obtain a cycle you just need to get the linear structure and turn in clockwise 90 degree until is horizontal, the carbonyl or keto will fall in the right part of the structure and the last C atom in the left part. After that you need to draw the structure in a boat shape, as the Oh from carbon 5 will attach carbon number one and produce a cycle.



D-Mannose

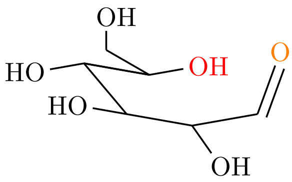


D-Mannose rotated

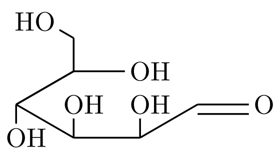


D-Mannose as a boat

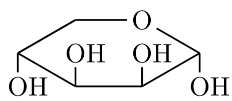
The final cycle looks like:



D-Mannose

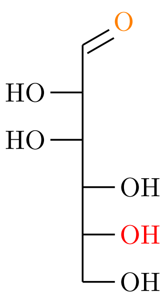


D-Mannose rotated

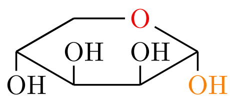


α-Mannose

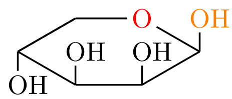
In general the cycles formed from linear monosaccharides can have five of six atoms, depending on the aldo/keto nature of the monosaccharides. Aldo monosaccharides produce six-atoms rings and keto monosaccharides produce five-atoms rings. A six-atom ring example:



D-Mannose

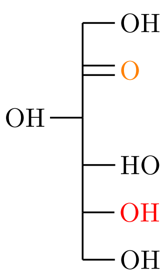


α-d-Mannose

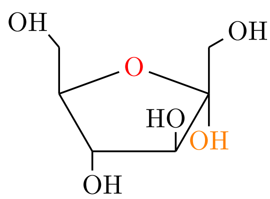


β-d-Mannose

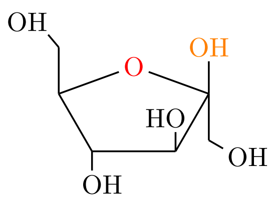
A five-atom ring example:



d-fructose



α-d-fructose

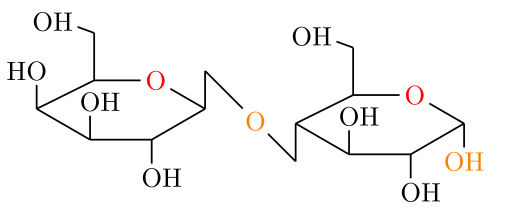


β-d-fructose

12.3 Disaccharides and polysaccharides

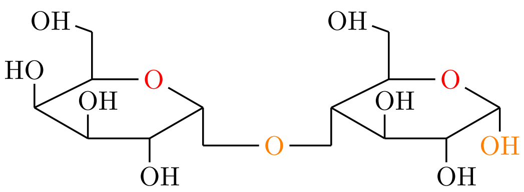
This next section covers the properties of the di and polysaccharides. These are more complex saccharides made of two of more units of sugar. At the same time, these form the most common carbs. For example, sugar cane is a disaccharides, that is, contains two monosaccharides units. Differently, the starch in potatoes is a polysaccharide, made of numerous saccharides units, and cellulose is a structural polysaccharide used by plants. All this complex saccharides have very different properties and for example humans can easily digest sugar cane and starch but not cellulose. The difference is in the way the monosaccharides are connected among themselves by means of a connection names glycosidic bond.

Disaccharides Disaccharides are the combination of two saccharides. An example of a disaccharides is α-lactose, found in the milk:



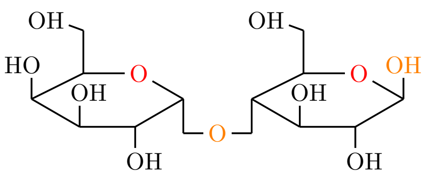
α-lactose

Another example is α-maltose, found in beer:

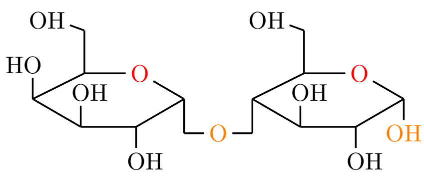


α-maltose

α- and β-Disaccharides Same as monosaccharaides, disaccharides can also be α- and β, depending on the position of the first OH after the -O- atom in the right cycle. For example:



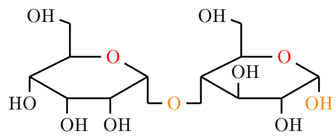
β-maltose



α-maltose

Sample Problem 128

Classify the following disaccharide as α or β:

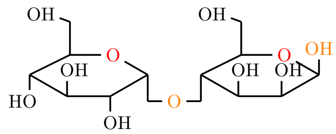


SOLUTION

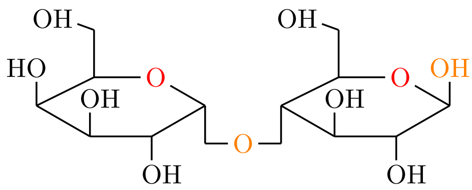
This is an α disaccharide as the first OH after the -O- atom in the right cycle is pointing down.

STUDY CHECK

Classify the following disaccharide as α or β:

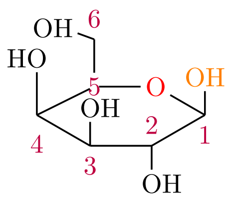


`Numbering saccharides` Disaccharaides are the result of the connection between two monosaccharides at a specific location of the molecule. Depending on the location of the connection and the orientation of the connection, the disaccharide would be easy or impossible to digest. Is is important to number the carbons in each monosaccharide in order to be able to identify the atoms involved in the connection. Here is an example of a monosaccharide:

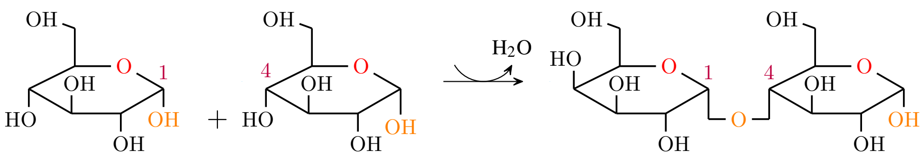


β-maltose

In order to number the chain we start by the -O- atom. This atoms would be zero. After that you continue numbering clockwise. The results is:



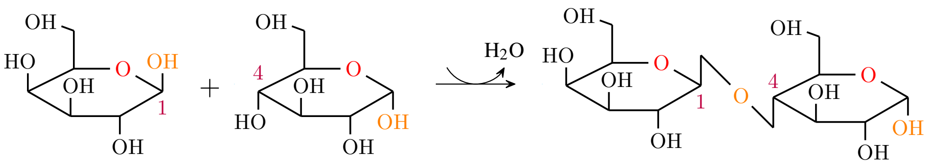
`Glycosidic bond` Monosaccharides connect to form disaccharides by means of very special bond. This bond is the result of combining two OH groups and looks like -O-. This is the glycosidic bond and the formation of this bond produced water. For example, maltose is the result of the connection of two glucose molecules by means of a 1-4-glycosidic bond.



