College Chemistry

A Comprehensive Set of Imperfect Notes

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Periodic Table of Chemical Elements

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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 each and 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. These notes start with the fundamentals and–at this point–end with solids and liquids. 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 class is not an easy subject. It often gets frustrating due to the terminology or the math content. This guide is developed in chapters and sections in order to break down the very basics of the chemistry concepts. One of my mail 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 in order to find the solution based on the information presented. As you study this set of lecture, I encourage you to read the different section of a chapter, highlight the main ideas and find key words that represent new concepts. Numerous examples are presented along the chapters with the full solution. Numerous examples are also presented without the worked solution, just including the answer. Numerous end of the chapter problems are 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 along this guide to help you focus on the most relevant content t. For example, *yellow notes* are used to indicate important formulas or tables. Also, when the numerical problems get to 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 student often rely on textbooks to learn. These valuable educational resources are often used for a very limited period of time 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 for 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 lecture is indeed imperfect, and hence its title. Yet, it is the result of many hours of work–indeed months of work. Still, it contains typos and often times incorrect answers. 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 use terms and a way of thinking that correspond to the educators'point of view.

This set of lecture does not intent to replace any textbook. Indeed, there are many high-quality textbooks in the literature that I recommend:

- 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

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–helps makes chemistry more apparent to the viewer. One of the weak points of many open education chemistry guides are the visuals. They tend to be simplistic with low quality. This guide extensively relies on images and diagrams and uses Tikz and other open-source tools to generate diagrams. All other images 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
New York City

PART A

1
Measurements

MEASURING is an important part of our everyday lives, and very probably you took several measurements today. You might have drunk a cup of coffee in the morning, or checked the outside temperature with a thermometer. Perhaps you baked a cake following a recipe and measured the mass of flour and sugar using a scale. A cup, a thermometer or a scale are measuring devices.This chapter will cover how to accurately measure properties and more importantly, how to convert convert units. You will, for example, learn how to convert milliliters to centimeters. This chapter will also cover density and you will learn how to compute the density of a substance. By learning how to measure and perform operation with units, you will gain experience performing basic chemistry calculations.

GOALS

1. Identify units and prefixes
2. Introduce/eliminate prefixes
3. Switch prefixes
4. Calculate significant figures
5. Carry density calculations

1.1 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 in different *systems of units* and for example the unit *meter* belongs to a different system than the unit *mile*. In particular here we will address two main systems: the English System used in the US and the Metric System used by most of the industrialized world. The *Metric System* (MS) is used by scientists throughout the world and is the most common measuring system based on the meter. The *International System of Units* (SI) adopted the metric system in 1960 in order to provide additional uniformity for units used the sciences. This chapter will be mostly based on the SI units. In the following we will introduce some common units.

Discussion: why is chemistry important for your career objective? List three reasons why chemistry connects with your career objective.

`Length` What is your height? Length refers to space 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: meter, inch and mile.

`Volume` How much milk do you usually buy? Maybe a gallon. Volume is the amount of space that a substance occupies. A liter (L) 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 every-day life. L, mL and gallon are units of volume. Units of volume are in general cubic units, so for example one litter is the same as one dm^3. We will cover cubit units furthe rin this chapter.

Figure 1.1: Scales measure mass

`Mass` How much do you 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 than with mass.

Figure 1.2: Watches measure time

However, mass and weight are not exactly the same, as weight is a measure of the gravitational pull on an object. It differs depending on the planet you live in. In the metric system, the unit for mass is the gram (g). The SI unit of mass, the kilogram (kg), is used for larger masses such as body weight. Pound, lb, is another unit of mass. The most important units of mass are: g, kg and lb.

`Temperature` How is the weather in NYC today? Is it cold or hot? You use a thermometer to measure temperature and for example assess how hot is an object, or how cold is it 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 from home, or maybe minutes. You probably think of time as years, days, minutes, or seconds. Of all these units, the International System if 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.

Sample Problem 1

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

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

SOLUTION

(a) length; (b) weight or mass; (c) volume; (d) Temperature;

STUDY CHECK

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

(a) 800F (b) 3cL (c) 3cm

Answer: (a) Temperature; (b) volume; (c) length;

Table 1.1: The metric, international system and english units

Measuremts	Metric System	International System (SI)	English System
Length	Meter (m)	Meter (m)	Foot (f)
Volume	Liter (L)	Cubic meter (m^3)	Gallon (gal)
Mass	Gram (g)	Kilogram (kg)	Pound (lb)
Time	Second (s)	Second (s)	Second (s)
Temperature	Celsius (°C)	Kelvin (K)	Fahrenheit (°F)

1.2 Prefixes & Conversion Factors

Let's consider: **k**m, **c**m, m, which can be read as centimeter, kilometer and meter. The word kilo (**k**) and centi (**c**) are called prefixed whereas meter (m) is the unit. Kilometer is larger than meter, whereas centimeter is smaller than a meter. A prefix such as kilo or centi can be attached to any unit to increase or decrease its size. Hence we can talk about a centimeter (**c**m) but also about a centisecond (**c**s) or centiliter (**c**L). The following table lists some of the metric prefixes, their symbols, and their decimal values.

Add Table 1.2 to your flashcard.

Table 1.2: Table containing some Prefixes. For example, 1km is a thousand (1×10^3) meters, and 1ms is (1×10^{-3}) seconds. The prefixes on top of the table are larger than the unit, and for examples 1Tbite is larger than a bite. The prefixed on the bottom are smaller than the unit, and 1fs is smaller than a second.

Prefix	Symbol	Value
tera	T	1×10^{12}
giga	G	1×10^{9}
mega	M	1×10^{6}
kilo	k	1×10^{3}
deci	d	1×10^{-1}
centi	c	1×10^{-2}
milli	m	1×10^{-3}
micro	μ	1×10^{-6}
nano	n	1×10^{-9}
pico	p	1×10^{-12}
femto	f	1×10^{-15}

`How to identify the prefix?` Look for example in cm. Centi (c) is the prefix and means 1×10^{-2} and meter (m) is the unit which refers to lenth. Another example, kg means kilogram. 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.

What would you prefer to own a kilodollar, a dollar or a centidollar? A unit with a prefix can be bigger or smaller than the plain unit–this is the unit without prefix–, depending on the prefix. The following prefixes make the unit smaller: deci, centi, milli, micro, nano, pico and femto. For example a fs (femtosecond) is smaller than a s (second). Differently, the following prefixes make the unit larger: Tera, Giga, Mega. For example a Tb (terabite) is larger than a b (bite). Bite is a unit used in computer science.

Remember: the **pre**fix always comes first as the **c** in **cm**

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

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

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

Remember: **equalities** are written in line whereas **conversion factors** with a fraction.

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

Add Table 1.3 to your flashcard.

Table 1.3: Table containing some common unit equalities.

Unit	Equality
Inches (in)-centimeters (cm)	2.54 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

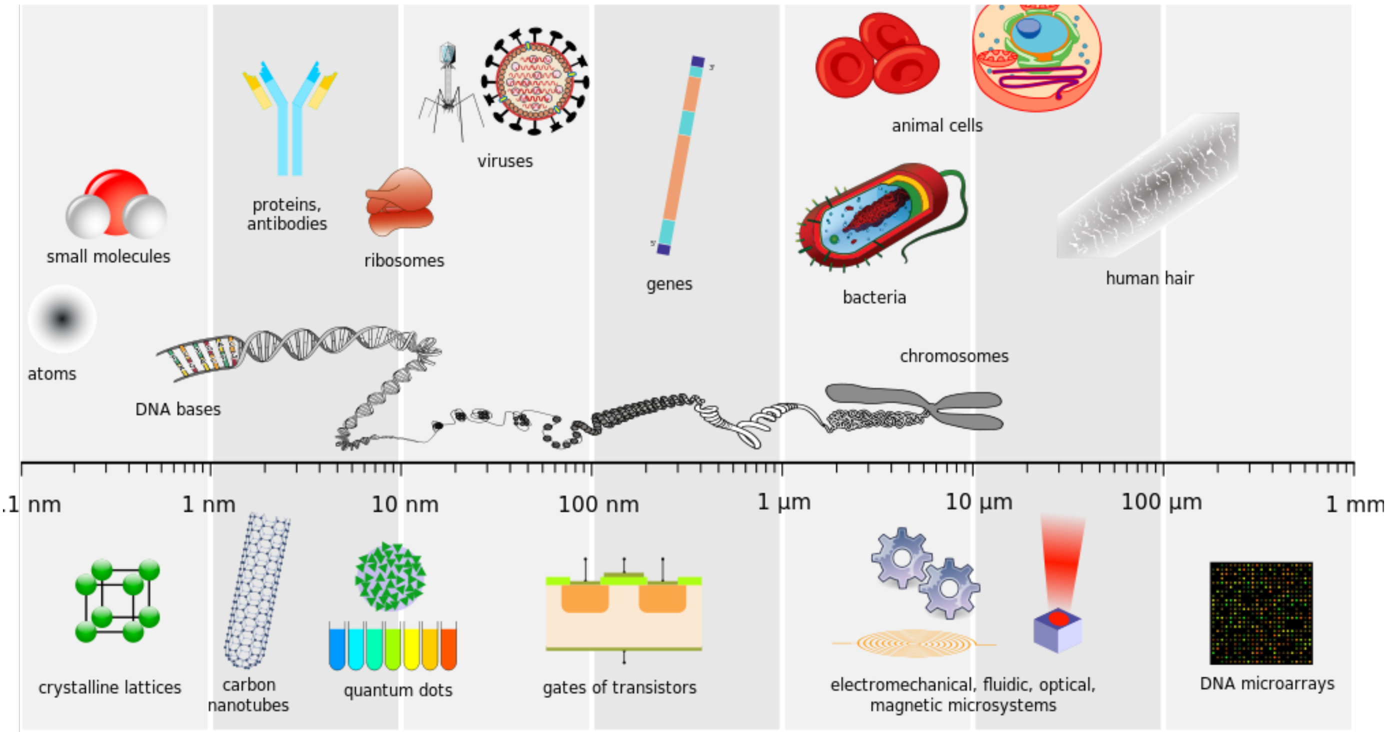


Figure 1.3: The different scales of the matter

Figure 1.4: How many mL do you add to your eye?

Sample Problem 2

Complete each of the following equalities and conversion factors:

(a) 1dm=______m

(b) 1km=______m

(c) $\dfrac{1nm}{\_\_\_\_\_\_ m}$

(d) $\dfrac{\_\_\_\_\_\_ m}{1cm}$

SOLUTION

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

STUDY CHECK

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

(a) 1cs=______s

(b) $\dfrac{\_\_\_\_\_\_ s}{1Ts}$

(c) $\dfrac{\_\_\_\_\_\_ s}{1Ms}$

Answer: (a) 1cs=$1 \times 10^{-2}s$; (b) $\dfrac{1 \times 10^{12}s}{1Ts}$; (c) $\dfrac{1 \times 10^{6}s}{1Ms}$;

1.3 Using Conversion Factors

Unit equalities in the form of conversion factors are used to convert a unit into another. 75
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. An example would be converting centimeters (cm) to millimeters (mm). Let's work on some example.

`Removing or adding prefixes` Imagine that you need to remove a prefix from a unit, and convert 3 km in meters. First, you would need the conversion factor corresponding to the prefix (centi) from Table 1.3. Then you need to arrange the conversion factor placing the prefix in the bottom of the fraction. This will cancel out the prefix from the number and the conversion factor leaving the plain unit on top of the conversion factor (the final unit). 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:

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

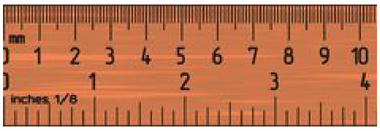


Figure 1.5: Rulers normally have the cm and inch conversion written.

Sample Problem 3

The length of a textbook page is 20cm. Convert 20cm to m.

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 $\frac{1 \times 10^{-2} m}{1cm}$:

$$20\cancel{cm} \times \frac{1 \times 10^{-2} m}{1\cancel{cm}} = 0.2m$$

The units on top and on the bottom of the formula cancel and we get meters.

STUDY CHECK

Convert 100m to km.

Answer: $100\cancel{m} \times \frac{1 \times 10^3 m}{1\cancel{km}} = 100000m.$

`Switching prefixes` In order to switch a prefix into another prefix, such as transforming 30 millimeters (30 mm) into centimeters (cm), you will need two conversion factors: the first corresponds to the prefix to be removed (milli), whereas the second conversion factor corresponds to the prefix to be introduced (centi). You will get the conversion factors from Table 1.2. You will arrange the first conversion factor so that the prefix to be removed cancels from top to bottom and the second conversion factor so that on the top part of the conversion factor you have the prefix to be introduced. For this example:

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

Remember: the number 1×10^{-3} is **scientific notation** and must be typed in the calculator as: 1 $\boxed{\text{EE}}$ -3.

Sample Problem 4

The length of a textbook page is 20cm. How many mm is this lenth.

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:

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

STUDY CHECK

Convert 100mm to km.

$$\text{Answer:}100\cancel{mm}\times\frac{1\times 10^{-3}\cancel{m}}{1\cancel{mm}}\times\frac{1km}{1\times 10^{3}\cancel{m}}=1\times 10^{-4}km.$$

Figure 1.6: The per-feet-square price in Tribeca where The Borough of Manhattan Community College is located in NYC is $\$1,750$

`Using other equalities` How many hours is 300 minutes, or how many centimeters is 2 inches? Some of the units conversion are not based on a power of ten relationship such as the ones in Table 1.2. Table 1.3 lists some of the common equalities that can be easily converted into conversion factor. As an example, the unit equivalency between hours and minutes is $60min = 1h$ and the conversion factor would be $\frac{60min}{1h}$ or $\frac{1h}{60min}$.

Sample Problem 5

Convert 20 in to cm.

SOLUTION

We want to convert 20 inches into centimeters. Inch is not in Table 1.2 but in 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:

$$20\cancel{in}\times\frac{2.54cm}{1\cancel{in}}=50.80cm$$

STUDY CHECK Convert 200mL to drops.

$$\text{Answer: }200\cancel{mL}\times\frac{15drops}{1\cancel{mL}}=3000drops=3\times 10^{3}drops$$

Remember: if you use a power on a power of ten, the power and the ten exponent multiplies, and for example $1\times(10^{-2})^2$ is 1×10^{-4} or $1\times(10^{-4})^3$ is 1×10^{-12}.

Also the power key in your calculator is $\boxed{\wedge}$

`Square or cubic units` How big is your apartment? You might be living in a $750ft^2$ loft in Brooklyn or in a larger house Upstate. Often times we encounter cubic or square units such as cubic centimeter (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 $1cm^2 = 1\times(10^{-2})^2m^2 = 1\times 10^{-4}m^2$. Similarly, if $1mm = 1\times 10^{-3}m$, hence $1mm^3 = 1\times(10^{-3})^3m^3 = 1\times 10^{-9}m^3$.

Sample Problem 6

How many m^2 is $20cm^2$.

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.

$$\text{Answer:} 100\cancel{m^3} \times \frac{1dm^3}{1 \times 10^{-3}\cancel{m^3}} = 1 \times 10^5 dm^3.$$

`Units of volume` 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. In fact one litter is the same as one dm^3 and one ml is the same as one cm^3. In the allied health field, the units *mL* is also written as *cc* as in cubic centiliters.

$$\boxed{1L = 1dm^3 \text{ and } 1cL = 1cm^3}$$

Sample Problem 7

Convert 30 m^3 into L.

SOLUTION

In order to convert m^3 into L we just need to remember that the units Liter 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.

$$\text{Answer: } 40\cancel{L} \times \frac{1\cancel{mL}}{1 \times 10^{-3}\cancel{L}} \times \frac{1cm^3}{1\cancel{mL}} = 0.04cm^3.$$

1.4 Significant Figures

Numbers that results from the measurement of an experimental property are subject to uncertainty. Think about how many eggs are there in your refrigerator, there might be three and this number is an exact and certain number. Differently if you weight a single egg in an scale depending of the type of scale you used 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, calculate the number of significant figures of a number (we will refer to significant figures as SF). Another goal is to estimate significant figures in calculation in order to express the result with the right number of digits.

`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.

Exception 2 *A zero is not significant when used as a placeholder in a large 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 neither the first nor the second rule.

Sample Problem 8

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.

Answer: 4560 has 3SF, 0.123 has 3SF, 1000 has 1SF and 0.0030 has 2SF.

`Significant figures in calculations` There are two different rules that allow you to express the result of a 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:

$$34.3451 + 34.5 = 68.8$$

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.

Rule 2 *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:

$$4500 \times 342 = 1500000$$

the number 4500 has two significant figures, whereas the number 342 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, however, this number needs to be rounded into two significant figures.

Sample Problem 9

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 digit: $88.5 - 87.57 = 0.93 = 0.9$. After that we have:

$$\frac{0.9}{345.13 \times 100}$$

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

$$\frac{0.9}{345.13 \times 100} = 3 \times 10^5$$

STUDY CHECK

Do the following calculation with the correct number of figures: $(24.56 + 2.433) * 0.013$

Answer: 0.35.

1.5 Matter and density

Matter makes up all substances. The materials we use such as glass or wood are all made of matter. Because there are so many kinds of materials, we classify matter by the types of components it contains.This section covers the classification of matter according to pure substances and mixtures. It also elaborates on the types of mixtures one can find.

`Pure Substances` *Pure substances* have a definite composition, that is, are only made of one thing. There are two different types of pure substances: elements and compounds. *Elements* are composed of only one type of atom. Examples are silver, iron, and aluminum that all contain one type of substance, and iron is only made of iron atoms, for example. *Compounds* are combinations of different elements. For example, water, H_2O is made of a combination of hydrogen and oxygen atoms.

`Mixtures` *Mixtures* are physical combinations of different substances. The air we breathe is a mixture of oxygen and nitrogen.

`Types of Mixture` Mixtures can be homogeneous or heterogeneous. In a *homogeneous mixture*–also known as solutions–the composition is uniform throughout the sample. An example of an homogeneous mixture is air, which contains oxygen and nitrogen or salt 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 differenciate the dough and the chocolate.

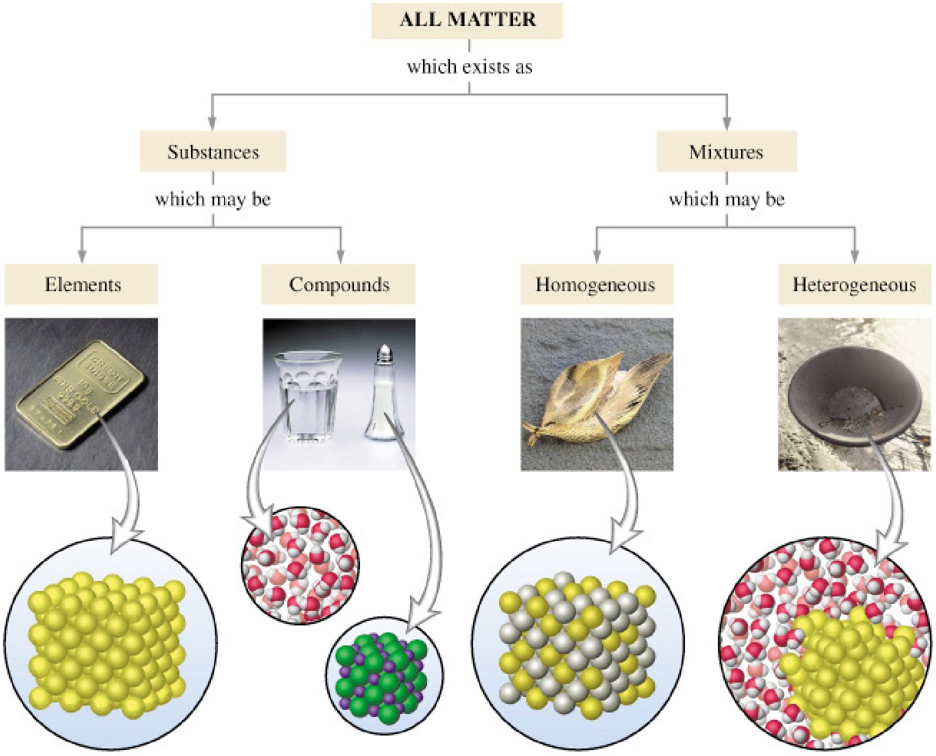


Figure 1.7: Classification of the matter

Sample Problem 10

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 and on the eye you see all the elements of the mixture as a single substance; (c) Sugar is a compound as it is 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 see several of the constituents such as the liquid and the miso paste and the tofu it is a heterogeneous mixture.