α-glucose and α-glucose

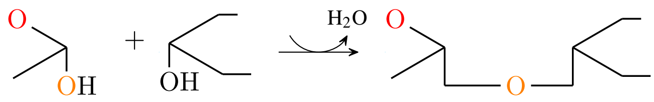
Maltose (α-1-4-Glycosidic bond)

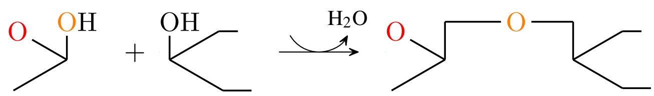
So if you combine an α-glucose molecule with another α-glucose molecule you form a 1-4-glycosidic bond. As the first glucose is α the resulting bond will also be an α bond, and the full name for the bond will be α-1-4-glycosidic bond. Now, depending on the orientation of the OH groups involved in the glycosidic bond we can have four different orientations of the bond. For example, lactose results from combining β-D-galactose and α-D-glucose. The separate monomers are:

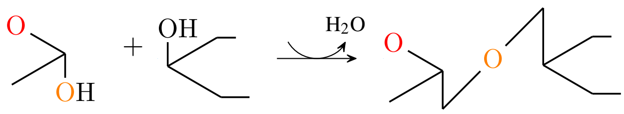


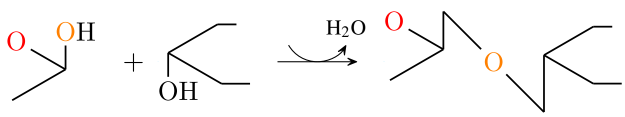
β-galactose and α-glucose   Lactose (β-1-4-Glycosidic bond)

In general the four possible glycosidic bonds are:


α-1,4-Glycosidic bond

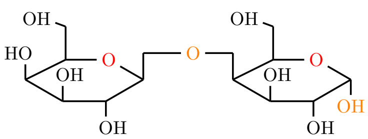

β-1,4-Glycosidic bond


α-1,4-Glycosidic bond


β-1,4-Glycosidic bond

Sample Problem 129

Classify the following glycosidic bond:

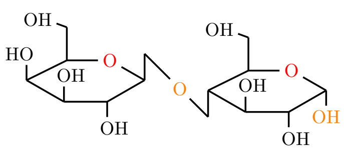


SOLUTION

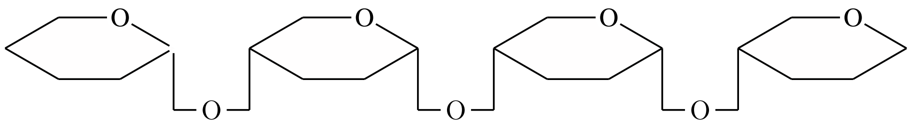
This is a β-1-4-glycosidic bond, as the first bond in the glycosidic bond points up.

STUDY CHECK

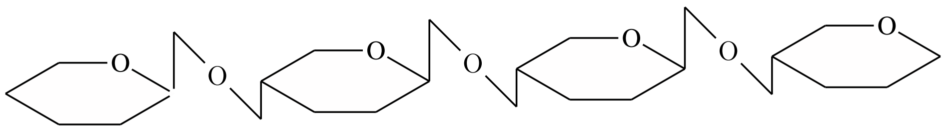
Classify the following glycosidic bond:



`Polysaccharides: starch and cellulose` In the previous section you saw the structure of maltose and lactose and these were made of monosaccharides connected by means of a -O- bond called glycosydic bond. A polysaccharide is a long polymer formed by means of numerous units of sugar connected by means of glycosidic bonds. Starch for example is a polysaccharide found in plants, rice, wheat or potatoes is made of α-glucose molecules connected by means of a α-1-4Glycosidic bond. These bonds are easy to break and humans easily digest starch. Starch hydrolyze in water and acid ultimately producing glucose.

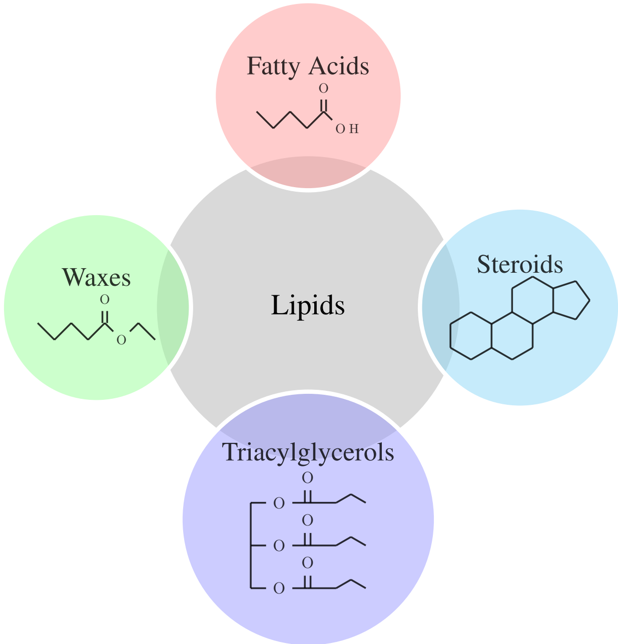


Cellulose is also a polymer of α-glucose molecules. However, this polysaccharide contains glucose molecules bonded by means of a β-1-4Glycosidic bond. These bonds are stronger than the α-1-4Glycosidic bond, and hence humans can not digest cellulose. Only animals can break this polymer with the help of a specific bacteria.

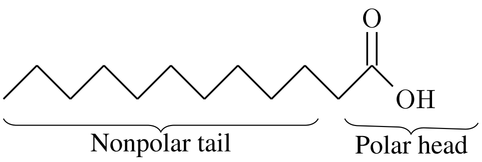


12.4 Lipids

This section covers the chemical properties and structure of lipids. These biomolecules are often referred as fats as they are insoluble in water and soluble in organic solvents. Think about oil, butter or even cholesterol. All these are different types of lipids with distinct structure. One one hand, oil and butter contain fatty acids, which are organic acids–carboxylic acids–with a carboxylic group connected to a very long zigzag hydrocarbon chain. On the other hand, cholesterol is a lipid that do not contain fatty acids. The structure of cholesterol results from an steroid nucleus made of three cyclohexane and a cyclopentane fused together. Perhaps less known fats are waxes and triacylglicerols. Both result from the reaction of fatty acids and alcohols. We will briefly cover the structure of those lipids as well. The following diagrams gives the classification for the different kinds of lipids.



Structure of fatty acids Fatty acids are molecules with two different parts: a polar head and a nonpolar tail. On one hand, a carboxylic acid group $(\overset{O}{\overset{\|}{C}}-OH)$ make these molecules acidic and polar. On the other hand, a hydrocarbon () chain makes these molecules nonpolar, that means and hence insoluble in water. As the chain tends to be long these molecules over all are nonpolar that means they do not dissolve in water. These fatty acids are called saturated fatty acids, as they only contain simple C-C bonds.



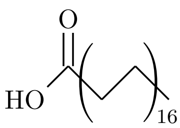
Saturated Fatty acid

Sample Problem 130

Geddic acid is a saturated fatty acid with 34 carbons found in animal fat. Write down its skeletal formula.

SOLUTION

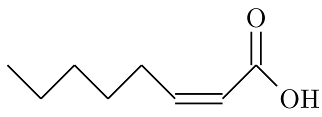
Saturated fatty acids have a polar acid head and a hydrocarbon chain with no double bonds. The polar head has one carbon and hence the tail should have 33 carbons:



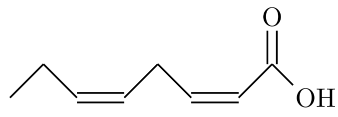
STUDY CHECK

Lignoceric acid is a saturated fatty acid with 24 carbons found in wood tar and peanut oil contain. Write down its skeletal formula.

Mono and polyunsaturated fatty acids Often times fatty acids contains double bonds called unsaturations. They can have a single double bond or more than one double bond. Monounsaturated fatty acids only contain a single unsaturation and polyunsaturated fatty acids have more than one double bond. Examples are palmitoleic acid, a monounsaturated fatty acid found in butter or linolenic acid, a polyunsaturated fatty acid found in corn oil.



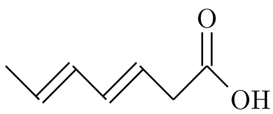
Monounsaturated fatty Acid



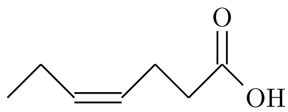
Polyunsaturated fatty Acid

Sample Problem 131

Classify the following fatty acids as saturated, monounsaturated or polyunsaturated:



A



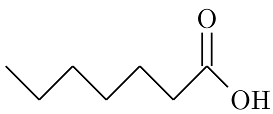
B

SOLUTION

A saturated fatty acid has no double bond on its tail, whereas a monounsaturated fatty acid has a single double bond in the tail and a polyunsaturated fatty acid has more than two double bonds. According to this, the fatty acid A is a polyunsaturated as it contains two double bonds on its tail. The fatty acid A is monounsaturated, as it only contains a single double bond.

STUDY CHECK

Classify the following fatty acid as saturated, monounsaturated or polyunsaturated:



Melting point of fats Some fats such as olive oil are liquid at room temperature. Other such as butter or lard are solid at room temperature. Liquid fats have low melting point so that they are already melted in a liquid form at room temperature. Solid fats have in general larger melting point and hence at room temperature they preserve its solid form. Two different factors affect the melting point of a fat. The first factor is the number of carbons. The larger the number of carbons the higher melting point and hence the fat will most likely be solid at room temperature. For example, lauric acic, found in coconut oil, has 12 carbon atoms and a boiling point of 44°C. Palmitic acid, an oil found in palm, on the other hand has 16 carbon atoms and its boiling point would be therefore higher (63°C).

$$CH_3 \left(CH_2 \right)_{12} COOH$$

Lauric Acid (MP=44°C)

$$CH_3 \left(CH_2 \right)_{14} COOH$$

Palmitic Acid (MP=63°C)

The second factor is the presence of double bonds–remember these are called unsaturations. Double bonds decrease the melting point. For example, stearic acid is an unsaturated fatty acid–with no double bonds–with 18 carbons and a melting point of 69°C. On the other hand, linoleic acid is a monounsaturated fatty acid with the same 18 carbons but a lower melting point of -5°C. Due to the presence of a single unsaturation the melting point is highly reduced and hence linoleic acid will be liquid at room temperature whereas stearic acid will be a solid.

$$\mathrm{CH_3} \left(\mathrm{CH_2} \right)_4 \mathrm{CH{=}CH{-}CH_2{-}CH{=}CH} \left(\mathrm{CH_2} \right)_7 \mathrm{COOH}$$

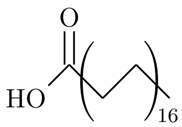
Linoleic Acid (MP=−5°C)

$$\mathrm{CH_3} \left(\mathrm{CH_2} \right)_{16} \mathrm{COOH}$$

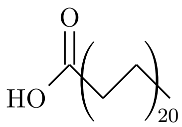
Stearic Acid (MP=69°C)

Sample Problem 132

Compare the melting point of the following fatty acids and predict whether they will likely be solid or liquid at room temperature.



A



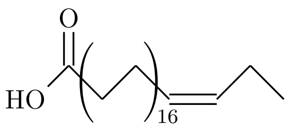
B

SOLUTION

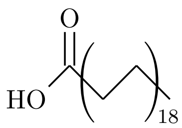
The larger the number of atoms in the fatty acid the higher its melting point and hence the higher the chances the fat is solid at RT. The fatty acid A has 33 carbon atoms whereas B has 41 so B has higher melting point than A and has more chances to be solid at RT.

STUDY CHECK

Compare the melting point of the following fatty acids and predict whether they will likely be solid or liquid at room temperature.



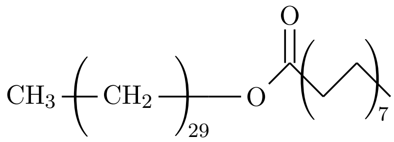
A



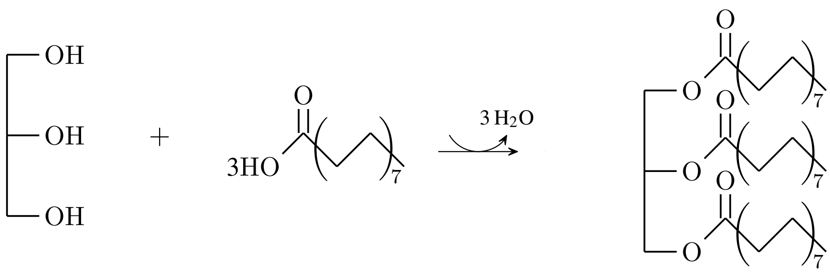
B

Carnauba wax is obtained from the leaves of the carnauba palm.

Carnauba wax

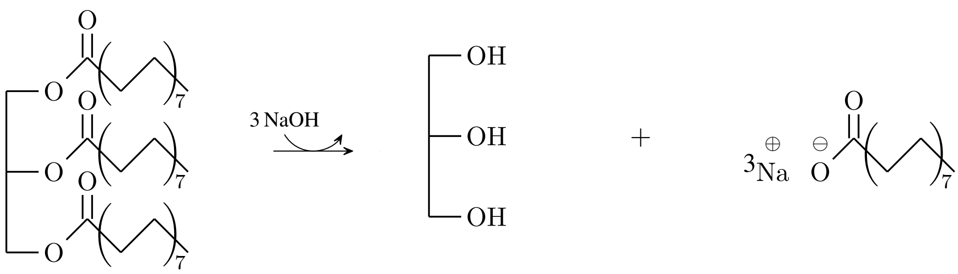


`Waxes, animal fat and soaps` Waxes and animal fat are the results of a reaction between fatty acids and alcohols. Waxes are ester of a fatty acid produced by mixing a fatty acid with long chain alcohols. Waxes are found in animals and plants. Plants secrete waxes in order to control evaporation, reduce wettability and increase hydration. The most important commercial plant wax is carnauba wax, a hard wax obtained from Brazilian palm employed for example in car and furniture polish, floss coating, and surfboard wax. Animal fat is also known as triacylglycerols as they result of the reaction between fatty acids and glycerol, an alcohol with three carbon atoms. Animals store energy in the form of fat. For example, during hibernation they survive by consuming fat.



Glycerol Fatty acids triacylglycerol

Soap results from the treatment of fat with a strong base such as NaOH or KOH. Due to this basic treatment fats break down into glycerol and a salt of the fatty acid. Soaps are indeed salts of fatty acids. Depending on the nature of the base employed in the saponification and the nature of the oil one can produce solid or more more liquid soaps.

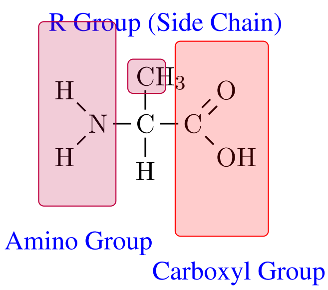


triacylglycerol glycerol Soap

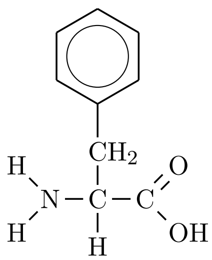
12.5 Proteins

This new section covers the chemical properties of proteins, which are polymers amino acids. Amino acids are molecules with both an amino (NH_2) and a carboxylic acid group (COOH). Amino acids have also a side chain (R)–this is a organic group different than amino or acid group. Side chains give amino acids acid-base hydrophilic-hydrophobic properties and properties. Amino acids bond together by means of a peptide amide bond producing small peptides or large proteins.