STUDY CHECK

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

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

Answer: (a) homog. mix.; (b) compound; (c) element; (d) heterog. mix.

Density This section also covers density of matter. Density refers to the mass of a substance with respect to its volume. This is an unique property for each substance. 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. The formula for density is

Add Formula 1.1 to your flashcard.

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

Density and mixing Ice is solid and it floats on water. The reason for that is density: density of ice is smaller than density of water and hence ice stays on top of water. Objects with density larger than 1 $g \cdot ml^{-1}$ will sink whereas objects with density smaller than 1 $g \cdot ml^{-1}$ will float. If you add a drop of vegetable oil to a glass of water, the drop will float. This is because density of oil that is smaller than 1 $g \cdot ml^{-1}$.

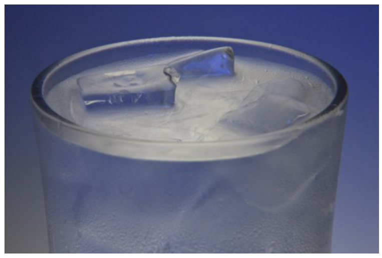


Figure 1.8: Ice floats on water as the density of ice is lower than 1 $\frac{g}{mL}$

Sample Problem 11

In the figure we mixed three liquids of density: A (0.5 $g \cdot ml^{-1}$), B(2 $g \cdot ml^{-1}$) and C(1 $g \cdot ml^{-1}$). Identify each liquid.

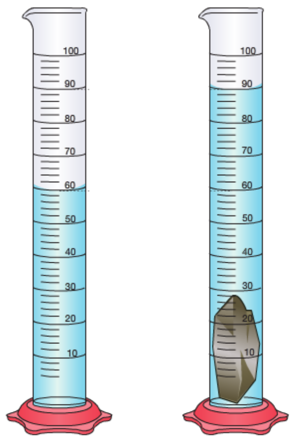
SOLUTION

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.

We can also use density to obtain the volume of a solid without having to physically having to measure the dimensions of the object. We will explain how to do this in the following example:

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: $d = \frac{30g}{30mL} = 1\ g \cdot ml^{-1}$.

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 $(g \cdot ml^{-1})$?

Answer:$d = \frac{226g}{20mL} = 11.3\ g \cdot ml^{-1}$.

CHAPTER 1

Units and measurements

1. A value of 2 L is a measure of

(a) length (d) time
(b) volume
(c) mass (e) temperature

Ans: (b)

2. A value of 5 cm is a measure of

(a) length (d) time
(b) volume
(c) mass (e) temperature

Ans: (a)

3. The metric base unit for volume is the

(a) m^3 (d) mL
(b) L
(c) cm^3 (e) m^2

Ans: (b)

4. The metric base unit for mass is the

(a) g (d) L
(b) Kg
(c) lb (e) cm

Ans: (a)

5. Which of the following is the basic unit of mass in the SI?

(a) pound (c) milligram
(b) kilogram (d) microgram

Ans: (b)

6. Which of the following is a measurement of mass in the metric system?

(a) milliliter (d) Celsius
(b) centimeter
(c) gram (e) meter

Ans: (c)

7. The amount of space occupied by a substance is its

(a) mass (d) length
(b) density
(c) weight (e) volume

Ans: (e)

Prefixes & Conversion Factors

8. Fill the gap in the following unit equalities or conversion factors: $1dm =$ ______m

(a) 1×10^{-1} (d) 1×10^{-6}
(b) 1×10^{3}
(c) 1×10^{-2} (e) 1×10^{12}

Ans: (a)

9. Fill the gap in the following unit equalities or conversion factors: $1cm =$ ______m

(a) 1×10^{-1} (d) 1×10^{-6}
(b) 1×10^{3}
(c) 1×10^{-2} (e) 1×10^{12}

Ans: (c)

10. Fill the gap in the following unit equalities or conversion factors: $\frac{1nm}{\_\_\_\_\_\_ m}$

(a) 1×10^{-1} (d) 1×10^{-6}
(b) 1×10^{-9}
(c) 1×10^{-2} (e) 1×10^{12}

Ans: (b)

11. Fill the gap in the following unit equalities or conversion factors: $\frac{1fs}{\_\_\_\_\_\_ s}$

(a) 1×10^{-1} (d) 1×10^{-15}
(b) 1×10^{-9}
(c) 1×10^{-2} (e) 1×10^{-12}

Ans: (d)

12. Fill the gap in the following conversion factors:

$$20\cancel{cm} \times \frac{\_\_\_\_\_\_ m}{1\cancel{cm}} = 0.2m$$

(a) 1×10^{-2} (d) 1×10^{3}
(b) 1×10^{-3}
(c) 1×10^{2} (e) 1×10^{-6}

Ans: (a)

13. Fill the gap in the following conversion factors:

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

(a) 200 (d) 0.2
(b) 2
(c) 2000 (e) 0.02

Ans: (a)

14. Fill the gap in the following conversion factors:

$$20\cancel{cm} \times \frac{\_\_\_\_\_\_}{\_\_\_\_\_\_} = 7.87in$$

(a) $\frac{1in}{0\cancel{cm}}$ (d) $\frac{2.54in}{1\cancel{cm}}$
(b) $\frac{1in}{1\cancel{cm}}$ (e) $\frac{1in}{2.54\cancel{cm}}$
(c) $\frac{2.54in}{2.54\cancel{cm}}$

Ans: (e)

Using Conversion Factors

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

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

Ans: $1 \cdot 10^{-6}$

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

$$40\cancel{m} \times \frac{1nm}{\underline{\qquad}\,\cancel{m}} = 4 \times 10^{-8} nm$$

Ans: $1 \cdot 10^{-9}$

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

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

Ans: $1 \cdot 10^{-2}; 10 \cdot 10^{3}$

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

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

Ans: $1 \cdot 10^{-1}; 1 \cdot 10^{-2}$

19. Set up the conversion factor to convert 500cm into inches:

$$50\cancel{cm} \times \frac{\underline{\qquad}}{\underline{\qquad}} = \underline{\qquad}\, in$$

Ans: $19.68in$

20. Set up the conversion factor to convert $400cm^2$ into m^2 :

$$400\cancel{cm^2} \times \frac{\underline{\qquad}}{\underline{\qquad}} = \underline{\qquad}\, m^2$$

Ans: $0.04m^2$

Significant Figures

21. Which of the following measurements has three significant figures?

(a) 0.0032 m
(b) 5100 m
(c) 0.510 m
(d) 0.0051 m
(e) 500 m

Ans: (c)

22. Which of the following numbers contains the designated CORRECT number of significant figures?

(a) 0.05600 5 SF
(b) 0.03043 SF
(c) 456 000 3 SF
(d) 1.304 2 SF
(e) 3.12050 4 SF

Ans: (c)

23. The number of significant figures in the measurement of 45.030 mm is

(a) None
(b) 3SF
(c) 4SF
(d) 5SF
(e) 6SF

Ans: (d)

Matter and Density

24. 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 density of the gold?

(a) 10.4 $\frac{g}{mL}$
(b) 6.77 $\frac{g}{mL}$
(c) 1.00 $\frac{g}{mL}$
(d) 23.68 $\frac{g}{mL}$
(e) 19.3 $\frac{g}{mL}$

Ans: (e)

25. Which one of the following substances will float in gasoline, which has a density of 0.66 g/mL?

(a) table salt (2.16 $\frac{g}{mL}$)
(b) balsa wood (0.16 $\frac{g}{mL}$)
(c) sugar (1.59 $\frac{g}{mL}$)
(d) aluminum (2.70 $\frac{g}{mL}$)
(e) mercury (13.6 $\frac{g}{mL}$)

Ans: (b)

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

Ans: $5g \cdot ml^{-1}$

27. 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?

Ans: $31.56mL$

28. Classify a copper wire as

(a) Element
(b) Compound
(c) Homogeneous mixture
(d) Heterogeneous mixture
(e) none of the above

Ans: (a)

29. Classify a chocolate cookie as

(a) Element
(b) Compound
(c) Homogeneous mixture
(d) Heterogeneous mixture
(e) none of the above

Ans: (d)

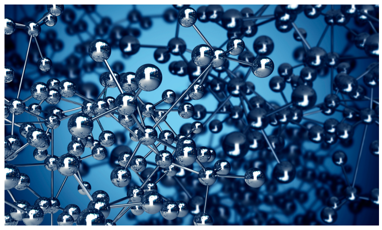
30. Classify vinegar as

(a) Element
(b) Compound
(c) Homogeneous mixture
(d) Heterogeneous mixture
(e) none of the above

Ans: (c)

2

The periodic table: atoms and Elements



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 elements can be different, having distinct number of proton and neutrons. This chapter covers the principles of atomic and electronic structure. You will learn what makes an atom and will be able to quantify the number of particles inside an atom.

GOALS

1. Navigate the periodic table
2. Calculate the number of electrons, protons and neutrons in an atom
3. Calculate average atomic masses
4. Calculate empirical formulas from mass
5. Calculate molecular formulas from empirical formulas

2.1 The periodic table

The periodic table contains all elements that form the matter arranges in columns and rows. Each element has a different given and a symbol that represents their name. The tabular arrangement in the form of rows and columns allow 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 more simple substances. For example aluminum is an element and as it is only made of aluminum atoms and if you analyze the composition of a piece of gold you would only find gold 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 exist in the element's name, the second letter should be lowercase. The chemical symbol for aluminum is Al with capital A and lowercase l. The symbols of all elements can be found in the periodic table (Figure 2.1).

Periods and groups The periodic table contains elements arranged in rows and columns. The horizontal rows are called periods and the vertical columns are called groups. For example, the second period contains lithium (Li), beryllium (Be), boron (B), carbon (C), nitrogen (N), oxygen (O), fluorine (F), and neon (Ne), and the second group contains Beryllium (Be), Magnesium (Mg), Calcium (Ca), Strontium (Sr), Barium (Ba) and Radium (Ra). There are eight periods (period 1 to period 8) and 18 groups. Some of the groups are labeled with an A whereas others are labeled with a B. The groups numbers can be written with roman numbers and a letters (A or B) or with a modern group numbering of 1-18 going across the periodic table. For example, the group 2 (Mg-Ra) can also be called IIA, and the group 13 (B-Ti) is also known as IIIA.

Classification of groups Some of the groups in the periodic table have specific names such as alkali metals, alkaline earth metals, transition metals, 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

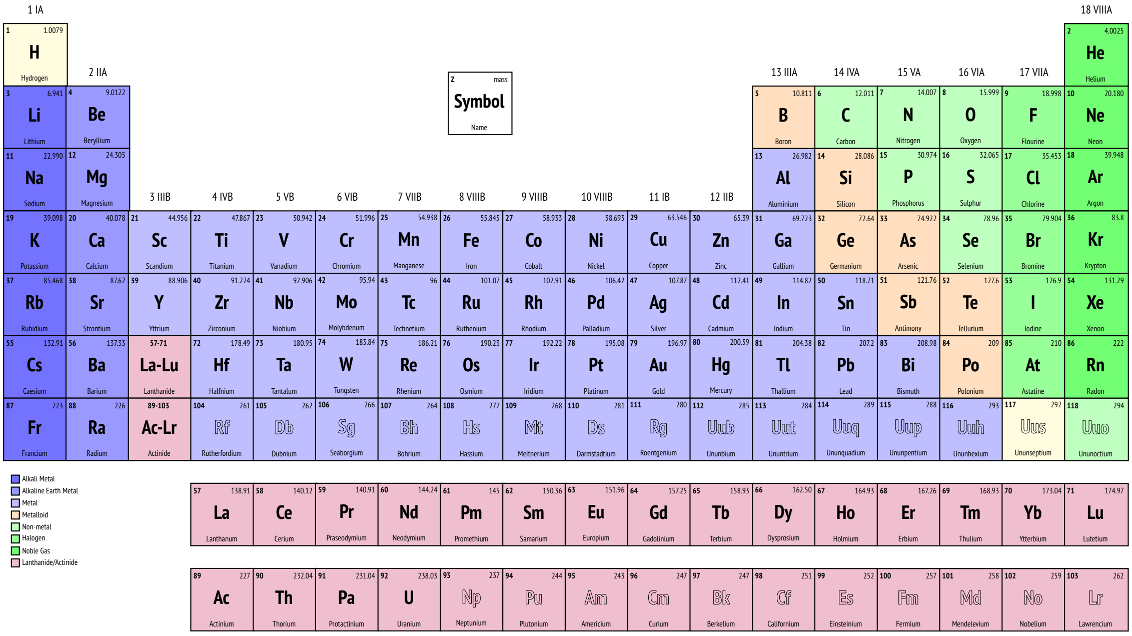


Figure 2.1: The Periodic Table

metals. They are also good conductors of heat and electricity and have low melting points. Alkali earth metals are the group 2A (2) elements: beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra). Transition metals are the elements from group 3 to 12 and they are in the middle of the table. Halogens are the group 7A (17) elements: fluorine (F), chlorine (Cl), bromine (Br), iodine (I), and astatine (At). Halogens are strongly reactive elements. Finally, noble gases are the group 8A (18) elements: helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), and radon (Rn). They are inert and seldom combine with other elements in the periodic table.

How to classify Hydrogen Hydrogen (H) sometimes seems to be put in the wrong spot at the periodic table. 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 a nonmetal.

Metals, Nonmetals, and Metalloids The elements of the periodic table can also be classify as metals, nonmetals, and metalloids. Metals are those elements on the left of the table with a blue color and nonmetals are the elements on the right of the table with a green color. 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 not very shiny, 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 the nonmetals. For example, they are better conductors of heat and electricity than the nonmetals, but not as good as the metals. The metalloids are semiconductors because they can act as both conductors and insulators under certain conditions. A example of metalloids is Silicon (Si) that should not be confused from silicone, a chemical–and not an element–employed in prosthetics.

Sample Problem 13

Answer the following questions:

(a) Give the symbol or name the following elements: Au, Iron, Na and Iodine.

(b) Give the group and period of the following elements, and give the name: Ca, Ir, and C.

(c) Classify as alkali metal, alkali earth metal, transition metal, halogen or noble gas, and give the name: Mg, Li, Co, He, F.

(d) Classify as metal, nonmetal or metalloid, and give the name: Ba, N, Si.

SOLUTION

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

STUDY CHECK

Answer the following questions:

(a) Give the symbol or name the following elements: Ni.

(b) Give the group and period of the following elements, and give the name: Cl.

(c) Classify as alkali metal, alkali earth metal, transition metal, halogen or noble gas, and give the name: Ne.

(d) Classify as metal, nonmetal or metalloid, and give the name: W.

Answer: (a) Nickel; (b) Chlorine: G 17 (VIIA) P3; (c) Neon Noble gas ; (d) Tungstein metal .

2.2 The atom

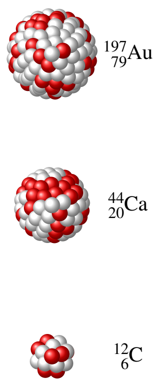


Figure 2.2: The bigger atomic mass the larger the nucleus

Each of the elements in the periodic table are made of atoms, which is the smallest piece of the element that retains the characteristics of that element. Atoms are also the building blocks of all the matter and the materials you use in your everyday life. This section covers the structure of the atom. You will learn how to calculate the number of particles that made an atom and differentiate the atoms of an element.

Atomic Structure Atoms contain three atomic particles: the proton, neutron, and electron. Protons have positive charge (+), and electrons carry negative charge (–). Neutrons on the other hand are neutral–they have no electrical charge. The protons and neutrons are located in the core of the atom, which is called the nucleus, and account for the mass of the atom. Electrons are located in the exterior part of the atoms. When an atoms is neutral it has no charge and the number of electrons and protons are the same. Some atoms have positive charge, resulting of removing electrons. Others can have negative charge as a result of accepting a negatively charged electron.

Atomic and mass number Elements are made of atoms, and each atom of an element is characterized by the atomic number (Z) and the mass number (A). The atomic number (Z) of an element indicates the number of electrons of an atom. The mass number (A) of an element indicates the combined number of protons and neutrons of an atom. Both A and Z for an atom X are indicated in the following form:

$$^{A}_{Z}\mathrm{X}$$

This is called the isotope notation. As an example $^{24}_{12}\mathrm{Mg}$ means that the atomic number of Mg is Z=12 and the mass number is A=24. The atomic number can be found in the periodic table whereas the mass number A is not on the table. By means of 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, we can quickly identify the number of electrons in an atom. 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, corresponds to the mass number minus the atomic number.

Sample Problem 14

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

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

SOLUTION

(a) $^{27}_{12}\mathrm{Mg}$ 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) $^{22}_{10}\mathrm{Ne}$ has 10 electrons and 10 protons, and 12 neutrons. (c) $^{20}_{10}\mathrm{Ne}$ has10 electrons and 10 protons, and 10 neutrons as well. $^{22}_{10}\mathrm{Ne}$ and $^{20}_{10}\mathrm{Ne}$ are both isotopes of Neon.

STUDY CHECK

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

(a) $^{32}_{16}\mathrm{S}$ (b) $^{34}_{16}\mathrm{S}$ (c) $^{36}_{16}\mathrm{S}$

Answer: (a) 16p, 16e and 16 n; (b) 16p, 16e and 18 n; (a) 16p, 16e and 20 n.

> “Nothing exists except atoms and empty space; everything else is opinion.”
> *Democritus*

Isotopes All atoms of the same element are not the same. Some are heavier whereas others are lighter. Isotopes are atoms of the same element with different numbers of neutrons. For example: $^{24}_{12}\mathrm{Mg}$, $^{25}_{12}\mathrm{Mg}$ and $^{26}_{12}\mathrm{Mg}$ are three isotopes of Mg. $^{27}_{12}\mathrm{Mg}$ is heavier than $^{24}_{12}\mathrm{Mg}$ as it contain more neutrons and protons in the nucleus. Each of the isotopes has a specific abundance. Some are more abundant than other. For example, the abundance of $^{24}_{12}\mathrm{Mg}$ is 79%, and the abundance of $^{25}_{12}\mathrm{Mg}$ and $^{26}_{12}\mathrm{Mg}$ is 10% and 11%, respectively.