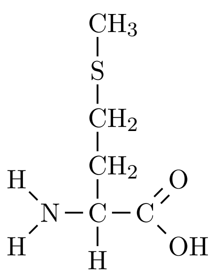
`Amino acids` Proteins are made of amino acids. Amino acids have the following parts: (1) a central carbon called α carbon; (2) an amino (NH_2) group connected to the α carbon; (3) a carboxylic acid (COOH) group connected to the α carbon; (4) a hydrogen atom (H) connected to the α carbon; (5) a side chain (R). All amino acids have the same α carbon, amino group and acid group as well as hydrogen connected to the α carbon. They differ in the side chain. Here is an example of an amino acid called Alanine, indicating the amino, the acid and the side chain.



Two more amino acids examples:



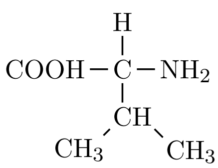
Phenylalanine



Methionine

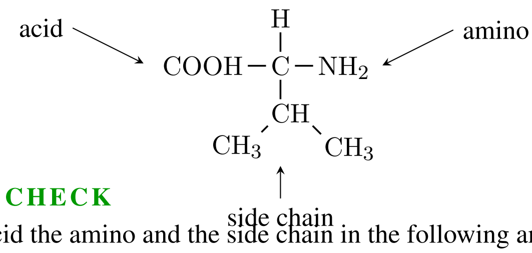
Sample Problem 133

Identify the acid the amino and the side chain in the following amino acid:



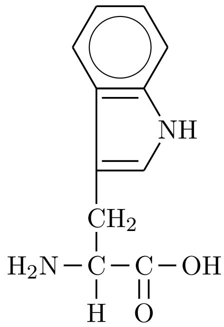
SOLUTION

The acid group is COOH and the amino is NH_2. The side chain is the other group connected to the central C different than hydrogen:

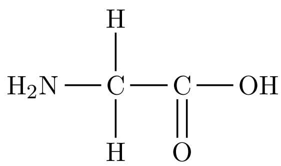


💎 **STUDY CHECK**

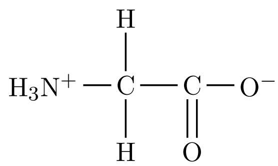
Identify the acid the amino and the side chain in the following amino acid:



`Amino acids as zwitterions` Amino acids are acid and basic, that means they have two different PK_a. For example the two different PK_a values for Glycine is 2.34 and 9.60. This means that in solution amino acids will exist in a protonated form in which the acid group looses its proton H^+ becoming negatively charged and the amino group has an extra proton and becomes positive. This positive and negative state of an amino acid is called zwitterion.



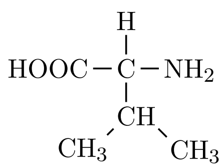
Neutral Glycine
(Gly)



Glycine as zwitterion

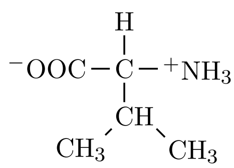
Sample Problem 134

Write down the zwitterion state of the following AA



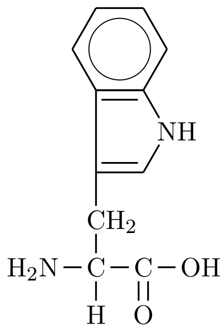
SOLUTION

In its zwitterion state an AA has a netaive acidic group and a positive amino group. The acid group has one less H and the amine gains that proton:

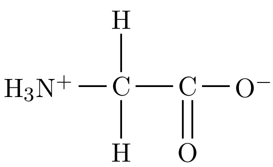


💎 **STUDY CHECK**

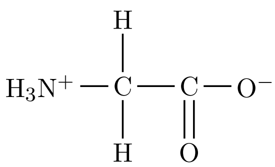
Write down the zwitterion state of the following AA



Hydrophilic and hydrophobic Amino acids There 20 common amino acids. They can be classified according to they hydrophilic and hydrophobic. Hydrophilic means that they solve in water whereas hydrophobic means that they do not solve in water. The classification is result of their side chain, as some of their side chains are hydrophilic whereas others are hydrophobic. For example, Glycine is an hydrophobic amino acid as its side chain is mainly made of carbon and hydrogen and is hence nonpolar. Differently, Serine is an hydrophilic amino acid as its side chain contains oxygen and is hence polar.

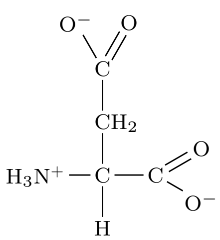


Glycine (Gly)
Hydrophobic AA

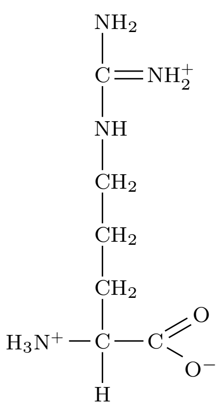


Histidine (His)
Hydrofilic AA

Acid and basic Amino acids Amino acids can also be classified according to their acid-base character. This property depends on the acid-base character of the side group. If this group is acidic then the amino acid will be acidic and basic side chains correspond to basic amino acids. For example Aspartic acid is an acidic amino acid as its side chain is acidic and arginine is a basic amino acid as its side chain is basic.



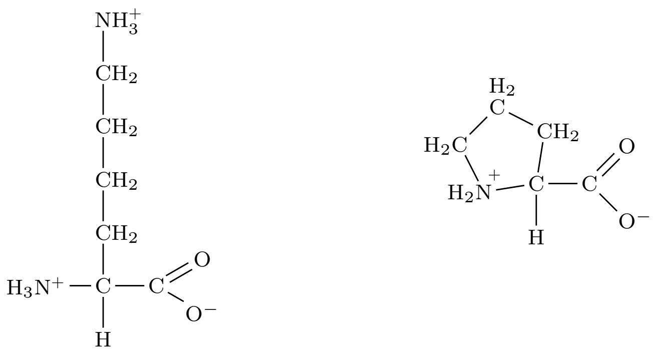
Aspartic Acid (Asp)
Acidic AA



Arginine (Arg)
Basic AA

Sample Problem 135

Classify the following AA as acidic, basic, hydrophilic or hydrophobic:

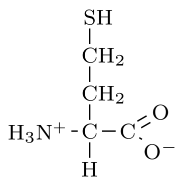


SOLUTION

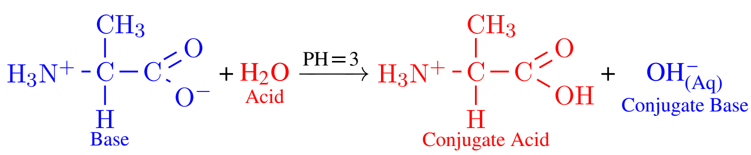
They key to classify the different AA is on its side chain. The AA on the left has a nonpolar side chain made of C and H and hence it will hydrophobic. The AA on the right has a polar side chain with N and at the same time a basic side chain. This AA will be hydrophilic and at the same time basic.

STUDY CHECK

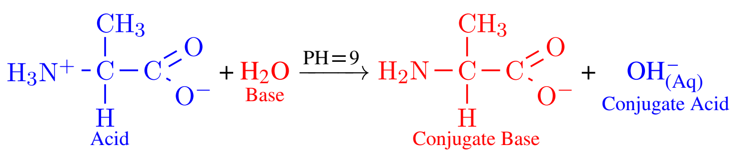
Classify the following AA as acidic, basic or neutral, and hydrophilic or hydrophobic:



`Isoelectronic point of an Amino acid` Amino acids have acid-base character and that means one can modify its charge by modifying the PH. There is a specific PH value at which an AA is neutral. This is called the isoelectronic point (IP) of an Amino acid. For PH value above this point the AA will have negative charge (will be in the form of a conjugate base) and for PH values below this point the AA will have positive charge as the AA will be in a conjugate acid form. For example, the IP of Alanine is 6.1. This means that for PH lower than 6.1 the AA will have positive charge and will be in the form of a conjugate acid:



Differently, if the PH is above 6.1 the AA will have negative charge as the medium will be basic:

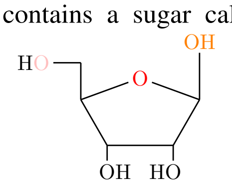


Finally, at the isoelectronic point the AA will be neutral.

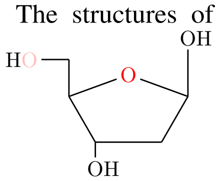
12.6 Nucleic acids

Nucleic acids are long polymers of a unit called nucleotide connected by means of a bond called phosphodiester. There are tow different types of nucleic acids: DNA and RNA. The term DNA stands for deoxyribonucleic acid, whereas the term RNA stands for ribonucleic acid. A nucleotide ies made of a phosphate acid and a nucleoside, and at the same time a nucleoside is made of a sugar and a base. This section cover the properties of nucleic acids.

Sugars in nucleic acids The structure of nucleic acids is indeed sugar. Two different sugars can be found in nucleic acids. Deoxyribonucleic acid (DNA)contains deoxyribose, hence its name, whereas ribonucleic acid (RNA) contains a sugar called ribose. The structures of ribose and deoxyribose are:



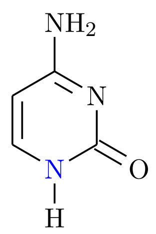
Ribose (R), RNA



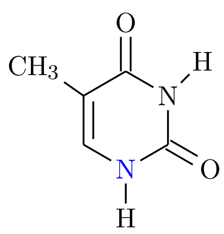
Deoxyribose (DR), DNA

Bases in DNA and RNA Nucleic acids contain a sugar, and also contain a base. Remember that bases are chemicals that produce hydroxyls in water. There are tow different types of nucleic bases: purine bases and pyrimidine bases. At the same time, there are three pyrimidine bases: Cytosine, Thymine, and Uracil. There are two purine bases: Adenine and Guanine. Some of these bases can only be found as part of DNA such as thymine, whereas others can be found only in RNA such as Uracil. Other nucleic bases can be found either on DNA or on RNA. An example of the later is Cytosine. Because of the complex name of these bases often times we refer to them by the first letter. For example, Uracil is also referred as U.

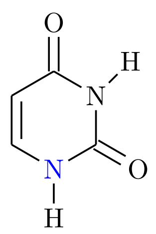
PYRIMIDINE BASES



Cytosine (C)
DNA & RNA

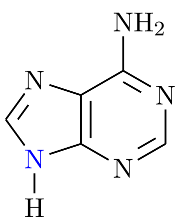


Thymine (T)
DNA

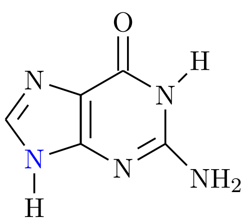


Uracil (U)
RNA

PURINE BASES

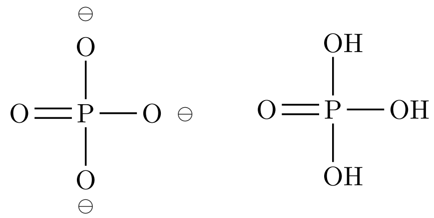


Adenine (A)
DNA & RNA



Guanine (G)
DNA

DNA and RNA contain phosphate At this point we saw tow different sugars part of RNA and DNA. We also saw some bases that also are part of the two different nucleic acids. Still, DNA and RNA are complex molecules that also contain a third element: a phosphate unit. Phosphate results from phosphoric acid: H_2PO_4 after removing two hydrogens:



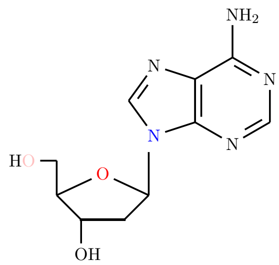
Phosphate   Phosphoric Acid

Nucleoside: a base connected to sugar Nucleic acids are polymers. This means the are the result of the connection between smaller units that over all make the nucleic acids. These units are called nucleosides. A nucleoside is the result of connecting a sugar (either R or DR) with a base (C, T, U, A or G). These tow units are connected by means of a bond called N-glycosidic. This is a similar bond that connect sugars in disaccharides, however involved a nitrogen instead of an oxygen:

$$\mathrm{R-OH} \;+\; \mathrm{HN-R'} \xrightarrow{H_2O} \mathrm{R-N-R'}$$

N-Glycosidic bond

Let us create a nucleoside. In order to do this, we need to select a sugar (for example DR) and a nucleic base (for example adenine). Both will combine be means of carbon number 1 of the sugar and eliminate a water molecule:



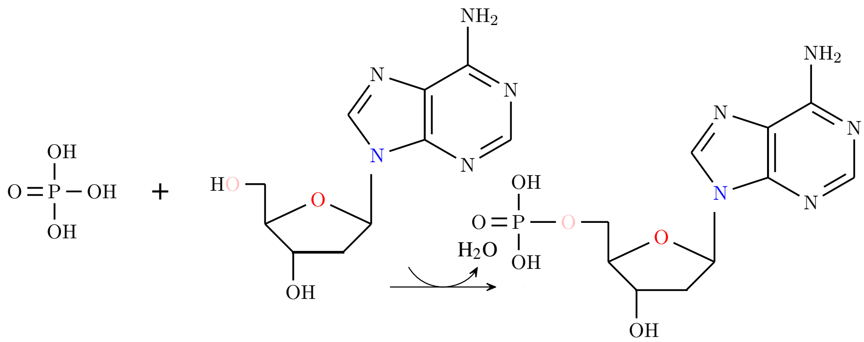
Nucleoside (Sugar+Base)

Nucleotide: a nucleoside connected to phosphate DNA and RNA are composed of numerous units called nucleotides. Nucleotides are the result of a nucleoside reacting with phosphoric acid by means of carbon number 5'. The bond in a nucleotide is called phosphoester bond:

$$\mathrm{R-\overset{\overset{\displaystyle O}{||}}{P}-OH} \;+\; \mathrm{HO-R'} \xrightarrow{H_2O} \mathrm{R-\overset{\overset{\displaystyle O}{||}}{P}-O-R'}$$

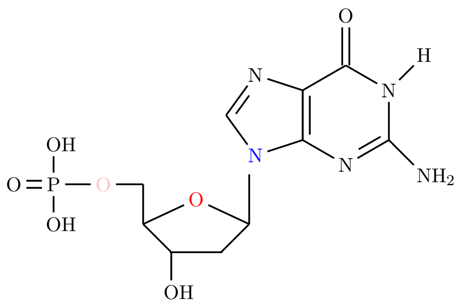
Phosphoester bond

For example, let us consider the nucleoside formed by the reaction of the Deoxyribose sugar and adenide. This nucleoside is called adenosine. After the reaction with phosphoric acid nucleotide is formed called adenosine-5'-monophosphate.