Atomic mass The atomic mass represents the mass of the atoms of an element. The units of atomic mass are called *amu*, which stands for atomic mass units. This value can be simply found at any periodic table. Using the periodic table provided in this manual 2.1, you can find the atomic mass of each element on top of the symbol at the right side. For example, the atomic mass of oxygen (O) is 15.999 amu and the atomic mass of nitrogen (N) is 14.007 amu. As atoms are made of numerous isotopes–this means different atoms of the same element but with different number of neutrons and hence different weight–the atomic mass you find in the periodic table is the result of including the mass of the different isotopes. That is you need to do an average of the mass of each isotope using values of abundance. In another words: the *atomic mass* of an element, expressed in amu (atomic mass units), is the weighted average of the masses of the individual isotopes of the element. For an element with n isotopes with different masses (A_1, A_2, ..., A_n) and different fractional abundances for each isotope (f_1, f_2, ..,f_n), the atomic mass is given by

$$\text{Atomic mass} = \sum_{i=1}^{n} A_i \cdot f_i$$

Table 2.1: Isotope abundance of some elements

Element	Isotope	% Abundance	Element	Isotope	% Abundance
Hydrogen	1-H	99.9885%	H	1-H	99.9885%
Hydrogen	1H	99.9885%	Silicon	28Si	92.2297%
	2H	0.0115%		29Si	4.6832%
Helium	3He	0.000137%		30Si	3.0872%
	4He	99.999863%	Sulfur	32S	94.93%
Lithium	6Li	7.59%		33S	0.76%
	7Li	92.41%		34S	4.29%
Boron	10B	19.9%		36S	0.02%
	11B	80.1%	Chlorine	35Cl	75.78%
Carbon	12C	98.93%		37Cl	24.22%
	13C	1.07%	Argon	36Ar	0.3365%
Nitrogen	14N	99.632%		38Ar	0.0632%
	15N	0.368%		40Ar	99.6003%
Oxygen	16O	99.757%	Potassium	39K	93.2581%
	17O	0.038%		40K	0.0117%
	18O	0.205%		41K	6.7302%

Sample Problem 15

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:

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

STUDY CHECK

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

Answer: 7.59% and 92.41%.

2.3 An introduction to molecules

The periodic table contains all elements in nature. Elements can combine to form molecules. For example, in the air you can find traces of Argon–this is an element–but you can also find water (H_2O), that results from the combination of two elements, hydrogen (H) and oxygen (O). This section will first introduce you some properties of molecules, without covering their names–this will covered in the following.

Molecular weight Here are two example of molecules: molecular oxygen O_2 and molecular nitrigen N_2. The subscript "2" indicate that each molecule contains two atoms. For example, a O_2 molecular is made of two oxygen atoms. These two molecules have different weight. How do we calculate the weight of a set of molecules? This is actually called the molecular weight (MW), and you can also use different terms to refer to the same property such as: molecular mass, molar mass. All these terms indeed mean the weight a set of molecules. We can calculate the MW by adding the weight of each atom that form the molecule.

Units of molecular weight The units of molecular weight are the same as the units of atomic weight: amu, atomic mass units.

Sample Problem 16

Calculate: (a) The atomic weight of O; (b) the molecular mass of molecular oxygen, O_2

SOLUTION

(a) According to the periodic table the atomic weight (AW) of Mg is 15.999 amu. (b) The molar mass of O_2 is the result of adding the atomic masses of 2O atoms, that is 31.998 amu, close to 32 amu.

STUDY CHECK

Calculate the molar mass of water H_2O and ammonia, NH_3

Answer: 18 and 17 amu.

2.4 Empirical and molecular formula of a chemical

You will find two types of formulas: molecular formulas and empirical formulas. A molecular formula (MF) is the real formula of a chemical such as for example hydrogen peroxide H_2O_2. Differently, the empirical formula (EF) is a simplified formula resulting of dividing the molecular formula by the smallest integer number–different than one. So if H_2O_2 is the molecular formula of sodium peroxide, HO is the empirical formula of the same chemical; as you can see the empirical formula is simplified. The use of empirical formulas comes from the fact that the formulas of all chemicals actually come from experiments, and from experiments one normally can only obtain empirical formulas. The word empirical means "from an experiment".

Molecular weight of empirical formulas and molecular formulas The molecular weight of MF and EF are related by the following formula:

Add this formula to your flashcard.

$$\boxed{n \cdot MW_{EF} = MW_{MF}}$$

where:

MW_{EF} is the molecular weight of the empirical formula

MW_{MF} is the molecular weight of the molecular formula

n is a integer number such as 1, 2, 3...

Empirical formula are just simplified formulas. So when we think about molecular weight we normally have the molecular weight using the real molecular formula in mind. Let us work on an example:

Sample Problem 17

The empirical formula of dichloromethane is $ClCH_2$ and the molecular weigh of the chemical is 98amu. Calculate the molecular formula of dichloromethane.

SOLUTION

Given the empirical formula of dichloromethane one can think of many different molecular formulas, for example: $Cl_3C_3H_6$ or $Cl_2C_2H_4$. From these, and many other, there is only on real molecular formula. How do we calculate the real molecular formula? By comparing the MW of the molecular and empirical formula we can figure out the number of time we need to multiply the MF to obtain the EF. We know the MW is 98amu. Using the EF we can also calculate a MW: $35 + 12 + 2 \cdot 1 = 49$ amu. If we compare both number using the formula:

$$n \cdot MW_{EF} = MW_{MF}$$

we have: $49 \cdot n = 98$; solving we have $n = 2$. Therefore the MF is: $Cl_2C_2H_4$.

STUDY CHECK

The empirical formula of dinitrogen tetroxide is NO_2 and the molecular weigh of the chemical is 88amu. Calculate the molecular formula of dinitrogen tetroxide.

Answer: N_2O_4.

2.5 Determining empirical formulas

We said that the real formula of a chemical is the molecular formula and therefore the real molecular weight of a chemical comes from these formulas. The empirical formulas are obtained from experiments in which a chemical is fragmented and analyzed so that the elements in the molecule and the percentage of each element is determined. Let us work on an example in order to learn the procedure of obtaining molecular formulas.

`Calculating molecular formulas` By means of a experiment, we want to calculate the empirical formula of a chemical given that the chemical contains 2.8 g of nitrogen and 6.4 g of oxygen. In order to calculate the EF we will set up a table. In each column we will add each of the elements that form the molecule. In the first row we will include the grams of each element, in the second we will divide the grams of each element by its atomic weight (AW(N)=14amu,AW(O)=16amu). Among all numbers of the second row, we will select the smallest. Once we have the smallest, we will divide all numbers by the smallest and that will give us round numbers that are the numbers in a empirical formula.

Empirical Formula Table		
	N	O
Grams	2.8g	6.4g
335 AW	14	16
Grams/AW	0.2	0.4
÷ by smallest	1	2
Formula	N_1O_2=NO_2	

Sample Problem 18

The mass percentage composition of a compound is: 18.59% O, 37.25% S, and 44.16% F. Calculate its empirical formula.

SOLUTION

We will set up the the molecular formula table, knowing that the percentage are mass percentages, that is the mass of each element in the chemical, hence they should go in the grams row. Also the atomic weights of O, S and F are 16, 32 and 19 amu.

Empirical Formula Table			
	O	S	F
Grams	0.1859g	0.3725g	0.4416g
AW	16	32	19
Grams/AW	0.0116	0.0116	0.0232
÷ by smallest	1	1	2
Formula		OSF_2	

STUDY CHECK

What is the empirical formula of a compound if a sample contains 10.28 g of C, 1.71 H and 12.71 g of oxygen?

Answer: CH_2O.

CHAPTER 2

The Periodic Table

1. The atomic symbol for Gold is:

(a) Go (c) G (e) Ol
(b) Au (d) Ca

Ans: (b)

2. The atomic symbol for aluminum is:

(a) Al (c) A (e) Ag
(b) Am (d) Sn

Ans: (a)

3. The atomic symbol for iron is:

(a) Ir (c) Fe (e) Ir
(b) Fs (d) In

Ans: (c)

4. Ca is the symbol for:

(a) Carbon (c) Cobalt (e) Cadmium
(b) Calcium (d) Copper

Ans: (b)

5. Which of the following elements is a metal?

(a) Nitrogen (c) Calcium (e) Iodine
(b) Lithium (d) Iron

Ans: (d)

6. Which of the following elements is a alkaline metal?

(a) Nitrogen (c) Calcium (e) Ruthemium
(b) Lithium (d) Iron

Ans: (b)

7. Which of the following elements is a nonmetal?

(a) Nitrogen (c) Calcium (e) Iodine
(b) Lithium (d) Iron

Ans: (a)

8. Which of the following elements is a halogen?

(a) Nitrogen (c) Calcium (e) Iodine
(b) Lithium (d) Iron

Ans: (e)

9. What is the symbol of the element in Period 4 and Group 2?

(a) Be (c) Ca (e) Si
(b) Mg (d) C

Ans: (c)

The Atom

10. 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.

Ans: (b)

11. The atomic number of an atom is equal to the number of

(a) nuclei (d) electrons plus protons.
(b) neutrons
(c) neutrons plus protons. (e) electrons

Ans: (e)

12. The mass number of an atom is equal to the number of

(a) nuclei (d) electrons plus protons.
(b) neutrons
(c) neutrons plus protons. (e) electrons

Ans: (c)

13. The mass number of an atom is equal to the number of

(a) electrons (c) neutrons plus protons.
(b) neutrons (d) protons

Ans: (c)

14. Consider a neutral atom with 30 protons and 34 neutrons. The atomic number of the element is

(a) 30 (c) 34 (e) 94
(b) 32 (d) 64

Ans: (a)

15. Consider a neutral atom with 30 protons and 34 neutrons. The mass number of the element is

(a) 30 (c) 34 (e) 94
(b) 32 (d) 64

Ans: (d)

16. 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.

Ans: 64%

An introduction to molecules

17. Calculate the molecular mass of the following molecule: CCl_2F_2

Ans: 121 amu

18. Calculate the molecular mass of the following molecule: C_4H_{10}

Ans: 58 amu

19. Calculate the molecular mass of the following molecule: C_4H_{10}

Ans: 58 amu

20. Calculate the molecular mass of the following molecule: $C_6H_{10}O_8$

Ans: 210 amu

Empirical and Molecular Formulas

21. What is the empirical formula of a compound if a sample contains 2.8 g of nitrogen and 3.2 g of oxygen?

Ans: NO

22. What is the empirical formula and the molecular formula of a compound if a sample contains 3 g of C, 0.5 H and 4 g of oxygen? MW=60amu

Ans: CH_2O

23. What is the empirical and molecular formula of a compound with a percent composition of 49.47% C, 5.201% H, 28.84% N, and 16.48% O, if its molecular mass is 194.2 amu.

Ans: $C_4H_5N_2O$

24. A 1.587 g sample of a compound containing N and O was analyzed finding a composition of 0.483 g of Nitrogen and 1.104 g of Oxygen. Calculate the empirical formula of the compound.

Ans: NO_2

REVIEW PART A

1. The metric base unit for length is the
   (a) meter (c) millimeter (e) foot
   (b) inch (d) kilometer
2. The amount of space occupied by a substance is its
   (a) mass. (c) weight. (e) volume.
   (b) density. (d) length.
3. Which of the following numbers is the smallest?
   (a) 4.0×10^{-6} (c) 4.0×10^{-2} (e) 4.0×10^{-12}
   (b) 4.0×10^{-8} (d) 4.0×10^{15}
4. Which of the following numbers contains the designated CORRECT number of significant figures (SF)?
   (a) 0.04300 (5 SF) (d) 1.04 (2 SF)
   (b) 0.00302 (2 SF) (e) 3.0650 (4 SF)
   (c) 156 000 (3 SF)
5. 5.21 cm is the same distance as
   (a) 0.0521 m. (d) 0.00521 km.
   (b) 52.1 dm. (e) 5210 m.
   (c) 5.21 mm.
6. Which of the following measurements are NOT equivalent?
   (a) 25 mg = 0.025 g (d) 84 cm = 8.4 mm
   (b) 183 L = 0.183 kL (e) 24 dL = 2.4 L
   (c) 150 msec = 0.150 sec
7. 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 density of the gold in $g \cdot mL^{-1}$?
   (a) 10.4 (c) 1.00 (e) 19.3
   (b) 6.77 (d) 0.0518
8. How many conversion factors can be derived from the equality $1\text{cm} = 1 \times 10^{-1}\text{m}$?
   (a) two (d) unlimited
   (b) three (e) none
   (c) four
9. The number 0.00300 expressed in scientific notation becomes
   (a) 3×10^{-2} (c) 3×10^{3} (e) 0.00300
   (b) 3.00×10^{-3} (d) 3.00×10^{3}
10. In which of the following pairs of numbers does each member of the pair contain the same number of significant figures?
    (a) 11.0 and 11.00 (d) 0.000066 and 660000
    (b) 600.0 and 60 (e) 0.003 and 61300
    (c) 0.05700 and 0.0570
11. What is the mass, in grams, of 30.0 mL of a liquid if its density is 1.20 g/mL?
    (a) 10.0 g (c) 10 g
    (b) 36 g (d) 360 g
12. The number 103.17 expressed in scientific notation becomes
    (a) 1.03×10^{-2} (c) 1.03×10^{2}
    (b) 1.0317×10^{-2} (d) 1.0317×10^{2}
13. The SI Unit standard for mass is the:
    (a) Pound (lb) (c) slug (sl)
    (b) gram (g) (d) kilogram (kg)
14. What quantity is measured using the following unit: cm^2
    (a) length (c) area
    (b) volume (d) mass
15. The number of significant figures in the measurement of 45.030 mm is
    (a) none. (c) six. (e) three.
    (b) four. (d) five.
16. What is 6.5 m converted to inches?
    (a) 1651 in (c) 260 in (e) 255.9 in
    (b) 39 in (d) 1700 in
17. Au is the symbol for
    (a) gold. (d) aluminum.
    (b) silver. (e) sodium.
    (c) argon.
18. Which of the following elements is a nonmetal?
    (a) nitrogen (d) silver
    (b) sodium (e) calcium
    (c) iron
19. What is the symbol of the element in Period 4 and Group 2?
    (a) Be (d) C
    (b) Mg (e) Si
    (c) Ca
20. A compound is 40.0 % carbon, 53.3 % oxygen, and 6.66 % hydrogen. What is its empirical formula?
    (a) $C_4O_5H_7$ (c) $C_1O_1H_2$
    (b) $C_1O_2H_3$ (d) None

21. A compound contains, by mass, 40.0% carbon, 6.71% hydrogen, and 53.3% oxygen. A 0.320 mole sample of this compound weighs 28.8 g. The molecular formula of this compound is:

22. Convert 10^3 nm to mm

23. Complete the following unit conversion :

$$10cm \times \frac{\underline{\qquad\quad} m}{\underline{\qquad\quad} cm} \times \frac{\underline{\qquad\quad} Km}{\underline{\qquad\quad} m} = \underline{\qquad} Km$$

24. Complete the following unit conversion involving square units:

$$50cm^2 \times \frac{\underline{\qquad\quad} m^2}{\underline{\qquad\quad} cm^2} \times \frac{\underline{\qquad\quad} mm^2}{\underline{\qquad\quad} m^2} = \underline{\qquad} mm^2$$

25. A bottle of maple syrup contains 70 mL of syrup. What is the mass of the syrup if the density of maple syrup is 1.33 $\mathrm{Kg \cdot L^{-1}}$.

26. A graduated cylinder contains 155 mL of water. A 15g piece of iron (density = 7.86 $\mathrm{g \cdot mL^{-1}}$) and a 20.0g piece of lead (density = 11.3 $\mathrm{g \cdot mL^{-1}}$) are added. What is the new water level, in milliliters, in the cylinder?

Answers:

1. (a)	**2.** (e)	**3.** (e)	**4.** (c)
5. (a)	**6.** (d)	**7.** (e)	**8.** (a)
9. (b)	**10.** (d)	**11.** (b)	**12.** (d)
13. (d)	**14.** (c)	**15.** (d)	**16.** (e)
17. (a)	**18.** (a)	**19.** (c)	**20.** (c)
21. CH_2O	**22.** 10^{-3}mm	**23.** 10^{-4}Km	**24.** 5000$\mathrm{mm^2}$
25. 93.1g	**26.** 158.7mL		

PART B

3

Chemical naming

ALL elements in the periodic table with the exception of the noble gases–He, Ne, Ar, Kr, Xe and Rn–combine to produce chemical compounds. Most of these chemicals are useful in your every day life, and you drink water, use Clorox to clean your house of baking soda to get rid of a stinky refrigerator. In this chapter you will learn not only how to name these chemicals but also to read chemical formulas. However, 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.

GOALS

1. Name and formulate ionic compounds
2. Name and formulate covalent compounds
3. Name and formulate acids and bases
4. Name and formulate oxosalts
5. Name and formulate common chemicals

3.1 Ions & ionic charges

Atoms gain and loose 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-}.

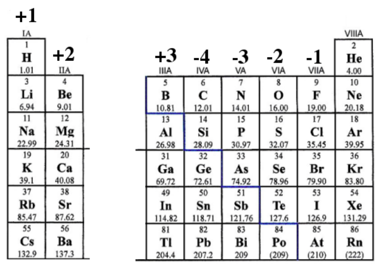


Figure 3.1: Figure with the different ionic charges.

`Discussion:` think about your household and the chemicals you use in your everyday. List three chemicals you found around you, with its correct chemical name

`Cations` Atoms that loose electrons become positively charged. These ions are called cations. Example of cations are Li^+ or Mg^{2+}.

`Anions` Atoms that gain electrons become negatively charged, as electrons have negative charge. These ions are called anions. Example 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.

`Ionic charges` How do we know that hydrogen produces a H^+ ion and nitrogen a N^{3-} anion. The charge of an ion is called ionic charges, and the numbers are coming from the periodic table. H, Na or K are in the group IA (left of the table) and hence the ionic charge will be 1+. Similarly, Mg or Ca are in the group IIA (left of the table) and hence the ionic charge will be 2+. Differently, F, Cl or Br are in the group 7A (right of the table) and its charge will be 1−. Oxygen is in group 6A (right of the table) and the ionic charge will be 2−. Figure 3.1 contains all ionic charges. What is the element is not in this list? In that case, very probably it will have several ionic charges and this charge has to be given to you. An example would be Fe, which ionic charge is not in

Figure 3.1 as iron can have several ionic charges.

Sample Problem 19

Identify the correct ionic state of: Cl, K, O and 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^{1-}, K^{1+}, O^{2-} and C^{4-}.

STUDY CHECK

Identify the correct ionic state of: N and Br.

Answer: N^{3-} and Br^{-}.

Figure 3.2: Water is a covalent compound

3.2 Covalent compounds

Covalent compounds are chemicals resulting from the combination of nonmetallic elements. And example is CO_2, which results of combining carbon (a non metal) with oxygen (a non metal).

The covalent bond The atoms in a covalent compound are connected by a bond and this bond is referred as covalent 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 in the bond.

Covalent naming In order 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 that indicate the number of atoms in the molecule. See Table **??** for a list of the different equivalencies between prefixes and number. The prefix mono is normally omitted. As an example, the formula CH_4 is named as carbon tetrachloride. Similarly, a covalent chemical name can be translated into a formula, and the formula for carbon monoxide would be CO. When the vowels o and o or 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.

Prefix	number
Mono	1
Di	2
Tri	3
Tetra	4
Penta	5
Hexa	6
Hepta	7
Octa	8
Nona	9
Deca	10

Figure 3.3: Prefixes used to name covalent compounds

Properties of covalent compounds At normal conditions, covalent compounds may exist as a solid, a liquid, or a gas. 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.

Sample Problem 20

Name of give the name of the following covalent chemicals:

Formula	Name
NO	
CS_2	
	Sulfur Dioxide
	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 oxide. 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: SCl_2 and diboron thrioxide.

Answer: sulfur dichloride and B_2Cl_3.

3.3 Ionic compounds

Ionic compounds are chemicals resulting from the combination of a nonmetallic element with a metallic element. And example is NaCl, which results of combining sodium (a metal) with chloride (a non metal).

Figure 3.4: Table salt is an ionic compound

The ionic bond The atoms of an ionic compound are connected by a bond and this bond is referred as ionic bond. In an ionic bond, one element gives away electrons and the other one receives electrons. 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 bond 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 and a negative ions. 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. 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 diving 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. Another example that involves simplifying the formula is the chemical resulting of combining Ca^{2+} and C^{4-}. After combining the charges we obtain Ca_4C_2 that needs to be simplified dividing by the smallest number leading to Ca_2C.

Sample Problem 21

Combine the following ions or give the ions given the final compound:

Ions	Combination
Li^{+} and O^{2-}	
Ca^{2+} and O^{2-}	
	Li_3N
	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: Na^{+} and F^{-} and Na_3N.

Answer: NaF; Na^{+} and N^{3-}.

Simple ionic naming In order 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. As an example, the formula NaCl is named as sodium chloride and $MgCl_2$ is named magnesium chloride. The subscripts in the formulas correspond to the ionic charges of the opposite atom and these charges cross from top to bottom. For example,$MgCl_2$ results of the combination of Mg^{2+} and Cl^{-} so that the number 2 in $MgCl_2$ nearthe Cl atom is coming from the Mg^{2+}. The sign of the charges are not important and only indicate which element goes first in the formula: the positive element first following by the negative element. For example the result of combining Na^{+} and Cl^{-} is NaCl and not ClNa as Na has positive ionic charge and has to appear first in the formula.

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

Sample Problem 22

Name or give the formula for the following ionic compounds:

Formula	Name
MgO	
Mg_3N_2	
	Lithium nitride
	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: Sodium fluoride and Na_3N.

Answer: NaF; Sodium nitride.

Complex ionic naming The ionic chemical NaCl results from the combination of Na^+ and Cl^-. The ionic charges of Na and Cl are given in Figure 3.1 according to the group. If the ionic chemical contains a transition metal with variable ionic charge, that is, which is not in Figure 3.1 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.

Formulate complex ionic chemicals In this section we will learn how to name ionic chemicals containing a metal with several possible charges, that is a metal which is not included in Figure 3.1. The charge of the metal has to be given in the name. As an example, the formula for Nickel(III) oxide is Ni_2O_3. The reason for the formula is the combination of Nickel(III) Ni^{3+} and oxygen O^{2-}, that gives Ni_2O_3, after crossing the charges from top to bottom. Another example is Nickel(II) oxide formulated as NiO. This results from combining Nickel(II) Ni^{2+} and oxygen O^{2-} that gives Ni_2O_2. After simplifying one obtains NiO.

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. In order 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 3.1, lets 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.

Sample Problem 23

Name or give the formula for the following ionic compounds:

Formula	Name
MnO	
Fe_3N_2	
	Cobalt(II) carbide
	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

Acid	Name
HCl	Hydrochloric acid
HI	Hydroiodic acid
HF	Hydrofluoric acid
HCN	Hydrocyanic acid

Figure 3.5: List of hydracids

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: Manganese(IV) oxide and AuCl.

Answer: MnO_2; Gold(I) chloride.

3.4 Naming acids & bases

In this section we will learn how to name acids and bases. Acids normally have common names and its naming does not follow modern rules. Differently, bases are named in a standard way.

`Bases` Bases are chemicals that end in OH or more specifically in H. Examples are NaOH or $Ca(OH)_2$. The name of a base starts by the name of the element of the first atom finishing by using the word *hydroxide*. An example is NaOH named as *sodium hydroxide*, or $Ca(OH)_2$, which is *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^-.

`Acids` Acids are chemicals that start its formula with an H. For example, HCl or H_2SO_4. HCl is an hydracid and is named as *hydrochloric acid*, whereas H_2SO_4 is an oxacid 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 3.1 contains a list of the most important oxoacids and Figure 3.3 lists a few hydracids.

Add this table into your flashcard.

Sample Problem 24

Name or give the formula for the following acids and bases. Indicate whether the compound is an acid or a base.

Formula	Acid/Base	Name
HCN		
KOH		
		Carbonic acid
		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: phosphoric acid and $Mg(OH)_2$.

Answer: H_3PO_4; magnesium hydroxide.

Table 3.1: Names of oxoacids and oxosalts

Element	Acid	Acid Name	Salt	Salt Name
Manganese	$HMnO_4$	Permanganic Acid	$MnO_4{}^-$	Permanganate
	$HMnO_3$	Manganic acid	$MnO_3{}^-$	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	$3\,BO_3{}^{3-}$	Borate

`Oxidation states of oxacids` Consider the following set of acids: $HClO$, $HClO_2$, $HClO_3$ and $HClO_4$. These acids have different number of oxygen atoms in its formula. We say Cl in these acids have different oxidation state or different oxidation number. This section will cover the calculation of the oxidation state of the central atom of an oxacid. Let us address the oxacid: $H\underline{Cl}O_3$. The goal if to calculate the oxidation number if the underline element, Cl. In order to do this we will follow a set of simple rules: we will use the valences as the oxidation number of the elements to the right and to the left of the central atom. We will assign an unknown oxidation state of x to the central atom. After that we will set up a equation so that the sum of all oxidation numbers equals to 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. The importance of the oxidation state of the central elements of an oxacid is due to the fact that acids with large oxidation state, that is with positive oxidation state, are called oxidizing acids. Those acids are in general very reactive and will tend to attack metals. Examples of these acids are: HNO_3 and H_2SO_4. Finally, is important to note that ultimately the oxidation state of an element is not necessarily related to the number of oxygens in its molecule. The oxidation state of an element is related to the number of electrons that the element has. The more electrons the smaller–the more negative–the oxidation state is.

Sample Problem 25

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 $H_2\underline{S}_2O_4$ named Sulfuric acid.

SOLUTION

We will set up the redox formula for the first acid, given that the redox number of H is $+1$ and the redox number of O -2.

$$2 \cdot (1) + 2 \cdot x + 6 \cdot (-2) = 0$$

The oxidation state of S in $H_2\underline{S}_2O_3$ is $+5$. For the second acid:

$$2 \cdot 1 + x + 4 \cdot (-2) = 0$$

that gives a redox of -6. If we compare both acids the larger the redox number the more oxidixing the acid is. Therefore, $H_2\underline{S}_2O_3$ is more oxidizing than $H_2\underline{S}_2O_4$.

STUDY CHECK

Calculate the redox number of the following acids: H_2MnO_4 and $H_2Cr_2O_7$.

Answer: $+6$.

3.5 Naming complex salts & common chemicals

Salts, or oxosalts are the result of mixing an oxacid and a base. They look more complex than simple ionic or covalent chemicals as they have at least three different elements. An example would be $CaSO_4$ called calcium carbonate. Another example is $NaHSO_4$ which is called sodium hydrosulfate. This section will cover the naming of complex oxosalts and hydrosaltas, as well as hydrates, that are oxosalts containing water molecules inside its structure. Before being able to name these complex chemicals it is convenient to practice combining ions.

`Combining ions` In order to combine two ions, you first arrange the positive ion followed by the negative ion, to then cross the ionic charges from the top of the ion to the bottom of the opposite ion. An example would be combining Mg^{2+} and N^{3-}, leading to Mg_3N_2. The number 3 next to Mg^{2+} comes from the charge of N, and the number 2 next to N comes from the charge of Mg. What if you need to combine more complex ions such as Mg^{2+} and OH^-. In this case the number 1 from OH^- will end up close to Mg and the number 2 from Mg would end up close to OH. As OH^- contains several atoms, then we need to use a parenthesis, and the final formula would be $Mg(OH)_2$, as the number 2 affects both the O and the H from OH.

Sample Problem 26

Combine the following ions or break down the following chemicals into ions:

Ions	Combination
Li^+ and $PO_4{}^{3-}$	
Ca^{2+} and $NO_3{}^-$	
	Li_2CO_3
	Mg_2NO_3

SOLUTION

By combining Li^+ and $PO_4{}^{3-}$ one obtains Li_3PO_4, and Ca^{2+} and $NO_3{}^-$ gives $Ca(NO_3)_2$. Li_2CO_3 results from Li^+ and $CO_3{}^{2-}$ and Mg_2NO_3 from combining Mg^{2+} with $NO_3{}^-$.

STUDY CHECK

Combine the following ions or give the ions given the final compound: Ca^{2+} and $NO_3{}^-$ and K_2SO_4.

Answer: Ca_2NO_3; K^+ and $SO_4{}^{2-}$.

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 3.1. For example, the name of $MgSO_4$ is magnesium sulfate, as Mg is magnesium and $SO_4{}^{2-}$ is called sulfate. Similarly, potassium nitrate results from the combination of K^+ potasium and $NO_3{}^-$, nitrate. By combining the two ions we obtain the final formula as $(K)_1(SO_4)_1$ that is KSO_4.

Sample Problem 27

Name of give the name of the following oxosalts:

Formula	Name
K_2SO_4	
Na_2CO_3	
	Nickel(II) carbonate
	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. Nickel(II) carbonate is $NiCO_3$ and sodium phosphate is Na_3PO_4.

STUDY CHECK

Name of give the name of the following oxosalts: $FeSO_4$ and Iron(III) sulfate.

Answer: Iron(II) sulfate and $Fe_2(SO_4)_3$.

Naming Hydrosalts Hydrosalts are very related to oxosalts. For example, sulfate ($SO_4{}^{2-}$) is an oxosalt whereas hyrogensulfate ($HSO_4{}^-$) is a hydrosalts. We obtain hydrosalts from oxosalts by adding hydrogen and for every hydrogen you add you have to subtract a negative charge for their total charge. For example, phosphate ($PO_4{}^{3-}$) is an oxosalt whereas hydrogenphosphate ($HPO_4{}^{2-}$) is an hydrosalt and dihydrogen-phosphate ($H_2PO_4{}^-$) is another hydrosalt. As phosphate has three negative charges, hydrogenphosphate has to have one less charge (that is 2−) and dihydrogenphosphate hast to have two less negative charges (that is −1).

Sample Problem 28

Name or formula the following hydrosalts:

Formula	Name
	Magnesium hydrogensulfate
	Sodium hydrogen carbonate
$LiHCO_3$	
MgH_2PO_4	

SOLUTION

The formula of Magnesium hydrogensulfate is $Mg(HSO_4)_2$ as the formula for hydrogen sulfate is $HSO_4{}^-$ and the valence of magnesium is Mg^{2+}. The formula for Sodium hydrogen carbonate is $NaHCO_3$ as it results from combining Na^+ and $HCO_3{}^-$. Mind hydrogencarbonate results from adding a hydrogen ion H^+ to a carbonate $CO_3{}^{2-}$ ion. The name for $LiHCO_3$ is lithium hydrogencarbonate, whereas the name for MgH_2PO_4 is magnesium dihydrogenphosphate.

STUDY CHECK

Name or formula the following hydrogensalts: $LiHS_2O_3$, LiH_2PO_4 and sodium hydrogenphosphate.

Answer: lithium hydrogenthiosulfate; Lithium dihydrogenphosphate and Na_2HPO_3.

Hydrates Some chemicals contain water molecules trapped in the solid structure of the chemical. By warming up the chemical we can let these water molecules go into the air. These types of chemicals that contain water are called *hydrates*. 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. In order to formulate hydrates you just need to use prefixes such as mono, di, tetra to indicate the number of water molecules in the chemical and end the name with *hydrate*.

Sample Problem 29

Name or formula the following hydrates:

Formula	Name
	Nickel(II) permanganate dihydrate
	Sodium nitrate monohydrate
$Na_2CO_3 \cdot 10\,H_2O$	
$MgSO_4 \cdot 7\,H_2O$	

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 formula the following hydrates: $LiNO_3 \cdot H_2O$, $Na_3PO_4 \cdot 3\,H_2O$ and sodium sulfate tetrahydrate.

Answer: lithium nitrate monohydrate; sodium phosphate trihydrate and $Na_2SO_4 \cdot 4\,H_2O$.

Common naming Some of the chemicals are normally referred by a common name that does not involve the use of any chemical naming rules. An example would be H_2O normally referred as water instead of its standard name that is dihydrogen oxide. You can find more names in Table 3.5.

Add this table into your flashcard.

Sample Problem 30

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: ammonia and methane.

Answer: NH_3 and CH_4.

Chemical	Name
H_2O	Water
NH_3	Ammonia
CH_4	Methane
CO_2	Dry ice
NaCl	Table salt
$NaHCO_3$	Sodium Bicarbonate
$Mg(OH)_2$	Milk of magnesia
N_2O	Laughing gas
$CaCO_3$	Marble
CaO	Quicklime
$NaHCO_3$	Baking Soda
$MgSO_4 \cdot 7\,H_2O$	Epsom Salt

Figure 3.6: List of common chemicals

CHAPTER 3

Ions & Ionic charges

1. Fe^{2+} is:

(a) an atom
(b) an element
(c) an anion
(d) a cation
(e) none of the above

Ans: (d)

2. F^- is:

(a) an atom
(b) an element
(c) an anion
(d) a cation
(e) none of the above

Ans: (c)

3. Cl is:

(a) an atom
(b) an ion
(c) an anion
(d) a cation
(e) none of the above

Ans: (a)

4. Identify the correct ionic state of Mg :

(a) Mg^{2+}
(b) Mg^{2-}
(c) Mg^+
(d) Mg^{2-}
(e) Mg

Ans: (a)

5. Identify the correct ionic state of O:

(a) O^{2+}
(b) O^{2-}
(c) O
(d) O^+
(e) O^-

Ans: (b)

6. Identify the correct ionic state of N:

(a) N^{2+}
(b) N^{2-}
(c) N^{3-}
(d) N
(e) N^-

Ans: (c)

Covalent Compounds

7. Name or formulate the following compound: NO

(a) nitrogen monoxide
(b) nitrogen(I) oxide
(c) nitrogen(II) oxide
(d) mononitrogen monoxide
(e) oxygen nitride

Ans: (a)

8. Name or formulate the following compound: phosphorus trichloride

(a) PCl_5
(b) PCl_3
(c) PCl_2
(d) P_3Cl
(e) PCl

Ans: (b)

9. Name or formulate the following compound: SO_3

(a) sulfur(III) oxide
(b) sulfur(VI) oxide
(c) sulfur trioxide
(d) trisulfur monoxide
(e) trisulfur oxide

Ans: (c)

10. Name or formulate the following compound: sulfur hexafluoride

(a) SF_6
(b) S_6F
(c) SF
(d) SF_2
(e) S_6F_3

Ans: (a)

Ionic Compounds

11. Name or formulate the following compound: Li_3N

(a) Trilithium nitride
(b) Lithium(I) nitride
(c) Lithium nitride
(d) Lithium Dinitride
(e) Lithium trinitride

Ans: (c)

12. Name or formulate the following compound: $MgCl_2$

(a) Magnesium(II) chloride
(b) Magnesium chloride(II)
(c) Magnesium Dichloride
(d) Dimagnesium chloride
(e) Magnesium chloride

Ans: (e)

13. Name or formulate the following compound: calcium sulfide

(a) Ca_2S_2
(b) CaS_2
(c) Ca_2S
(d) CaS
(e) Ca_2S_1

Ans: (d)

14. Combine the following ions: Ba^{2+} and P^{3-}

(a) BaP
(b) Ba_2P_3
(c) Ba_3P_2
(d) BaP_2
(e) Ba_3P_3

Ans: (c)

15. Name or formulate the following compound: VO

(a) Vanadium monoxide
(b) Vanadium(I) oxide
(c) Vanadium(II) oxide
(d) Vanadium oxide

Ans: (c)

16. Name or formulate the following compound: Ni_2O_3

(a) Nickel(II) oxide
(b) Nickel(I) oxide
(c) Nickel oxide
(d) Nickel monoxide

Ans: (a)

17. Name or formulate the following compound: Iron(III) chloride

(a) $FeCl_3$
(b) $FeCl_2$
(c) Fe_3Cl
(d) Fe_2Cl_3
(e) FeCl

Ans: (a)

Acids and Bases Naming

18. Name or formulate the following compound: hydrochloric Acid
   (a) HI
   (b) $HClO_4$
   (c) HCl
   (d) $HClO_3$
   (e) ClH

Ans: (c)

19. Name or formulate the following compound: hydroiodic Acid
   (a) HI
   (b) $HClO_4$
   (c) HCl
   (d) $HClO_3$
   (e) ClH

Ans: (a)

20. Name or formulate the following compound: perchloric acid
   (a) $HMnO_4$
   (b) H_2CO_3
   (c) HNO_3
   (d) H_2SO_4
   (e) $HClO_4$

Ans: (e)

21. Name or formulate the following compound: nitric acid
   (a) $HMnO_4$
   (b) H_2CO_3
   (c) HNO_3
   (d) H_2SO_4
   (e) $HClO_4$

Ans: (c)

22. Name or formulate the following compound: permanganic acid
   (a) $HMnO_4$
   (b) H_2CO_3
   (c) HNO_3
   (d) H_2SO_4
   (e) $HClO_4$

Ans: (a)

23. Name or formulate the following compound: sulfuric acid
   (a) $HMnO_4$
   (b) H_2CO_3
   (c) HNO_3
   (d) H_2SO_4
   (e) $HClO_4$

Ans: (d)

24. Name or formulate the following compound: $Ca(OH)_2$
   (a) calcium hydroxyde
   (b) calcium(II) hydroxyde
   (c) calcium(I) hydroxyde
   (d) Calcium dihydroxide
   (e) Calcium hydrate

Ans: (a)

Naming of oxosalts, hydrosalts, hydrates & common chemicals

25. Combine the chemical: $SO_4{}^{2-}$ and Ca^{2+}
   (a) $CaSO_4$
   (b) $Ca_2(SO_4)_2$
   (c) Ca_2SO_4
   (d) $CaSO_3$
   (e) CaS_4

Ans: (a)

26. Give the ions forming the chemical: KNO_3
   (a) $NO_4{}^-$ and K^+
   (b) $NO_3{}^{-2}$ and K^+
   (c) $NO_3{}^{-2}$ and K^{+2}
   (d) $NO_3{}^-$ and K^+
   (e) $NO_3{}^-$ and K^{2+}

Ans: (d)

27. Name or formulate the following compound: Na_2CO_3
   (a) Sodium bicarbonate
   (b) Sodium(I) carbonate
   (c) Sodium(II) carbonate
   (d) Sodium carbonate
   (e) Sodium carbon oxide

Ans: (d)

28. Name or formulate the following compound: $FeCO_3$
   (a) Iron(II) carbonate
   (b) Iron(I) carbonate
   (c) Iron carbonate
   (d) Iron carbon oxide
   (e) Iron(III) carbonate

Ans: (a)

29. Name or formulate the following compound: potassium permanganate
   (a) KMnO
   (b) K_2Mn
   (c) K_2MnO_4
   (d) $KMnO_3$
   (e) $KMnO_4$

Ans: (e)

30. Name or formulate the following compound: sodium bicarbonate(common name)
   (a) $NaCO_4$
   (b) $NaHCO_3$
   (c) $NaCO_3$
   (d) Na_2CO_3
   (e) $NaCO_2$

Ans: (b)

31. Name or formulate the following compound: ammonia(common name)
   (a) CH_4
   (b) NH_3
   (c) $NaCO_3$
   (d) $Mg(OH)_2$
   (e) NaCl

Ans: (b)

32. Name or formulate the following compound:
   (a) $MgSO_4$
   (b) $Ni(SO_4)_3$
   (c) Cobalt(II) nitrate
   (d) Cobalt(II) sulfate dihydrate
   (e) $KHCO_3$

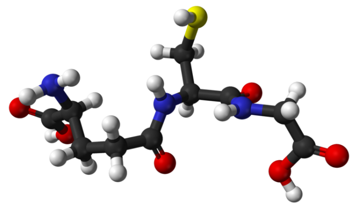
Ans: magnesium sulfate, nicke(III) sulfate, $Co(NO_3)_2$, $CoSO_4 \cdot 2\,H_2O$, potassium hydrogencarbonate.