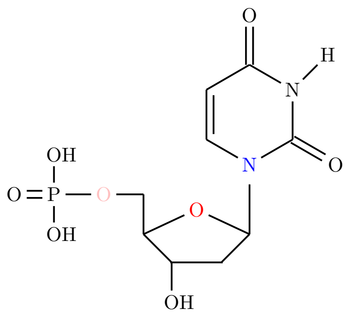


Phosphoric acid Nucleoside Nucleotide (Phosphate+Nucleosite)

Here some other examples of DNA nucleotides. Mind they are all based in Deoxyribose, the sugar that forms only DNA:

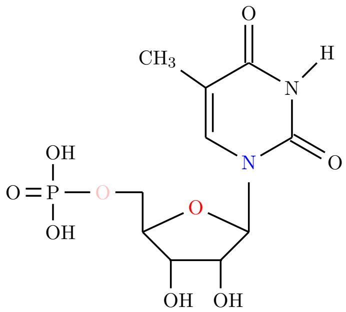


Deoxyguosine-5'-monophosphate

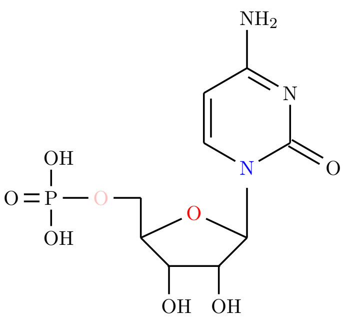


Deoxyuridine-5'-monophosphate

Here some other examples of RNA nucleotides. Mind they are all based in ribose, the sugar that forms only RNA:



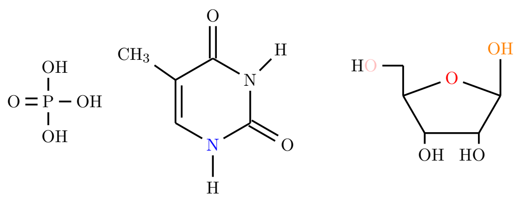
thymidine-5'-monophosphate



cytidine-5'-monophosphate

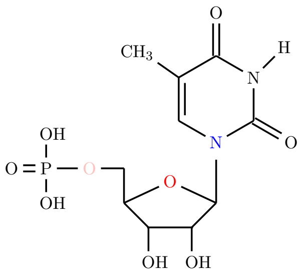
Sample Problem 136

Combine the following molecules to produce a nucleotide:



SOLUTION

A nucleotide is the result of combining a sugar–either ribose or deoxyribose–and a base with phosphate. The sugar given is ribose and the base is thymine, so the nucleotide will be part of RNA:

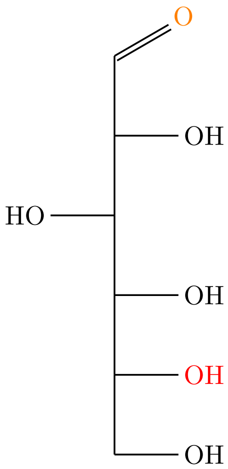


Thymidine-5'-monophosphate

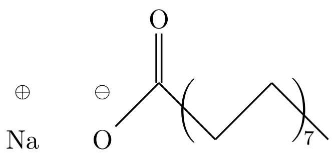
CHAPTER 12

Biomolecules

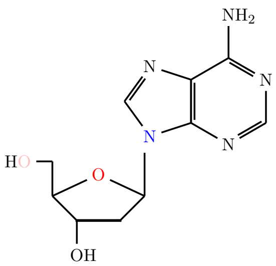
12.1 Classify the following biomolecules as a carbohydrate, a lipid, a protein or a nucleic acid:



12.2 Classify the following biomolecules as a carbohydrate, a lipid, a protein or a nucleic acid:

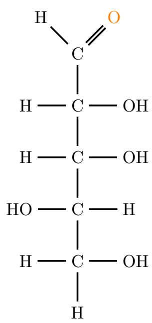


12.3 Classify the following biomolecules as a carbohydrate, a lipid, a protein or a nucleic acid:

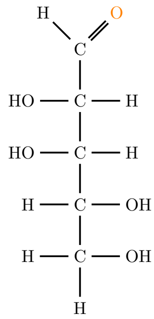


Simple carbohydrates

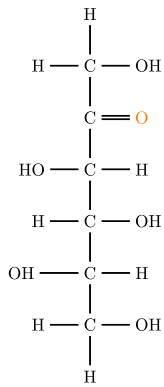
12.4 Classify the following carbohydrate as aldo or keto



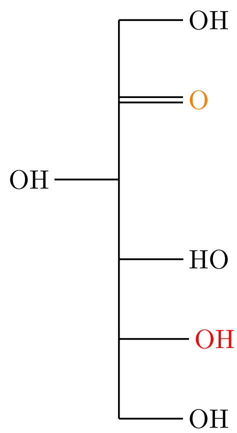
12.5 Classify the following carbohydrate as aldo or keto



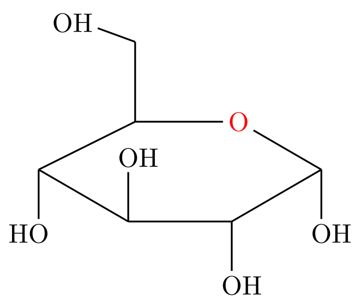
12.6 Classify the following carbohydrate as aldo or keto



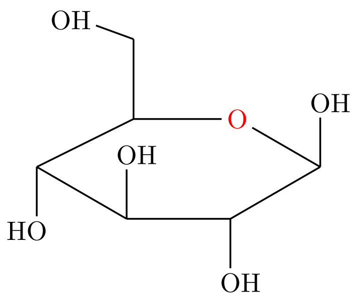
12.7 Find the molecular formula of the following monosaccharide:



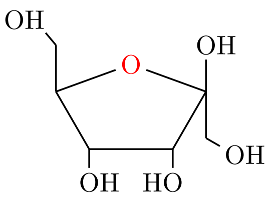
12.8 Classify the following Haworth monosaccharide as α or β.



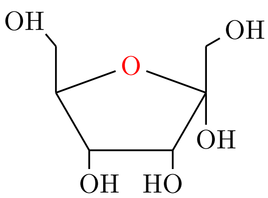
12.9 Classify the following Haworth monosaccharide as α or β.



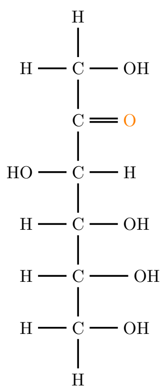
12.10 Classify the following Haworth monosaccharide as α or β.



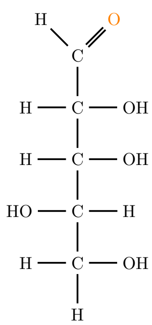
12.11 Classify the following Haworth monosaccharide as α or β.



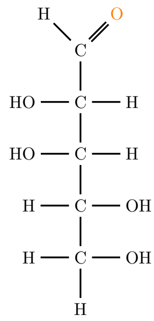
12.12 Classify the following monosaccharide as d or l.



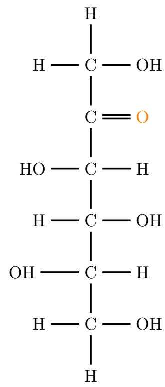
12.13 Classify the following monosaccharide as d or l.