33. Name or formulate the following compound:
   (a) $Ca(NO_3)_2$
   (b) $Na(HCO_3)_2$
   (c) Nickel(II) sulfate
   (d) Nicke(II) sulfate tetrahydrate
   (e) NaH_2PO_4

Ans: calcium nitrate, sodium hydrogen carbonate, $NiSO_4$, $NiSO_4 \cdot 4\,H_2O$, sodium dihydrogenphosphate.

4

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 or by the gram. This chapter will introduce the idea of a 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 oder to predict the amount of chemicals produced.

GOALS

1. Transform grams into moles and moles into molecules
2. Balance chemical reactions
3. Carry stoichiometric calculations
4. Identify the limiting reagent
5. Calculate the % yield

4.1 The mole

Some of the terms you use in your everyday life actually refer to a number. 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 beers–24 cans.
In a chemistry laboratory we normally do not weight small numbers of molecules of a chemical. In chemistry, molecules are counted by the mole, and the term mole–abbreviated as mol–refers to 6.022×10^{23}. 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, in reference to Amedeo Avogadro, the Italian physicists who coined the term.
In chemistry labs, chemicals are often measured by weight. In this section we will show how to convert moles into weights–into grams–by using a property called the 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 mol refers to a large number of molecules.

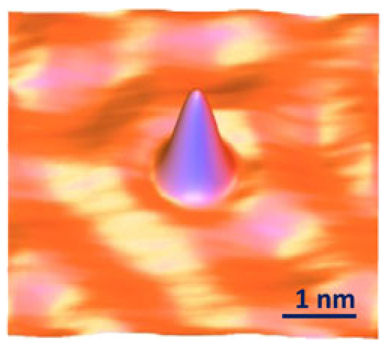


Figure 4.1: Molecules are counted by the mole and only in rare occasions we are able to count molecules one by one

`From moles to atoms` 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 follow unit equivalency:

$$\boxed{1\text{mol of H}_2\text{O} = 6.02 \times 10^{23} molecules of \text{H}_2\text{O}}$$

Add this relation into your flashcard.

or a conversion factor to transform moles into molecules or molecules into moles as well:

$$\boxed{\frac{1\text{mol of H}_2\text{O}}{6.02 \times 10^{23}\text{molecules of H}_2\text{O}} \text{ or } \frac{6.02 \times 10^{23}\text{molecules of H}_2\text{O}}{1\text{mol of H}_2\text{O}}}$$

In the same way, a mole contains 6.02×10^{23} molecules, as the term mole refers to the number 6.02×10^{23}. For example, in order to calculate the number of molecules of

Figure 4.2: Eggs are bought by the dozen and molecules are counted by the mol

H_2O in 3 moles of H_2O, (moles $\rightarrow$ molecules) you need to set up a conversion factor, starting by the give information (3 moles) and using the mol-to-molecule conversion factor with mol on the bottom:

$$3\,\cancel{\text{moles of } H_2O} \times \frac{6.02 \times 10^{23} \text{ molecules of } H_2O}{1\,\cancel{\text{mol of } H_2O}} = 1.80 \times 10^{24} \text{molecules of } H_2O$$

In 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 molecule in the bottom. For example 3×10^{20} H_2O molecules equals to 4.98×10^{-4}moles of H_2O as

$$3 \times 10^{20}\,\cancel{\text{molecules of } H_2O} \times \frac{1\text{moles of } H_2O}{6.02 \times 10^{23}\,\cancel{\text{molecules of } H_2O}} = 4.98 \times 10^{-4}\text{moles of } H_2O$$

From molecules to atoms Molecules are made of atoms, and for example the CO_2 molecule contains an atom of C and two atoms of O. In order to convert from molecules to atoms (molecules $\rightarrow$ atoms) you need to use the coefficients in the molecular formula. For example, and 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:

$$\boxed{\frac{1 \text{ molecule of } H_2O}{1 \text{ atom of O}} \text{ and } \frac{1 \text{ molecule of } H_2O}{2 \text{ atom of H}}}$$

Figure 4.3: Socks are bought as pairs and molecules are counted by the mol

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 1.204×10^{22} H atoms, as:

$$3\,\cancel{\text{moles of } H_2O} \times \frac{6.022 \times 10^{23}\,\cancel{\text{molecules of } H_2O}}{1\,\cancel{\text{moles of } H_2O}} \times \frac{2\text{H atoms}}{1\,\cancel{\text{molecules of } H_2O}}$$
$$= 1.204 \times 10^{22}\text{atoms of H}$$

Sample Problem 31

Calculate: (a) the number of number of CuO molecules in 3.4 moles of CuO; (b) the number of moles of CO in 5×10^{20} CO molecules; (c) The number of O atoms in 4.5 moles of NO_2.

SOLUTION

(a) 3.4 moles of CuO equals to 2.05×10^{22} molecules of CuO as:

$$3.4\,\cancel{\text{moles of CuO}} \times \frac{6.022 \times 10^{23} \text{ molecules of CuO}}{1\,\cancel{\text{mole of CuO}}} = 2.05 \times 10^{22}\text{moles of CuO}$$

(b) 5×10^{20} CO molecules equals to 8.3×10^{-4} moles of CO , as

$$5 \times 10^{20}\,\cancel{\text{CO molecules}} \times \frac{1\,\cancel{\text{mole of CO}}}{6.022 \times 10^{23} \text{ CO molecules}} = 8.3 \times 10^{-4}\text{moles of CO}$$

Figure 4.4: You can buy a six-pack of beers

(c) 4.5 moles of NO_2 contains 5.4×10^{24} O atoms, as

$$4.5\,\cancel{\text{moles of } NO_2} \times \frac{6.022 \times 10^{23}\,\cancel{NO_2 \text{ molecules}}}{1\,\cancel{\text{moles of } NO_2}} \times \frac{2\text{O atoms}}{1\,\cancel{NO_2 \text{ molecules}}}$$
$$= 5.4 \times 10^{24}\text{O atoms}$$

STUDY CHECK

The chemical formula for caffeine is $C_8H_{10}N_4O_2$. Calculate the number of C, H, N and O atoms in 3.5 moles of caffeine.

Answer: 3.4×10^{25} moles of C, 4.2×10^{25} moles of H, 1.7×10^{25} moles of N and 8.4×10^{24} moles of O.

4.2 The molecular mass

A standard way to measure chemicals in the lab is by weight. We can weight different quantities and the larger the quantity the larger the weight. For a chemical, the weigh of a mole is called the molar (or molecular) weight. For example if we weight a mole of water (H_2O) we will be weighting 18 grams of water, or if you weight 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).

Figure 4.5: The molecular mass of cinnamic acid ($C_9H_8O_2$), used in the manufacture of flavors, is $148.16\frac{g}{mol}$

Molar mass of a chemical Chemicals are made of atoms, and each atom has an 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 H atoms. The units for molecular weight is $\frac{g}{mol}$, also written as $g \cdot mol^{-1}$. In order 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 $C_2H_4O_2$ that means a vinegar molecule contains 2C, 4H and 2O atoms. If you add the atomic masses of 2C, 4H and 2O you will get $60g \cdot mol^{-1}$. 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 $164.09g \cdot mol^{-1}$.

Figure 4.6: Car batteries contains sulfuric acid (H_2SO_4), a corrosive chemical with a molar mass of $180.06\frac{g}{mol}$

Sample Problem 32

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 $24.31g \cdot mol^{-1}$. (b) The molar mass of H_2SO_4 is the result of adding the atomic masses of 2H (AW=$1g \cdot mol^{-1}$) atoms, 1 S (AW=$32g \cdot mol^{-1}$) and 4O (AW=$16g \cdot mol^{-1}$) atoms, that gives $98.08g \cdot mol^{-1}$.

STUDY CHECK

Calculate the molar mass of glucose $C_6H_{12}O_6$

Answer: $180.06g \cdot mol^{-1}$.

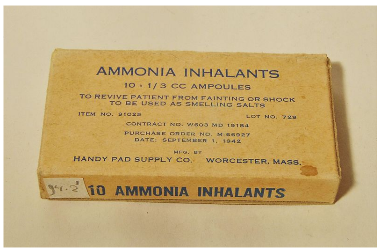


Figure 4.7: Ammonia smelling salts ($(NH_4)_2CO_3$, MW=$96g \cdot mol^{-1}$) were historically employed to wake up injured athlete during a sport game.

`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 $18g \cdot mol^{-1}$. This means:

$$\boxed{\text{1mole of } H_2O = \text{18g of } H_2O}$$

that is the same as

$$\boxed{\frac{\text{1 mole of } H_2O}{\text{18 grams of } H_2O} \text{ or } \frac{\text{18 grams of } H_2O}{\text{1 mole of } H_2O}}$$

Sample Problem 33

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: $2 \times 5 + 8 \times 1 + 1 \times 12 + 3 \times 16 = 96g \cdot mol^{-1}$. In order to calculate the moles given the gram, you need to use the molar mass a conversion factor:

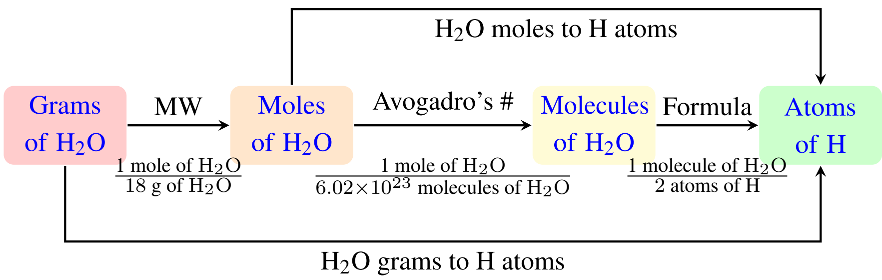
$$100\,\cancel{\text{g of } (NH_4)_2CO_3} \times \frac{\text{moles of } (NH_4)_2CO_3}{96\,\cancel{\text{g of } (NH_4)_2CO_3}} = 1.04 \text{ moles of } (NH_4)_2CO_3$$

STUDY CHECK

Calculate the MW of table salt (NaCl) and the grams in 20 moles of this salt.

Answer: $58.4g \cdot mol^{-1}$; 1168g.

`From grams to atoms` In the previous sections we covered how to convert grams to moles, and moles to molecules, or molecules to atoms. You can follow the diagram below in order 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.

Sample Problem 34

Convert 10 grams of ammonia (NH_3, MW=$17 g \cdot mol^{-1}$) 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 into atoms:

$$10\cancel{\text{g of NH}_3} \times \frac{1 \text{ mole of NH}_3}{17 \cancel{\text{ g of NH}_3}} \times \frac{6.022 \times 10^{23} \text{ NH}_3 \text{ molecules}}{1 \cancel{\text{ mole of NH}_3}} \times \frac{3 \text{ H atoms}}{1 \text{ NH}_3 \cancel{\text{ molecule}}} = 1.8 \times 10^{24} \text{H atoms}$$

STUDY CHECK

Methane is a chemical used as a fuel. Calculate how many grams of methane CH_4 contains 5×10^{25} H atoms.

Answer: 332.1g.

4.3 Chemical reactions

When we eat we burn food with oxygen 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. 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 to produce fertilizers. This section covers the basic of chemical reactions. You will learn how to balance reaction and how classify reaction in different types.

Figure 4.8: A termite reaction between iron(III) oxide and Al: $Fe_2O_3 + 2\,Al \longrightarrow 2\,Fe + Al_2O_3$

`Simple chemical reactions` Magnesium is a metal that react 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 together–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 after. 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` Chemicals reactions can be read in words. In order 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. In order 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,

Figure 4.9: An image of the combustion of Mg: $2\,Mg(s) + O_2(g) \longrightarrow 2\,MgO(s)$

$$2\,Mg(s) + O_2(g) \longrightarrow 2\,MgO(s).$$

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 to two. On the other hand, the number of O atoms in the reactants and products is the same, being 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 to 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 In order to balance a reaction, we need to introduce the stoichiometric coeficientes that make the number of atoms of reactants and products the same. In order 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 35

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: $Fe_2O_3(s) + C(s) \longrightarrow Fe(s) + CO(g)$

Answer: $Fe_2O_3(s) + 3\,C(s) \longrightarrow 2\,Fe(s) + 3\,CO(g)$.

Figure 4.10: Iron rust is the result of a combination reaction: $4\,Fe_2{}^{+}O_2 \longrightarrow 2\,Fe_2O_3$

`Five types of reactions` Most of the 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:

$$2\,Mg(s) + O_2(g) \longrightarrow 2\,MgO(s)\ \text{(combination)}$$

Figure 4.11: Wood burning is a combustion reaction

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:

$$\mathrm{Zn(s) + 2\,HCl(aq) \longrightarrow \mathit{Zn}Cl_2(aq) + H_2(g)}\ \text{(Single replacement)}$$

In a double replacement reaction, the first element in the reacting compounds switch places. An example is the reaction between $AgNO_3$ and NaCl, in which Ag from $AgNO_3$ replaces Na in NaCl:

$$\mathit{Ag}\mathrm{NO_3(aq)} + \mathit{Na}\mathrm{Cl(aq)} \longrightarrow \mathit{Na}\mathrm{NO_3(aq)} + \mathit{Ag}\mathrm{Cl(s)}\ \text{(Double replacement)}$$

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):

$$\mathrm{CH_4(g) + O_2(g) \longrightarrow H_2O(g) + CO_2(g)}\ \text{(Combustion)}$$

4.4 Mass calculations

In the previous section, you learned how to balance a chemical reaction. In order 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:

$$\mathrm{2\,Mg(s) + O_2(g) \longrightarrow 2\,MgO(s)}$$

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` 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) is:

$$\boxed{\frac{\text{2 moles of Mg}}{\text{1 moles of O}_2} \text{ or } \frac{\text{1 moles of O}_2}{\text{2 moles of Mg}}}$$

Similarly, the mole ratio between Mg and MgO is:

$$\boxed{\frac{\text{2 moles of Mg}}{\text{2 moles of MgO}} \text{ or } \frac{\text{2 moles of MgO}}{\text{2 moles of Mg}}}$$

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 product.

`Reactants to products` We will calculate how much MgO will be produced from 5 moles of Mg by converting Mg into MgO by means of the conversion factor between both chemicals. As we want to transform the Mg into MgO we will use the conversion factor with Mg on the bottom of the fraction. This way the units will cancel out to give moles of MgO:

$$5\ \cancel{\text{moles of Mg}} \times \frac{2 \text{ moles of MgO}}{2\ \cancel{\text{moles of Mg}}} = 5 \text{ moles of MgO.}$$

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 do we need to react with Mg, we will proceed as:

$$5\ \cancel{\text{moles of Mg}} \times \frac{1 \text{ mole of } O_2}{2\ \cancel{\text{moles of Mg}}} = 2.5 \text{ moles of } O_2.$$

This result means that 2.5 moles of O_2 will react with 5 moles of Mg.

Sample Problem 36

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

Calculate the number of moles of water produced by 4 moles of oxygen.

Answer: 8 moles.

`Discussion:` What weights more one kilo of Sulfur or one kilo of Gold? Now, what weights more, one mol of Sulfur or one mole of Gold

4.5 Percent yield and limiting reagent

Reactions sometimes do not fully complete. So if one expect to obtain a given quantity of chemical at the end of a long experiment, chances are one really obtains less than expected. The percent yield tells how much of the final chemical do you really produce. The larger this value the more chemical the reaction produces and the less waste it generates. At the same time when mixing two reactants sometimes one of them remains in the form of leftover. In another words, sometimes of the reagents limits the reaction and the other one is simply in excess. This section will cover the very important ideas of percent yield and limiting reagent.

Percent yield Chemical reactions are not perfect and most of the time all staring material does convert into products. The reasons are that when we perform real chemistry, and for example measure quantities, transfer reactants from one container into another, some material is lost along the way. Also some chemicals are not pure and when you weight is more than the pure chemical. Besides that, there are often times, side reactions that compete with the main reaction generating byproducts. The theoretical yield is the amount of products one will expect in a hypothetically perfect chemical reaction, while the actual yield is the amount of product that is indeed produced in reality. The percent yield is just the fraction between the actual yield and the theoretical yield in percent form:

$$\text{Percent Yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

In order to compute the percent yield, you need to calculate the theoretical yield, that is the amount of product that will be produced following the rules of stoichiometry, whereas the actual yield has to be give, as this quantity is measured in the lab.

Sample Problem 37

Ammonia is produced by reacting nitrogen and 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 is smaller than the theoretical. 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

Calculate the percent yield of ammonia if 1 moles of nitrogen gives 1.5 moles of ammonia

Answer: 75%.

Limiting reagent Let us consider the reaction between carbon and 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 oxygen to produce one mole of carbon dioxide". Mind that to produce one mole of CO_2 you need one more of C and one mole of 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. 0.5 moles of carbon dioxide will be formed and 0.5 moles of C will remain. Differently, no oxygen will remain and the whole 0.5 moles will consume. C is the excess reactant and oxygen the limiting reactant. Often times a reaction will be presented with the quantities of two reactants and you will have to determine the limiting reagent. The limiting 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 calculate the limiting reagent.// Consider the following reaction, in which 3 moles of Mg react with 2 moles of oxygen to produce magnesium oxide:

$$\underset{\text{3 moles}}{2\,Mg} + \underset{\text{2 moles}}{O_2} \longrightarrow 2\,MgO$$

In order to identify the limiting reagent, we will chose one of the given reagent quantities and calculate the moles needed of the other reagent. For example, if we choose the 3 moles of Mg, the amount of oxygen needed to react with this quantity will be:

$$3\cancel{\text{ moles of Mg}} \times \frac{1 \text{ moles of } O_2}{2\cancel{\text{ moles of Mg}}} = 1.5 \text{ moles of } O_2.$$

This means that to react with 3 moles of Mg you need 1.5 moles of oxygen. On the other hand, you have two moles of oxygen, and that is more than what you need to react with the 3 moles of Mg. Hence, oxygen is the excess reagent and Mg the limiting reagent.

Sample Problem 38

In the synthesis of ammonia (NH_3):

$$\underset{\text{2 moles}}{3\,H_2(g)} + \underset{\text{2 moles}}{N_2(g)} \longrightarrow 2\,NH_3(g)$$

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

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.

Answer: N_2 is the limiting reagent and H_2 the excess. 0.5 moles of hydrogen will remain.

4.6 Percent yield and limiting reagent

In the previous sections, given the moles of reactants, we learn how to use chemical reactions to predict the amount of product formed, or calculate the energy involved in a reaction. 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 in given grams. We will based the following examples in the reaction of Mg and O_2:

$$2\,\mathrm{Mg(s)} + \mathrm{O_2(g)} \longrightarrow 2\,\mathrm{MgO(s)}$$

`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 $18g \cdot mol^{-1}$. If we need to convert 12 grams of water into mols we should do:

$$12\cancel{\text{g of H}_2\text{O}} \times \frac{\cancel{\text{moles of H}_2\text{O}}}{\cancel{18\text{ g of H}_2\text{O}}} = 0.66 \text{ g}$$

In order 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 in order to operate with them you can only use moles.

`Grams to moles in a reaction` Consider the reaction between Mg and O_2 to produce MgO. We need to calculate the grams of MgO produced from 3 moles of Mg. We will proceed first starting with the moles of Mg and using the conversion factor that related moles of Mg and moles of MgO, with 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=$24g \cdot mol^{-1}$). This times, we will have first to convert the grams of Mg into moles to then use the mole ration between Mg and MgO:

$$5\cancel{\text{ g of Mg}} \times \frac{1 \text{ moles of Mg}}{24\cancel{\text{ g of Mg}}} \times \frac{2 \text{ moles of MgO}}{2\cancel{\text{ moles of Mg}}} = 0.21 \text{ g of MgO}.$$

`Grams to grams in a reaction` Imagine now, we start with 6 moles of Mg (AW=$24g \cdot mol^{-1}$) and we want to know the mass of MgO(MW=$40g \cdot mol^{-1}$) produced. The importance of this type of calculations is based on the fact that in a chemistry laboratory chemicals produced or employed are normally weighted in grams. For 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. Aft this point, we will finish the calculation by converting the moles of MgO into grams using its molecular weight:

$$5\ \cancel{\text{grams of Mg}} \times \frac{1\ \cancel{\text{moles of Mg}}}{24\ \cancel{\text{grams of Mg}}} \times \frac{2\ \cancel{\text{moles of MgO}}}{2\ \cancel{\text{moles of Mg}}} \times \frac{40 \text{ grams of MgO}}{1\ \cancel{\text{moles of MgO}}} = 10 \text{ grams of MgO}.$$

Sample Problem 39

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=$2g \cdot mol^{-1}$); (b) Calculate the number of grams of NH_3 (MW=$17g \cdot mol^{-1}$) produced from 10 grams of nitrogen (MW=$28g \cdot mol^{-1}$)

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:

$$10\ \cancel{\text{grams of H}_2} \times \frac{1\ \cancel{\text{moles of H}_2}}{2\ \cancel{\text{grams of H}_2}} \times \frac{2 \text{ moles of NH}_3}{3\ \cancel{\text{moles of H}_2}} = 3.33 \text{ moles of NH}_3.$$

that is: 10 grams of hydrogen produce 3.33 moles of ammonia.(a) 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

Calculate the number of grans of nitrogen needed to react with 3 grams of hydrogen to produce ammonia.

Answer: 14 grams.

CHAPTER 4

The Mole

1. Calculate the number of molecules in 8 moles of CO:

(a) 2.4×10^{24}
(b) 4.8×10^{24}
(c) 6.0×10^{24}
(d) 9×10^{24}
(e) 6×10^{23}

Ans: (b)

2. How many moles equal to 3.2×10^{21} molecules:

(a) 4983.39 moles
(b) 0.33 moles
(c) 5.3×10^{-3} moles
(d) 8.3×10^{-9}moles
(e) 0.99 moles

Ans: (c)

3. How many moles equal to 2×10^{23} molecules:

(a) 4983.39 moles
(b) 0.33 moles
(c) 5.3×10^{-3} moles
(d) 8.3×10^{-9}moles
(e) 0.99 moles

Ans: (b)

4. Calculate the number of H atoms in 1×10^{-20} molecues of CH_4:

(a) 1×10^{-20} atoms
(b) 2×10^{-20} atoms
(c) 4×10^{-20} atoms
(d) 1.3×10^{25} atoms
(e) 0.23 moles

Ans: (c)

5. Calculate the number of C atoms in 3 moles of $C_{10}H_{14}N_2$:

(a) 1×10^{-20} atoms
(b) 3.6×10^{-22} atoms
(c) 2.5×10^{-25} atoms
(d) 1.8×10^{25} atoms
(e) 0.99 moles

Ans: (d)

6. Calculate the number of H atoms in 3 moles of $C_{10}H_{14}N_2$:

(a) 1×10^{-20} atoms
(b) 3.6×10^{-22} atoms
(c) 2.5×10^{-25} atoms
(d) 1.8×10^{-25} atoms
(e) 0.99 moles

Ans: (c)

The Molecular Mass

7. Compute the molar mass of NH_3:

(a) 17 $g \cdot mol^{-1}$
(b) 32 $g \cdot mol^{-1}$
(c) 28 $g \cdot mol^{-1}$
(d) 291.71 $g \cdot mol^{-1}$
(e) 2 $g \cdot mol^{-1}$

Ans: (a)

8. Compute the molar mass of CO:

(a) 17 $g \cdot mol^{-1}$
(b) 32 $g \cdot mol^{-1}$
(c) 28 $g \cdot mol^{-1}$
(d) 291.71 $g \cdot mol^{-1}$
(e) 2 $g \cdot mol^{-1}$

Ans: (c)

9. Compute the molar mass of $Fe_2(CO_3)_3$:

(a) 17 $g \cdot mol^{-1}$
(b) 32 $g \cdot mol^{-1}$
(c) 116 $g \cdot mol^{-1}$
(d) 291.71 $g \cdot mol^{-1}$
(e) 2 $g \cdot mol^{-1}$

Ans: (d)

10. How many grams are there in 3 moles of Silver (AW=$107.9 g \cdot mol^{-1}$):

(a) 117.3 g
(b) 176.8 g
(c) 3 g
(d) 323.7 g
(e) 156 g

Ans: (d)

11. How many grams are there in 3 moles of Potassium (AW=$39.10 g \cdot mol^{-1}$):

(a) 117.3 g
(b) 176.8 g
(c) 3 g
(d) 323.7 g
(e) 156 g

Ans: (a)

12. How many grams are there in 3 moles of Cromium (AW=$52 g \cdot mol^{-1}$):

(a) 117.3 g
(b) 176.8 g
(c) 3 g
(d) 323.7 g
(e) 156 g

Ans: (e)

13. How many grams are there in 3 moles of atomic hydrogen (AW=$1 g \cdot mol^{-1}$):

(a) 117.3 g
(b) 176.8 g
(c) 3 g
(d) 323.7 g
(e) 156 g

Ans: (c)

14. How many grams are there in 4 moles of $C_6H_{12}O_6$:

(a) 100.10 grams
(b) 834.02 grams
(c) 720.64 grams
(d) 602.50 grams
(e) 55.20 grams

Ans: (c)

15. How many C atoms are there in 3 moles of $C_6H_{12}O_6$:

(a) 2×10^{27} atoms
(b) 2.13×10^{26} atoms
(c) 1.08×10^{25} atoms
(d) 2.17×10^{25} atoms
(e) 5×10^{25} atoms

Ans: (a)

16. How many O atoms are there in 3 moles of $C_6H_{12}O_6$:

(a) 2×10^{27} atoms
(b) 2.13×10^{26} atoms
(c) 1.08×10^{25} atoms
(d) 3.9×10^{27} atoms
(e) 5×10^{25} atoms

Ans: (d)

17. Fill the conversion factor that calculates the final property:

$$4\ \cancel{\text{moles of } CO_2} \times \frac{\underline{\hspace{2cm}}\ \text{g of } CO_2}{\underline{\hspace{2cm}}\ \cancel{\text{moles of } CO_2}} = \underline{\hspace{2cm}}\ \text{g of } CO_2.$$

18. Fill the conversion factor that calculates the final property:

$$10\ \cancel{\text{g of NO}} \times \frac{\underline{\hspace{2cm}}\ \text{moles of NO}}{\underline{\hspace{2cm}}\ \cancel{\text{g of NO}}} = \underline{\hspace{2cm}}\ \text{moles of NO}.$$

19. Fill the conversion factor that calculates the final property:

$$5\ \cancel{\text{moles of } C_6H_{12}O_6} \times \frac{\underline{\hspace{2cm}}}{\underline{\hspace{2cm}}} = \underline{\hspace{2cm}}\ \text{g of } C_6H_{12}O_6.$$

20. Fill the conversion factor that calculates the final property:

$$7\ \cancel{\text{g of } CH_4N_2O} \times \frac{\underline{\hspace{2cm}}}{\underline{\hspace{2cm}}} = \underline{\hspace{2cm}}\ \text{moles of } CH_4N_2O.$$

21. Fill the conversion factor that calculates the final property:

$$10^{24}\ \cancel{\text{molecules of } NO_2} \times \frac{\underline{\hspace{2cm}}\ \text{moles of } NO_2}{\underline{\hspace{2cm}}\ \cancel{\text{molecules of } NO_2}} = \underline{\hspace{2cm}}\ \text{moles of } NO_2.$$

22. Fill the conversion factor that calculates the final property:

$$3\ \cancel{\text{moles of NO}} \times \frac{\underline{\hspace{2cm}}\ \text{molecules of NO}}{\underline{\hspace{2cm}}\ \cancel{\text{moles of NO}}} = \underline{\hspace{2cm}}\ \text{molecules of NO}.$$

23. Fill the conversion factor that calculates the final property:

$$6\ \cancel{\text{moles of } C_6H_{12}O_6} \times \frac{\underline{\hspace{2cm}}}{\underline{\hspace{2cm}}} = \underline{\hspace{2cm}}\ \text{molecules of } C_6H_{12}O_6.$$

24. Fill the conversion factor that calculates the final property:

$$10^{25}\ \cancel{\text{molecules of } CH_4N_2O} \times \frac{\underline{\hspace{2cm}}}{\underline{\hspace{2cm}}} = \underline{\hspace{2cm}}\ \text{moles of } CH_4N_2O.$$

25. Fill the conversion factor that calculates the final property:

$$10^{26}\ \cancel{\text{molecules of } NO_2} \times \frac{\underline{\hspace{2cm}}\ \text{atoms of O}}{\underline{\hspace{2cm}}\ \cancel{\text{molecules of } NO_2}} = \underline{\hspace{2cm}}\ \text{atoms of O}.$$

26. Fill the conversion factor that calculates the final property:

$$10^{22}\ \cancel{\text{atoms of O}} \times \frac{\underline{\hspace{2cm}}\ \text{molecules of } H_2O}{\underline{\hspace{2cm}}\ \cancel{\text{atoms of O}}} = \underline{\hspace{2cm}}\ \text{molecules of } H_2O.$$

27. Fill the conversion factor that calculates the final property:

$$6\ \cancel{\text{molecules of } C_6H_{12}O_6} \times \frac{\underline{\hspace{2cm}}}{\underline{\hspace{2cm}}} = \underline{\hspace{2cm}}\ \text{atoms of C}.$$

28. Fill the conversion factor that calculates the final property:

$$10^{21}\ \cancel{\text{atoms of N}} \times \frac{\underline{\hspace{2cm}}}{\underline{\hspace{2cm}}} = \underline{\hspace{2cm}}\ \text{molecules of } CH_4N_2O.$$

Chemical Reactions

29. Indicate the stoichiometric coefficients that balance next reaction:

$$P_4(s) + O_2(g) \longrightarrow P_4O_{10}(s)$$

(a) 1,1,1
(b) 1,1,10
(c) 1,1,5
(d) 1,5,1
(e) 1,2,5

Ans: (d)

30. Balance the following reaction:

$$Al(s) + O_2(g) \longrightarrow Al_2O_3(s)$$

Ans: 4,3,2

31. Balance the following reaction:

$$FeS(s) + O_2(g) \longrightarrow Fe_2O_3(s) + SO_2(g)$$

Ans: 4,7,2,4

32. Balance the following reaction:

$$NH_3(g) + O_2(g) \longrightarrow NO(g) + H_2O(g)$$

Ans: 4,5,4,6

Mole-Mole Relationships

33. Fill the mole ratio for the following reaction:

$$C_6H_{12}O_6(s) + 6\,O_2(g) \longrightarrow 6\,CO_2(g) + 6\,H_2O(g)$$

$$\frac{\underline{\hspace{2cm}}\ \text{moles of } C_6H_{12}O_6}{\underline{\hspace{2cm}}\ \text{moles of } O_2}$$

Ans: 1/6

34. Fill the mole ratio for the following reaction:

$$C_6H_{12}O_6(s) + 6\,O_2(g) \longrightarrow 6\,CO_2(g) + 6\,H_2O(g)$$

$$\frac{\underline{\hspace{2cm}}\ \text{moles of } O_2}{\underline{\hspace{2cm}}\ \text{moles of } CO_2}$$

Ans: 6/6

35. Fill the conversion factor that calculates the moles of oxygen needed to react with 2 moles of Silver producing AgO:

$$2\,Ag(s) + O_2(g) \longrightarrow 2\,AgO(s)$$

$$2\ \cancel{\text{moles of Ag}} \times \frac{\underline{\hspace{2cm}}\ \text{moles of } O_2}{\underline{\hspace{2cm}}\ \cancel{\text{moles of Ag}}} = 1\ \text{moles of } O_2.$$

Ans: 1/2

36. Fill the conversion factor that calculates the moles of AgO produced from 2 moles of Silver:

$$2\,\mathrm{Ag(s)} + \mathrm{O_2(g)} \longrightarrow 2\,\mathrm{AgO(s)}$$

$$2\,\cancel{\text{moles of Ag}} \times \frac{\underline{\qquad}\ \text{moles of AgO}}{\underline{\qquad}\ \cancel{\text{moles of Ag}}} = 2 \text{ moles of AgO.}$$

Ans: 2/2

37. Fill the conversion factor that calculates the moles of AgO produced from 10 moles of oxygen:

$$2\,\mathrm{Ag(s)} + \mathrm{O_2(g)} \longrightarrow 2\,\mathrm{AgO(s)}$$

$$10\,\cancel{\text{moles of O}_2} \times \frac{\underline{\qquad}\ \text{moles of AgO}}{\underline{\qquad}\ \cancel{\text{moles of O}_2}} = 20 \text{ moles of AgO.}$$

Ans: 20

38. Calculate how many moles of nitrogen are needed to react with 5 moles of hydrogen, to produce ammonia:

$$\underset{\text{5 moles}}{3\,\mathrm{H_2(g)}} + \underset{\underline{\qquad}\ \text{moles}}{\mathrm{N_2(g)}} \longrightarrow 2\,\mathrm{NH_3(g)}$$

Ans: 3.3 moles

PERCENT YIELD AND LIMITING REAGENT

39. Six moles of nitrogen gas react to produce three moles of ammonia according to the following reaction:

$$3\,\mathrm{H_2(g)} + \underset{\text{6 moles}}{\mathrm{N_2(g)}} \longrightarrow \underset{\text{3 moles}}{2\,\mathrm{NH_3(g)}}$$

Calculate the percent yield.

Ans: 25%.

40. Six moles of nitrogen gas react to produce two moles of ammonia according to the following reaction:

$$3\,\mathrm{H_2(g)} + \underset{\text{6 moles}}{\mathrm{N_2(g)}} \longrightarrow \underset{\text{2 moles}}{2\,\mathrm{NH_3(g)}}$$

Calculate the percent yield.

Ans: 16.6%.

41. We mix three moles of hydrogen gas with three moles of nitrogen gas.

$$3\,\mathrm{H_2(g)} + \underset{\text{6 moles}}{\mathrm{N_2(g)}} \longrightarrow \underset{\text{2 moles}}{2\,\mathrm{NH_3(g)}}$$

Calculate the percent limiting reagent.

Ans: H_2.

42. We mix three moles of hydrogen gas with half a mole of nitrogen gas.

$$3\,\mathrm{H_2(g)} + \underset{\text{6 moles}}{\mathrm{N_2(g)}} \longrightarrow \underset{\text{2 moles}}{2\,\mathrm{NH_3(g)}}$$

Calculate the percent limiting reagent.

Ans: N_2.

43. We mix two moles of hydrogen gas with five moles of nitrogen gas.

$$3\,\mathrm{H_2(g)} + \underset{\text{6 moles}}{\mathrm{N_2(g)}} \longrightarrow \underset{\text{2 moles}}{2\,\mathrm{NH_3(g)}}$$

Calculate the percent limiting reagent.

Ans: H_2.

MASS CALCULATIONS

44. Calculate the number of grams of nitrogen needed to react with 4 moles of hydrogen, to produce ammonia:

$$\underset{\text{4 moles}}{3\,\mathrm{H_2(g)}} + \underset{\underline{\qquad}\ \text{grams}}{\mathrm{N_2(g)}} \longrightarrow 2\,\mathrm{NH_3(g)}$$

Ans: 37.3 g

45. Calculate the number of grams of hydrogen needed to react with 0.3 moles of nitrogen, to produce ammonia:

$$\underset{\underline{\qquad}\ \text{grams}}{3\,\mathrm{H_2(g)}} + \underset{\text{0.3 moles}}{\mathrm{N_2(g)}} \longrightarrow 2\,\mathrm{NH_3(g)}$$

Ans: 37.3 g

46. Calculate the number of grams of hydrogen needed to react with 0.3 moles of nitrogen, to produce ammonia:

$$\underset{\underline{\qquad}\ \text{grams}}{3\,\mathrm{H_2(g)}} + \underset{\text{0.3 moles}}{\mathrm{N_2(g)}} \longrightarrow 2\,\mathrm{NH_3(g)}$$

Ans: 37.3 g

REVIEW PART B

1. The name of $CO_3{}^{-2}$ is:
   (a) carbon trioxide
   (b) carbon oxide
   (c) carbonic
   (d) carbonate
   (e) bicarbonate
2. Which of the following chemicals is covalent:
   (a) CO
   (b) NO
   (c) NO_2
   (d) H_2O
   (e) all above
3. Name the following chemical: NO_2
   (a) Nitrogen oxide
   (b) Nitrogen oxigen
   (c) Nitrate
   (d) Nitrite
   (e) Nitrogen dioxide
4. The correct name for the compound N_2O_3 is
   (a) nitrogen oxide.
   (b) nitrogen trioxide.
   (c) dinitride trioxide.
   (d) dinitrogen oxide.
   (e) dinitrogen trioxide.
5. What is the formula for aluminum nitrite?
   (a) Al_2NO_2
   (b) $AlNO_3$
   (c) $Al(NO_2)_3$
   (d) $Al_2(NO_3)_3$
   (e) $Al_2(NO_2)_2$
6. $Fe_2(SO_4)_3$ is called
   (a) iron sulfate.
   (b) iron (II) sulfate.
   (c) iron (III) sulfate.
   (d) diiron trisulfate.
   (e) iron trisulfate.
7. What is the formula of a compound that contains Na^+ and $PO_4{}^{3-}$ ions?
   (a) Na_3PO_4
   (b) $NaPO_4$
   (c) Na_2PO_3
   (d) Na_3PO_3
   (e) Na_3P
8. Calculate the molar mass of magnesium chloride, $MgCl_2$.
   (a) 24.3 g
   (b) 95.2 g
   (c) 125.9 g
   (d) 59.8 g
   (e) 70.0 g
9. What is the molar mass of $Mg_3(PO_4)_2$, a substance formerly used in medicine as an antacid?
   (a) 71.3 g
   (b) 118.3 g
   (c) 150.3 g
   (d) 214.3 g
   (e) 262.9 g
10. How many moles of carbon atoms are there in 0.500 mole of C_2H_6?
    (a) 0.500 moles
    (b) 1.00 moles
    (c) 3.00 moles
    (d) 6.02×10^{23} moles
    (e) 4.00 moles
11. How many grams of Fe_2O_3 are there in 0.500 mole of Fe_2O_3?
    (a) 79.8 g
    (b) 35.9 g
    (c) 63.8 g
    (d) 51.9 g
    (e) 160. g
12. How many oxygen atoms are present in 75.0 g of H_2O?
    (a) 75.0 atoms
    (b) 4.17 atoms
    (c) 7.53×10^{24} atoms
    (d) 2.51×10^{24} atoms
    (e) 5.02×10^{24} atoms
13. Which of the following correctly gives the best coefficients for the reaction below?