12.14 Classify the following monosaccharide as d or l.

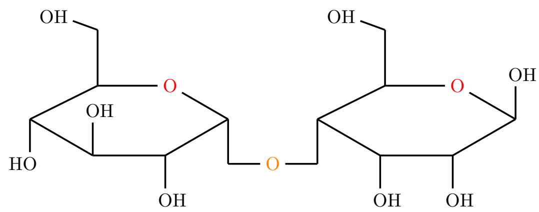


12.15 Classify the following monosaccharide as d or l.

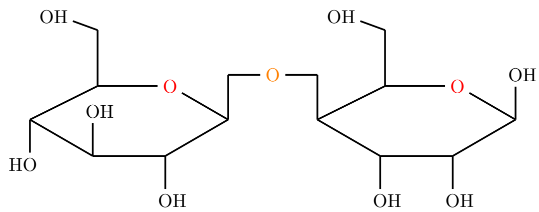


Disaccharides and polysaccharides

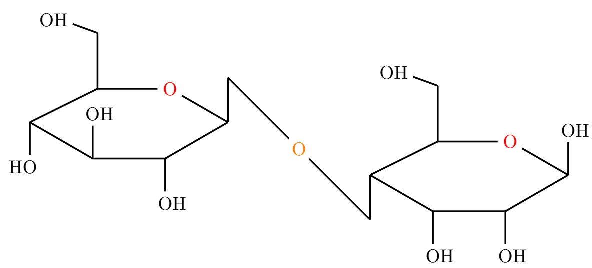
12.16 Classify the following glycosidic bond.



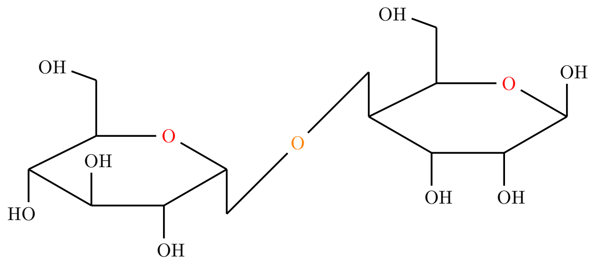
12.17 Classify the following glycosidic bond.



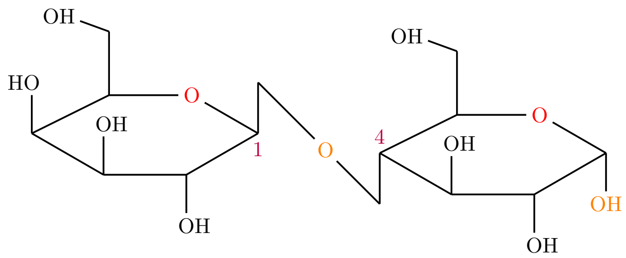
12.18 Classify the following glycosidic bond.



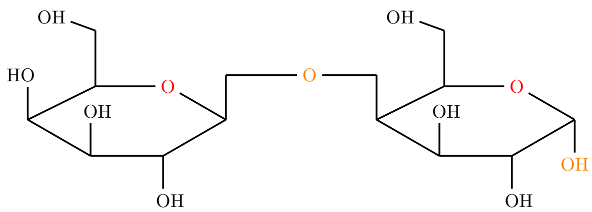
12.19 Classify the following glycosidic bond.

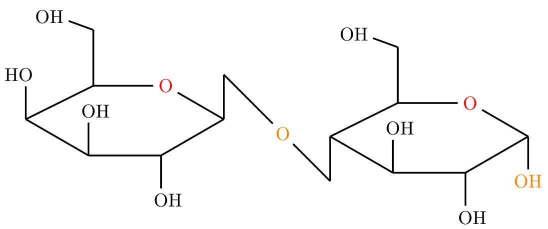


12.20 Enzymes that split disaccharides are only found in the upper reaches of the small intestine (called the duodenum). Find the monosaccharides resulting from the digestion of the lactose:

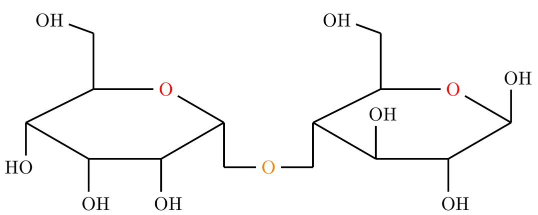


12.21 Which of the following disaccharides is easier to digest:

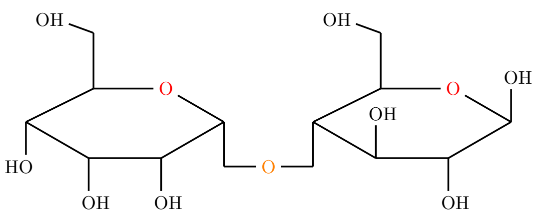




12.22 Classify the following disaccharide as α or β.

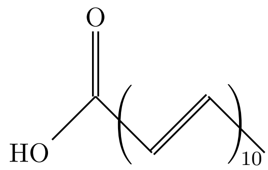


12.23 Classify the following disaccharide as α or β.

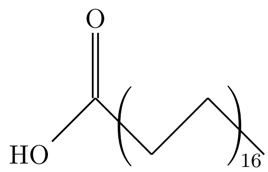


Lipids

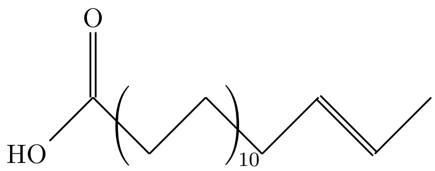
12.24 Classify the following fatty acid as saturated, monounsaturated and polyunsaturated:



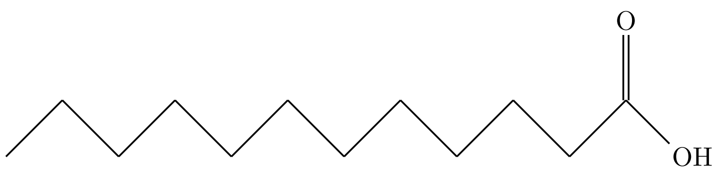
12.25 Classify the following fatty acid as saturated, monounsaturated and polyunsaturated:



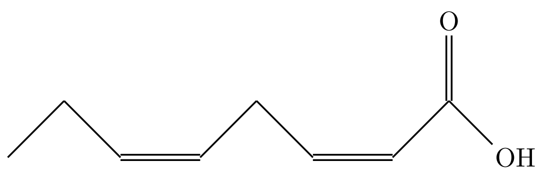
12.26 Classify the following fatty acid as saturated, monounsaturated and polyunsaturated:



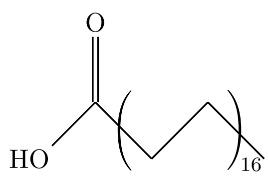
12.27 Give the molecular formula of the following fatty acid:



12.28 Give the molecular formula of the following fatty acid:

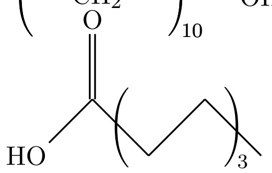


12.29 Give the molecular formula of the following fatty acid:

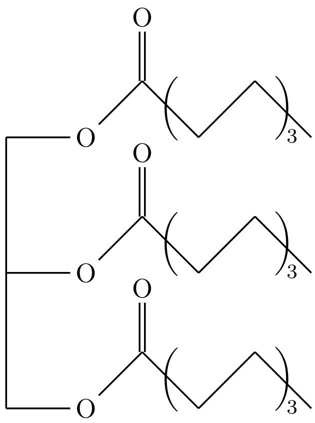


12.30 Give the structure of the wax resulting of mixing the following fatty acid and alcohol:

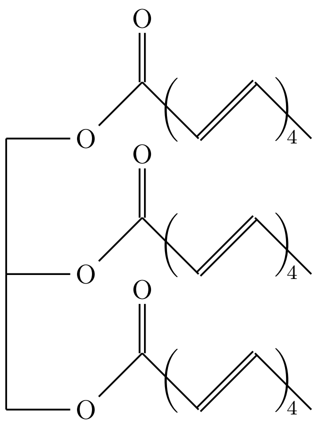
$CH_3-(CH_2)_{10}-OH$ and



12.31 Give the structure of the soap resulting of saponification of the following fat:

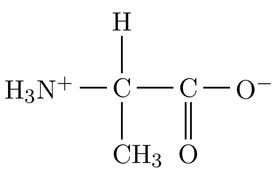


12.32 Give the structure of the fatty acid used to form the following fat:

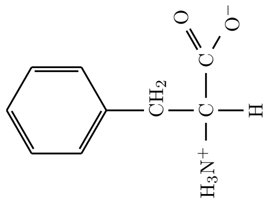


Proteins

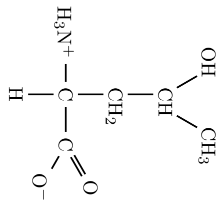
12.33 Identify the side chain of the following amino acid:



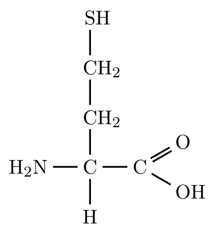
12.34 Identify the side chain of the following amino acid:



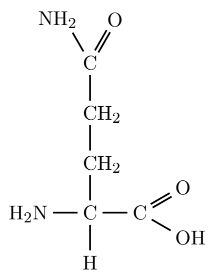
12.35 Identify the side chain of the following amino acid:



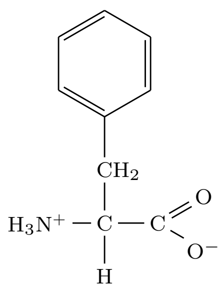
12.36 Write down the zwitterion form of the following AA:



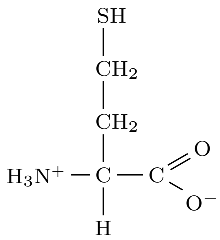
12.37 Write down the zwitterion form of the following AA:



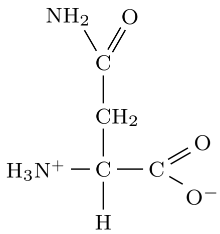
12.38 Classify the following AA as hydrophilic or hydrophobic. If hydrophilic, classify as acidic, basic or neutral:



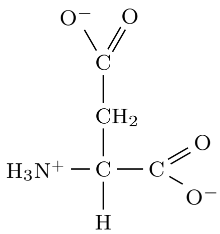
12.39 Classify the following AA as hydrophilic or hydrophobic. If hydrophilic, classify as acidic, basic or neutral:



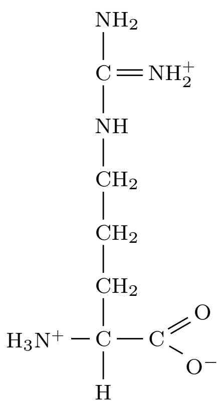
12.40 Classify the following AA as hydrophilic or hydrophobic. If hydrophilic, classify as acidic, basic or neutral:



12.41 Classify the following AA as hydrophilic or hydrophobic. If hydrophilic, classify as acidic, basic or neutral:

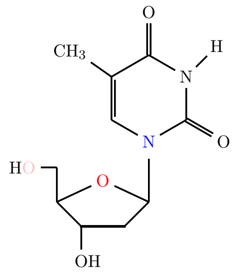


12.42 Classify the following AA as hydrophilic or hydrophobic. If hydrophilic, classify as acidic, basic or neutral:

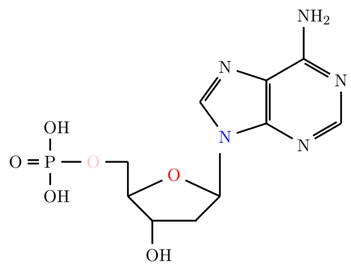


Nucleic Acids

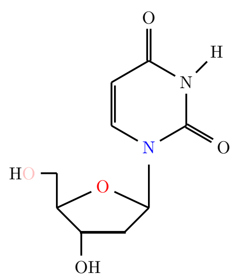
12.43 Classify the following nucleic acid as nucleotide or nucleoside.



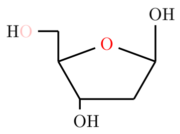
12.44 Classify the following nucleic acid as nucleotide or nucleoside.

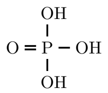


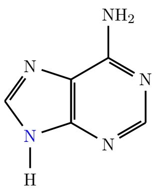
12.45 Classify the following nucleic acid as nucleotide or nucleoside.



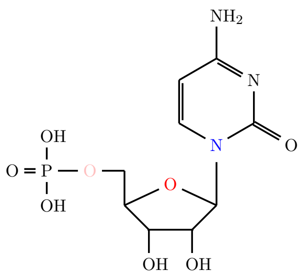
12.46 Combine the following molecules to produce a nucleoside:



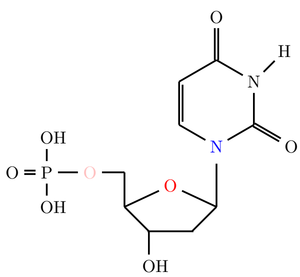


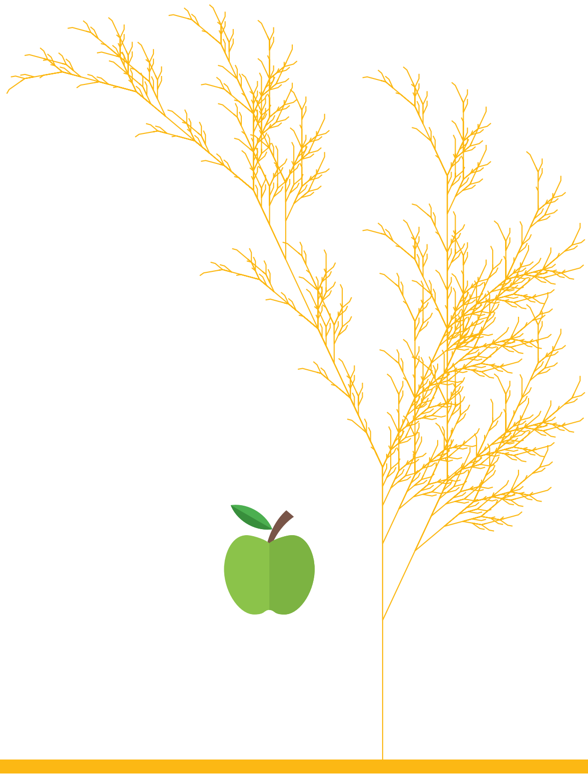


12.47 Indicate whether the following nucleic acid is part of DNA or RNA.



12.48 Indicate whether the following nucleic acid is part of DNA or RNA.





GOB Chemistry

A Comprehensive Set of Imperfect Notes

This set of lectures present content in a simple and clear way, while including numerous worked examples and many problems with solution. In particular, this current version of the manuscript contains more than 90 solved problems and more than 200 problems with solution. It also contains numerous diagrams and graphs specifically developed to clarify the content. The organization of the note intends to help the reader digest the large content typically covered in a GOB or college Chemistry class. Every part ends with a review quiz that assesses content.