    $$N_2H_4 + H_2O_2 \longrightarrow N_2 + H_2O$$

    (a) 1, 1, 1 ,1
    (b) 1, 2, 1, 4
    (c) 2, 4, 2, 8
    (d) 1, 4, 1, 4
    (e) 2, 4, 2, 4
14. How many moles of iron are present in 3.15×10^{24} atoms of iron?
    (a) 5.23
    (b) 1.90
    (c) 292
    (d) 0.523
    (e) 1.90×10^{48}
15. What coefficient is placed in front of O_2 to complete the balancing of the following equation?

    $$C_5H_8 + x\,O_2 \longrightarrow 5\,CO_2 + 4\,H_2O$$

    (a) 1
    (b) 3
    (c) 5
    (d) 7
    (e) 9
16. Consider the following equation.

    $$2\,Mg + O_2 \longrightarrow 2\,MgO$$

    How many grams of MgO are produced when 40.0 grams of O_2 react completely with Mg?
    (a) 30.4 g
    (b) 50.4 g
    (c) 60.8 g
    (d) 101 g
    (e) 201 g
17. How many grams of CO_2 are produced from 125 g of O_2 and excess CH_4?

    $$CH_4 + 2\,O_2 \longrightarrow CO_2 + 2\,H_2O$$

    (a) 125 g of CO_2
    (b) 62.5 g of CO_2
    (c) 172 g of CO_2
    (d) 85.9 g of CO_2
    (e) 250. g of CO_2
18. When 3.05 moles of CH_4 are mixed with 5.03 moles of O_2 the limiting reactant is

    $$CH_4 + 2\,O_2 \longrightarrow CO_2 + 2\,H_2O$$

    (a) CH_4
    (b) O_2
    (c) CO_2
    (d) H_2O
    (e) None of the above
19. Consider the following equation.

    $$2\,Mg + O_2 \longrightarrow 2\,MgO$$

    Calculate the yield if 3 moles of Mg produce 1.5 moles of MgO
    (a) 100%
    (b) 150%
    (c) 50%
    (d) 25%
    (e) 75%

Answers:

1. (d)	**2.** (e)	**3.** (e)	**4.** (e)
5. (c)	**6.** (c)	**7.** (a)	**8.** (b)
9. (e)	**10.** (b)	**11.** (a)	**12.** (d)
13. (b)	**14.** (a)	**15.** (d)	**16.** (b)
17. (d)	**18.** (b)	**19.** (c)	

PART C

5

Reactions in solution

THE most common reactions happen in solution. Think for example when you add salt to your soup or when a metal exposed to air rusts when it gets wet. The first is a dissociation reaction, whereas the second is a redox reaction. This chapter first covers the properties of solutions and you will learn how to quantify the amount of solute in a solution. This chapter also covers some important types of reactions happening in solution. Another important concept covered in this chapter is the idea of electrolytes. Most of you will be surprised to know that water do not conduct the electricity. This is because pure water is a weak electrolyte. The importance of electrolytes is well know among the sports community. If you have ever played a sport, you have probably chugged an sports drink. These are actually electrolyte solutions with extra sugar added. However, few know the specifics of their function. Electrolytes are actually salts that conduct electricity in water by separating into positive and negative ions. Here you will be able to identify different types of electrolytes. Finally, here we will also cover some important reactions happening in solution. In particular reactions between acids and base and reaction that result in a solid.

GOALS

1. Carry composition calculations
2. Classify electrolytes
3. Identify insoluble compounds
4. Write down net ionic equations
5. Balance redox reactions

5.1 Solutions and composition

This section covers the basic of solutions. First, solutions are not just a simple mixture of two components. The state of the matter of both components of the mixture or the polarity affect the formation of a solution. For example, a solution does not result from mixing oil and water.

`Discussion:` List three solutions in your household containing just a single solute. Give the chemical formula of the solute and the name of the solvent.

`What makes a solution?` Solutions are homogeneous mixtures of solute and solvent. Homogeneous means that if you look at the mixture you will not be able to differentiate both components and you will only see the mixture as 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 a small amount of sugar with water. Sugar is sweet and water tasteless. When you mix both, you form a solution of sugar and you will not see the sugar in the solution. In this particular example, sugar will be the solute in the solution, as the sugar is in less amount than water. A solution is the result or mixing a solute and a solvent:

$$\boxed{\text{Solution} = \text{Solute} + \text{Solvent}}$$

Add this relation into your flashcard.

`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.

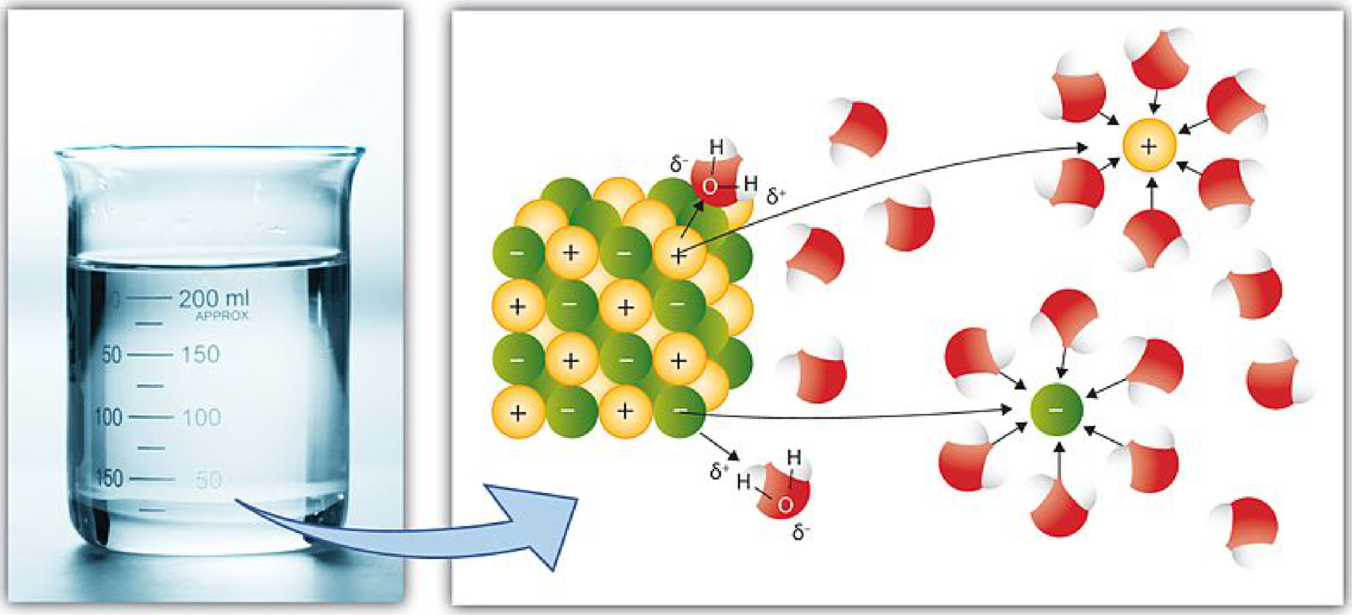


Figure 5.1: A solution results from dissolving a solute into a solvent

Empirical rules of polarity The affinity between a chemical and the mobile phase is connected 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. An example of this is H_2 molecule, which is non-polar. Differently, HF is a polar molecule, as F likes more the electrons than H and these will spend a longer time along fluorine (you will learn more about this effect at the end of the semester). The polar nature of substances is related to its miscibility and molecules with similar polar character will mingle and mix together creating a single visible phase. As an example, water (polar) and methanol (polar) will mix together. Differently, water (polar) and oil (non-polar) are immiscible due to its different polar nature and they will not mix. Even if the rules or polarity are based of 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:

- Diatomic molecules made of the same element are non-polar.
- Diatomic molecules made of different elements are polar.
- Poliatomic molecules (with more than four atoms) made of C and H are in general non-polar.
- Poliatomic molecules (with more than four atoms) containing C, H and a different atom are in general polar.

Sample Problem 40

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 only, 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.

Answer: polar, nonpolar, nonpolar, polar.

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) 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).

Solvent	Solute	Mixing?
Polar	Polar	Yes
Polar	Nonpolar	No
Nonpolar	polar	No
Nonpolar	Nonpolar	Yes

Sample Problem 41

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 non-polar. (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 $CCl_4(l)$

Answer: (a) will mix; (b) will not mix.

One of the most important properties of a solution is its concentration. In this section you will learn how to calculate solute concentrations. The larger the concentration of a solution the more solute in the solution. There are many different concentration units, such as molarity or mass percent concentration. This section will introduce you to some of the most important concentration units.

Meaning of concentration The concentration of a solution refers to the amount of solute with respect to the amount of solution. The larger concentration the larger number of solute particles with respect to the solution.

Mass percent concentration The mass percent (m/m) is the amount of solute in grams per grams of solution in percent form

Add this relation into your flashcard.

$$\boxed{m/m = \frac{\text{g of solute}}{\text{g of solution}} \times 100}$$

Add this relation into your flashcard.

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 42

A NaCl solution is prepared by mixing 4g of NaCl (MW=58.4$g \cdot mol^{-1}$) 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:

$$m/m = \frac{\text{g of solute}}{\text{g of solution}} \times 100 = \frac{\text{4 g of solute}}{\text{54 g of solution}} \times 100$$

The result is 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= $5.2 \cdot 10^{-2}L$. To calculate the moles of solute, we will use the grams of solute and the molar mass to convert this value into moles: $4g/58.4g \cdot mol^{-1}$ =0.068moles. Plugging all values into the molarity formula:

$$M = \frac{\text{moles of solute}}{\text{L of solution}} = \frac{\text{0.068 moles of solute}}{5.2 \cdot 10^{-2}\text{L of solution}} = 1.31M$$

STUDY CHECK

(a) A KCl solution is prepared by mixing 8g of NaCl (MW=74$g \cdot mol^{-1}$) with 250mL of H_2O. Calculate the molarity; (b) A KCl solution is prepared by mixing 5g of KCl with 200g of H_2O. Calculate the percent (m/m) of the solution.

Answer: 0.43M; 2.4 %.

Figure 5.2: Champagne is a solution of gas in a liquid

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 is 3 moles of solute every one litter of solution.

$$\boxed{3\text{M} \quad \text{or} \quad \frac{\text{3 mol of solute}}{\text{1 L of solution}} \quad \text{or} \quad \frac{\text{1 L of solution}}{\text{3 mol of solute}}}$$

Concentration	Conversion
3M	$\frac{\text{3 moles of solute}}{\text{1 L of solution}}$
5% (m/m)	$\frac{\text{5 grams of solute}}{\text{100 g of solution}}$

830

Figure 5.3: Conversion factors from concentration units

Similarly, if the mass percent of a solution is 5% this means that there is 5 grams of solute every 100 grams of solution. We often use concentration units as conversion factors when we need to transform between on unit on top (bottom) of the conversion factor and the unit on the bottom (top).

Sample Problem 43

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:

$$5\,\cancel{\text{moles of solute}} \times \frac{1 \text{ L of solution}}{4\,\cancel{\text{moles of solute}}} = 1.25\text{L}$$

This means that 1.25L of a 4M solution will provide 5 moles of solute.

STUDY CHECK

How many liters of a 6% (m/v)solution do you need to provide 5 grams of solute.

Answer: 8.3×10^{-2}L.

Figure 5.4: We dilute by adding water to a stock solution

Dilution Dilution is the process for preparing a diluted solution from a more concentrated solution. Solutions are often times stored in a stock room in concentrated form. These stocks should be diluted before use. This section covers the dilution process in order to estimate the amount of concentrated solution needed to prepare a more diluted solution. In order 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. In order to answer this we should use the following formula:

$$\boxed{c_1 \cdot V_1 = c_2 \cdot V_2}$$

Sample Problem 44

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.

Answer: 1.8 L.

835

Figure 5.5: Common ingredients found in energy drinks are caffeine or sugars

5.2 Electrolytes and insoluble compounds

Some chemicals when dissolved in water conduct the electricity, other do not. Electrolytes are solutes that once dissolved in water conduct the electricity. Differently, nonelectrolytes do not conduct the electricity in water. This section covers electrolytes, its properties and characteristics. You will be able to tell whether a solute conduct the electricity or not. Some other chemicals are insoluble. This means they do not dissolve in water. Here you will be able to identify insoluble compounds.

`Strong electrolytes` These are a type of electrolytes that completely dissociate in water. Hence in a solution of a strong electrolyte you will only have ions. Strong electrolytes are typically ionic compounds such as $MgCl_2$:

$$MgCl_2(s) \xrightarrow{H_2O} Mg^{2+}(aq) + 2\,Cl^-(aq).$$

`Weak electrolytes` These electrolytes partially dissociate in water, and this is indicated by means of a chemical reaction with a double arrow. Hence in a solution of a weak electrolyte you will have ions as well as molecules. Examples of weak electrolytes are hydrofluoric acid, water, ammonia or acetic acid. The dissociation of HF proceeds as:

$$HF(g) \xrightleftharpoons{H_2O} H^+(aq) + F^-(aq).$$

`Nonelectrolytes` Nonelectrolytes do not dissociate in water. Hence a nonelectrolyte solution only contains molecules. Examples of weak electrolytes are carbon-based chemicals such as methanol, ethanol, urea or sucrose. The dissociation of urea for example CH_4N_2O proceeds as:

$$CH_4N_2O(s) \xrightarrow{H_2O} CH_4N_2O(aq).$$

Sample Problem 45

For the following chemicals indicate whether you will have (a) only ions on solution, (b) ions and some molecules, or (c) molecules:

Chemical	Particles in solution
NH_3	
KOH	
$C_{12}H_{22}O_{11}$	

SOLUTION

Ammonia is a weak electrolyte and a solution of ammonia will contain ions and well as ammonia molecules. Differently KOH is a strong electrolyte and a solution of this chemical will contain only ions (K^+ and OH^-). Sucrose ($C_{12}H_{22}O_{11}$) is a nonelectrolyte and sucrose solution will only contain molecules. In table format:

Chemical	Particles in solution
$NH_{3(g)}$	$NH_{3(aq)} + NH_4^+ + OH^-$
$KOH_{(s)}$	$K^+ + OH^-$
$C_{12}H_{22}O_{11(s)}$	$C_{12}H_{22}O_{11(aq)}$

STUDY CHECK

For the following chemicals indicate whether you will have only ions on solution, ions and some molecules, or molecules: (a) H_2SO_4, HNO_3 and

CH_3OH.

Answer: (a) ions, (b) ions and (c) molecules.

Electrolyte Type	Dissociation	Particles in solution	Examples
Strong	Fully	Mostly ions	Ionic Compounds and most acids and bases: NaCl, NaOH, HCl, $MgCl_2$, H_2SO_4, etc ·
Weak	Partially	Ions & molecules	NH_3, CH_3COOH, 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)

Figure 5.6: Different types of electrolytes

Breaking down electrolytes into ions Electrolytes–in particular strong electrolytes–dissociate forming 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 break down electrolytes into ions correctly. In order to do this, you need to revert what you did while combining ions to name chemicals. First, it is important to point that you will only be able to break down strong electrolytes into ions, as weak electrolytes dissociate only partially and nonelectrolytes do not dissociate at all. For example, let us beak magnesium chloride $MgCl_{2(aq)}$ into ions. We can break down this chemical as it is a strong electrolyte. This is a strong electrolytes formed by magnesium and chloride. 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 if magnesium nitrite $Mg(NO_3)_2$. This strong electrolyte–as this is an ionic salt–is made of lithium with valence I and nitrate with valence negative one. 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 46

Break down the following chemicals into ions, if possible:

Chemical	Particles in solution
$K_2CrO_{4(aq)}$	
$Ba(NO_3)_{2(aq)}$	
$Ba(CrO_4)_{2(s)}$	
$KNO_{3(aq)}$	

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 $Ba(CrO_4)_{2(s)}$ 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 $2NO_3{}^-{}_{(aq)}$. Finally, potassium nitrate– $KNO_{3(aq)}$–will produce $K^+_{(aq)}$ and $NO_3{}^-{}_{(aq)}$. In the table:

Chemical	Particles in solution
$K_2CrO_{4(aq)}$	$2K^+_{(aq)} + CrO_4{}^{2-}{}_{(aq)}$
$Ba(NO_3)_{2(aq)}$	$Ba^{2+}_{(aq)} + 2NO_3{}^-{}_{(aq)}$
$Ba(CrO_4)_{2(s)}$	$Ba(CrO_4)_{2(s)}$
$KNO_{3(aq)}$	$K^+_{(aq)} + NO_3{}^-{}_{(aq)}$

STUDY CHECK

Break down the following chemicals into ions, if possible: $H_2O_{(l)}$, $NH_{3(l)}$, $AgNO_{3(aq)}$.

Answer: $H_2O_{(l)}$, $NH_{3(l)}$, $Ag^+_{(aq)}$, $NO_3{}^-{}_{(aq)}$.

Soluble and insoluble salts How do we know that $Ba(CrO_4)_{2(s)}$ is an insoluble salt? Soluble compounds are ionic compounds that dissolve in water. Insoluble salts, differently, will not solve in water. The following table will help you predict the solubility of a salt. In order to predict the solubility of salt you need to start by the right ion of the molecule, and look for it on the left column of the table. After that you need to check if the ion on the left is part of the exceptions in the same row but on the right column of the table. Let us predict for example the soluble/insoluble nature if $CaSO_4$. 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.

Sample Problem 47

Predict the soluble/insoluble nature of the following ionic compounds: (a) K_2CO_3, (b) $NaNO_3$ and (c) $Ca(OH)_2$.

SOLUTION

(a) K_2CO_3 is soluble, as $CO_3{}^{-2}$ is insoluble but when combined with K^+ the salt becomes soluble. (b) All nitrates are soluble, and there are no exceptions. (c) $Ca(OH)_2$ is soluble.

STUDY CHECK

Predict the soluble/insoluble nature of the following ionic compounds: (a) Li_3PO_4, (b) Na_2S and (c) AgCl.

Answer: (a) soluble, (b) soluble and (c) insoluble.

5.3 An introduction to reactions in solution

There are three different reactions in solution: acid-base reactions, precipitation reaction and redox reactions. The key to identify acid-base reactions is in the reactants, as an acid-base reaction results from the reaction between and acid and a base. Precipitation

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^-)	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 ($NH_4{}^+$)

reactions are reactions that produce a precipitate. Hence, the key to identify a precipitation reaction is in the products. Precipitation reactions always contains a solid as a product. Redox reactions contain two elements with different redox number in the reactants and products. The key to identify redox reactions is to be able to spot elements with different oxidation state, 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 they 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) 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 a insoluble chemical, that is, results in a solid chemical. An example would be:
$K_2CrO_{4(aq)} + Ba(NO_3)_{2(aq)} \longrightarrow Ba(CrO_4)_{2(s)}\downarrow + 2\,KNO_{3(aq)}$ (precipitation reaction)
The chemical $Ba(CrO_4)_{2(s)}$ is a solid that precipitates in the solution, hence the name of the type of reaction. The symbol on $Ba(CrO_4)_{2(s)}\downarrow$ represents the precipitation process. The solubility of a given solute such as $Ba(CrO_4)_{2(s)}$ is the amount of solute (in grams) that can be dissolve 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 reaction are different than acid-base or precipitation reaction. They contain the same chemical element in two different states resulting from the loss or win of electrons. Look for example:

$$2\,Al_{(s)} + 3\,Cu_{(aq)}{}^{+2} \longrightarrow 2\,Al_{(aq)}{}^{+3} + 3\,Cu_{(s)}\text{(redox reaction)}$$

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 product precipitate, hence this is not a precipitation reaction. Differently, 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 electron. Therefore in redox reaction there is 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 48

Classify the following reactions as acid-base, redox or precipitation.
(a) $Fe_{(s)} + Cu_{(aq)}{}^{+2} \longrightarrow Fe_{(aq)}{}^{+2} + Cu_{(s)}$
(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) $MgCl_{2(aq)} + 2\,AgNO_{3(aq)} \longrightarrow 2\,AgCl_{(s)} + Mg(NO_3)_{2(aq)}$

Answer: acid-base; redox; precipitation

5.4 Precipitation reactions and acid-base reactions

This section deal with two important types of reactions in solution. Precipitation reactions are characterized by the products and acid-base by the reactants. In an acid-base reaction, the reactants are an acid and a base, and they react to produce water and other chemical. Precipitation reactions produce a precipitate, that is, a solid.

Acid-base reactions Acid-base reactions result from the reaction of an acid with a base. Both they produce water and another chemical. An example is:

$$HBr_{(aq)} + KOH_{(aq)} \longrightarrow KBr_{(aq)} + H_2O_{(l)}\ \text{(acid-base reaction)}$$

Hydrobromic acid $HBr_{(aq)}$ is an acid and potassium hydroxide a base.

Precipitation reactions Precipitation reactions result in a insoluble chemical. An example would be:

$K_2CrO_{4(aq)} + Ba(NO_3)_{2(aq)} \longrightarrow Ba(CrO_4)_{2(s)} \downarrow + 2\,KNO_{3(aq)}$ (precipitation reaction)

The chemical $Ba(CrO_4)_{2(s)}$ is a solid that precipitates in the solution. The solubility of a given solute such as $Ba(CrO_4)_{2(s)}$ is the amount of solute (in grams) that can be dissolve 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.

Formula equations, ionic equations and net ionic equations Electrolytes in solutions contains ions–cations and anions–however, when we write chemical formulas we barely show those ions. Differently, we just write the formulas and that is the reason that chemical equations are referred as *formula equation*. Look for example:

$$K_2CrO_{4(aq)} + Ba(NO_3)_{2(aq)} \longrightarrow Ba(CrO_4)_{2(s)} \downarrow + 2\,KNO_{3(aq)} \text{ (formula equation)}$$

In this equation, $K_2CrO_{4(aq)}$ is actually in the form of ions: $2K^+_{(aq)}$ and $CrO_4{}^{2-}{}_{(aq)}$. At the same time, $Ba(NO_3)_{2(aq)}$ in the form of ions results in $Ba^{2+}_{(aq)}$ and $2NO_3{}^-{}_{(aq)}$. Also, $2KNO_{3(aq)}$ contains $2K^+_{(aq)}$ and $2NO_3{}^-{}_{(aq)}$. Finally, $Ba(CrO_4)_{(s)}$ does not produce any ions in solution, as it is a solid. Ionic equations result from writing all ions in a formula equation:

$$2\,K^+_{(aq)} + CrO_4{}^{2-}{}_{(aq)} + Ba^{2+}_{(aq)} + 2\,NO_3{}^-{}_{(aq)} \longrightarrow Ba(CrO_4)_{2(s)} \downarrow + 2\,K^+_{(aq)} + 2\,NO_3{}^-{}_{(aq)} \text{ (ionic equation)}$$

However, the ionic equation contains repeated ions. Look for example the previous equation with $2\,K^+_{(aq)}$ on the left and on the right of the equation. If we simplify the repeated ions

$$\cancel{2\,K^+_{(aq)}} + CrO_4{}^{2-}{}_{(aq)} + Ba^{2+}_{(aq)} + \cancel{2\,NO_3{}^-{}_{(aq)}} \longrightarrow Ba(CrO_4)_{2(s)} + \cancel{2\,K^+_{(aq)}} + \cancel{2\,NO_3{}^-{}_{(aq)}}$$

we obtain what is called as the *net ionic equation*:

$$CrO_4{}^{2-}{}_{(aq)} + Ba^{2+}_{(aq)} \longrightarrow Ba(CrO_4)_{2(s)} \downarrow \text{ (net ionic equation)}$$

Overall, we have that the formula equation, ionic equation and net ionic equation are just three different ways to express the same chemical equation. The first form includes only molecules whereas the second included all ions produced by each chemical. The last form, includes only ions that are not repeated in both sides of the equation.

Sample Problem 49

Write down the ionic equation and net ionic for the following formula equation:

$$HBr_{(aq)} + KOH_{(aq)} \longrightarrow KBr_{(aq)} + H_2O_{(l)}$$

SOLUTION

Mind we can only break down strong electrolytes. Hence, water will not be expressed in the form or ions as it is a weak electrolyte. If we break down the other chemicals we have the ionic equation:

$$H^+_{(aq)} + Br^-_{(aq)} + K^+_{(aq)} + OH^-_{(aq)} \longrightarrow K^+_{(aq)} + Br^-_{(aq)} + H_2O_{(l)}$$

If we eliminate the ions that are repeated in both sides:

$$H^+_{(aq)} + \cancel{Br^-_{(aq)}} + \cancel{K^+_{(aq)}} + OH^-_{(aq)} \longrightarrow \cancel{K^+_{(aq)}} + \cancel{Br^-_{(aq)}} + H_2O_{(l)}$$

we have the net ionic equation:

$$H^+_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(l)}$$

STUDY CHECK

Write down the ionic equation and net ionic for the following formula equation:

$$AgNO_{3(aq)} + NaBr_{(aq)} \longrightarrow AgBr_{(s)}\downarrow + NaNO_{3(aq)}$$

Answer: $Ag^+_{(aq)} + NO_3{}^-{}_{(aq)} + Na^+_{(aq)} + Br^-_{(aq)} \longrightarrow AgBr_{(s)}\downarrow + Na^+_{(aq)} + NO_3{}^-{}_{(aq)}$; $Ag^+_{(aq)} + Br^-_{(aq)} \longrightarrow AgBr_{(s)}\downarrow$

5.5 Redox reactions

Redox reaction are different than acid-base or precipitation reaction. They contain the same chemical element in two different states resulting from the loss or win of electrons. Look 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, therefore this is not an acid-base reaction. Also there is no product precipitate, hence this is not a precipitation reaction. Differently, 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. Therefore in redox reaction there is always elements in the chemicals that lose electrons and chemicals winning electrons.

`Oxidation state or redox number` How do we know if a chemical has lost of won electrons? The answer is by means of a number called redox number or oxidation state. There is four rules to identify the redox number of an element. First, single atoms or elements have zero redox number. Examples are Na or H_2, both with redox zero. Second, monoatomic ions have redox number equal to their charge. Examples are Na^+ or Cl^- with redox $+1$ and -1, respectively. Third, the redox number of fluorine is -1 and hydrogen on its covalent compounds $+1$. Fourth, the redox number of oxygen in normal oxides is normally -2, with the exception of peroxides (e.g. H_2O_2) in which is -1. We indicate redox numbers with roman number on top of the element. For example the redox number of manganese in this compound is $+7$: $\underline{Mn}^{VII}O_4{}^-$.

`Calculating the redox number` How do we calculate the redox number for example of manganese in this chemical: $\underline{Mn}O_4{}^-$, permanganate. IN order 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–equals 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 a redox number of manganese of VII.

Sample Problem 50

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) $\underline{Cr_2}O_3$

Answer: VI; III.

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 $\underline{Cr^{VI}{}_2}O_7{}^{2-}$ and $\underline{Cr^{III}{}_2}O_3$. 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 an element, and that means the element has lost electrons. The smaller the redox number the more reduced is the element and that means it has gained electrons. If we compare both case, 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 process, a reduction and the oxidation. These two processes can be separated into two half-reactions so that the combination of both half-reactions lead to the 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 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 negative charge and we add electrons to compensate 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. In the product side we have $Cu_{(s)}$ with zero redox number. Cu has undergone reduction as its redox number has decreases. 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 involve different number of electrons. Hence in order to be able to add both redox we need to time each half-reaction by a number so that the number of electrons cancel out. As the first reaction involved three electrons and the second two, we will do:

$$\begin{array}{lr} 2 \cdot \left(Al_{(s)} \longrightarrow {Al_{(aq)}}^{+3} + 3\,e^{-} \right) & \text{(oxidation)} \\ 3 \cdot \left({Cu_{(aq)}}^{+2} + 2\,e^{-} \longrightarrow Cu_{(s)} \right) & \text{(reduction)} \\ \hline 2\,Al_{(s)} + 3\,{Cu_{(aq)}}^{+2} + \cancel{6e^{-}} \longrightarrow 2\,{Al_{(aq)}}^{+3} + 3\,Cu_{(s)} + \cancel{6e^{-}} & \text{(redox)} \end{array} +$$

The overall balanced redox equation is:

$$\overbrace{2\ \overset{0}{Al}(s) + 3\ \overset{II}{Cu^{2+}}(aq) \longrightarrow 2\ \overset{III}{Al^{3+}}}^{\text{Oxidation:}-3e^{-}}(aq) + 3\ \overset{0}{Cu}(s)$$

$$\underbrace{3\ Cu^{2+}(aq) \longrightarrow 2\ Al^{3+}(aq) + 3\ Cu(s)}_{\text{Reduction:}+2e^{-}}$$

Balancing redox reactions in acidic medium Redox reactions happen in either basic or acidic medium. Here we will go over how to balance redox reactions in acidic medium. In order to do this we will first separate the reaction in two half-reactions. In each semi-reaction we will balance all elements but hydrogen and oxygen. When all elements are balanced, we will proceed to balance O by adding H_2O molecules and we will balance H by adding H^{+}. Finally, we all add electrons to compensate the charge of the reaction. Let us work on an example:

$${MnO_{4(aq)}}^{-} + {Fe_{(aq)}}^{2+} \longrightarrow {Mn_{(aq)}}^{2+} + {Fe_{(aq)}}^{3+}$$

One of the semi-reactions involve Manganese whereas the other involves Iron. The redox number of Mn in permanganate is VII hence Manganese is being reduced, as its redox number decreases from VII to II, whereas Iron is being oxidized as its redox number increases from II to III. The oxidation half-reaction does not contain hydrogen or oxygen hence we will only have to balance the charger with one electron:

$$Fe_{(aq)}{}^{2+} \longrightarrow Fe_{(aq)}{}^{3+} + e^- \text{ (oxidation)}$$

The reduction half-reaction contains oxygen. Hence, we will have to add H_2O molecules to balance oxygen and H^+ to balance hydrogen. In particular, we will need two water molecules–as $MnO_4{}^-$ has four oxygens and we will have to add four protons as we are adding two molecules of water. Finally, we need to add three electrons to equalize the charge:

$$MnO_{4(aq)}{}^- + 4\,H_{(aq)}{}^+ + 3\,e^- \longrightarrow Mn_{(aq)}{}^{2+} + 2\,H_2O_{(l)} \text{ (reduction)}$$

As the oxidation involves one electron and the reduction three, we need to time the oxidation by three:

$$\begin{array}{lr}
3 \cdot \left(Fe_{(aq)}{}^{2+} \longrightarrow Fe_{(aq)}{}^{3+} + e^- \right) & \text{(oxidation)} \\
MnO_{4(aq)}{}^- + 4\,H_{(aq)}{}^+ + 3\,e^- \longrightarrow Mn_{(aq)}{}^{2+} + 2\,H_2O_{(l)} & \text{(reduction)} \\
\hline
3\,Fe_{(aq)}{}^{2+} + MnO_{4(aq)}{}^- + 4\,H_{(aq)}{}^+ + \cancel{3e^-} \longrightarrow & \text{(redox)} \\
\qquad 3\,Fe_{(aq)}{}^{3+} + Mn_{(aq)}{}^{2+} + 2\,H_2O_{(l)} + \cancel{3e^-} &
\end{array}$$

$$4\ H^+(aq) + \overset{VII}{Mn}O_4(aq) + 3\ \overset{II}{Fe}{}^{2+}(aq) \longrightarrow 3\ \overset{III}{Fe}{}^{3+}(aq) + \overset{II}{Mn}{}^{2+}(aq) + 2\ H_2O(l)$$

$$\text{Oxidation: } -1e^- \qquad \text{Reduction: } +3e^-$$

Sample Problem 51

Balance the following redox in acidic medium:

$$Cr_2O_7{}^{2-} + SO_3{}^{2-} \longrightarrow Cr^{3+} + SO_4{}^{2-}$$

SOLUTION

We locate the same element in both sides of the reaction with different redox number. We found Chromium in the form of dichromate ($Cr_2O_7{}^{2-}$) with redox number VI and Chromium in the product side with redox III. Therefore Chromium is being reduced, as its redox number decreases. Differently, we found sulphur in the reactants side in the form of sulfite ($SO_3{}^{2-}$) with redox number of IV and in the product side with redox number of VI. Therefore Sulphur is being oxidized. We will first set up the oxidation half-reaction knowing that we have different amounts of Cr in both side and that we will have to add water molecules to balance O and protons to balance H. As we have seven oxygens in $Cr_2O_7{}^{2-}$ we will have to add seven water molecules. Also, as we add seven water molecules we will have to add fourteen protons. We will need six electrons to compensate the charge:

$$Cr_2O_7{}^{2-} + 14\,H^+ + 6\,e^- \longrightarrow 2\,Cr^{3+} + 7\,H_2O \text{ (reduction)}$$

For the reduction half-reaction, we have a difference of one oxygen atoms and hence we will need one water molecule and two protons; we will need two electrons to compensate the charge:

$$SO_3{}^{2-} + H_2O \longrightarrow SO_4{}^{2-} + 2\,H^+ + 2\,e^- \text{ (oxidation)}$$

In order to add both half-reactions as the reduction involves two electrons and the oxidation six, we will have to multiply the reduction half-reaction by three:

$$1 \cdot \left(Cr_2O_7{}^{2-} + 14\,H^+ + 6\,e^- \longrightarrow 2\,Cr^{3+} + 7\,H_2O \right) \qquad \text{(reduction)}$$
$$3 \cdot \left(SO_3{}^{2-} + H_2O \longrightarrow SO_4{}^{2-} + 2\,H^+ + 2\,e^- \right) \qquad \text{(oxidation)}$$

$$Cr_2O_7{}^{2-} + 14\,H^+ + 3\,SO_3{}^{2-} + 3\,H_2O + \cancel{6\,e^-} \longrightarrow 2\,Cr^{3+} + 7\,H_2O + 3\,SO_4{}^{2-} + 6\,H^+ + \cancel{6\,e^-} \qquad \text{(redox)}$$

After we cancel the electrons and eliminate protons and water molecules, the balanced reaction is:

Oxidation: $-1e^-$

$$\overset{\text{VI}}{Cr_2}O_7{}^{2-}(aq) + 8\,H^+(aq) + 3\,\overset{\text{IV}}{S}O_3{}^{2-}(aq) \longrightarrow 3\,\overset{\text{VI}}{S}O_4{}^{2-}(aq) + 2\,\overset{\text{III}}{Cr}{}^{3+}(aq) + 7\,H_2O\,(l)$$

Reduction: $+3e^-$

STUDY CHECK

Balance the following redox in acidic medium:

$$Cr_2O_2{}^{-7}{}_{(aq)} + NO_2{}^-{}_{(aq)} \longrightarrow Cr^{3+}_{(aq)} + NO_3{}^-{}_{(aq)}$$

Answer: $3\,NO_{2(aq)} + 8\,H^+_{(aq)} + Cr_2O_2{}^{7-}{}_{(aq)} \longrightarrow 3\,NO_3{}^-{}_{(aq)} + 2\,Cr^{3+}_{(aq)} + 4\,H_2O_{(l)}$.

Balancing redox reactions in basic medium In order to balance a redox in basic medium we need first to balance the reaction in acidic medium. After, we will compensate all H^+ with OH^- in both sides of the reaction. Mind that when combining H^+ with OH^- we obtain H_2O. For example, in order to balance the following reaction in basic medium:

$$3\,Fe_{(aq)}{}^{2+} + MnO_{4(aq)}{}^- + 4\,H_{(aq)}{}^+ \longrightarrow 3\,Fe_{(aq)}{}^{3+} + Mn_{(aq)}{}^{2+} + 2\,H_2O_{(l)}$$

we will add four OH^- in both sides:

$$3\,Fe_{(aq)}{}^{2+} + MnO_{4(aq)}{}^- + 4\,H_{(aq)}{}^+ + 4\,OH_{(aq)}{}^- \longrightarrow 3\,Fe_{(aq)}{}^{3+} + Mn_{(aq)}{}^{2+} + 2\,H_2O_{(l)} + 4\,OH_{(aq)}{}^-$$

And after cancelling the four protons with the four hydroxyls, we have:

$$3\,Fe_{(aq)}{}^{2+} + MnO_{4(aq)}{}^- + 4\,H_2O_{(l)} \longrightarrow 3\,Fe_{(aq)}{}^{3+} + Mn_{(aq)}{}^{2+} + 2\,H_2O_{(l)} + 4\,OH_{(aq)}{}^-$$

Now we have four water molecules in the left and two in the right. We will cancel two molecules from each side in order not to report water molecules twice:

$$3\,Fe_{(aq)}{}^{2+} + MnO_{4(aq)}{}^- + 2\,H_2O_{(l)} \longrightarrow 3\,Fe_{(aq)}{}^{3+} + Mn_{(aq)}{}^{2+} + 4\,OH_{(aq)}{}^-$$

CHAPTER 5

SOLUTIONS

1. A solution is prepared by mixing 4 g of C_6H_6(l) and 5 g of CCl_4(l). Indicate the true statement.

(a) C_6H_6 is the solute
(b) CCl_4 is the solute
(c) Both do not mix
(d) This is not a solution

Ans: (a)

2. Which of the following chemicals will mix with CH_3Cl

(a) C_6H_6
(b) CCl_4
(c) H_2O
(d) CH_2Cl_2

Ans: (c)

3. Which of the following chemicals will mix with CCl_4

(a) C_6H_6
(b) $CHCl_3$
(c) H_2O
(d) CH_3Cl

Ans: (a)

4. Air is :

(a) A solid solution
(b) A liquid solution
(c) A gas solution
(d) A heterogeneous mixture

Ans: (c)

5. Ammonia and water:

(a) A solid solution
(b) A liquid solution
(c) A gas solution
(d) A heterogeneous mixture

Ans: (c)

6. Oil and water do not mix due to a polarity difference. Explain why a detergent can help solve oil in water.

CONCENTRATION UNITS

7. Sodium hydroxide NaOH is a chemical used in drain cleaners. A solution is prepared by mixing 25g of NaOH in 250g of water.

(a) Water is the solute
(b) The percent (m/m) of solute is 9%
(c) The percent (m/m) of solvent is 80%
(d) The percent (m/m) of solvent is 10%

Ans: (b)

8. A solution is prepared by mixing 15g of NaOH in 50g of water.

(a) The percent (m/m) of solute is 30%
(b) The percent (m/m) of solute is 0.3%
(c) The percent (m/m) of solute is 23%
(d) The percent (m/m) of solvent is 10%

Ans: (c)

9. Alcohol-hydroxide is a mixture employed to clean glass. A mixture is prepared by mixing 60g of NaOH with 500g of ethanol.

(a) The percent (m/m) of solute is 8%
(b) The percent (m/m) of solute is 0.08%
(c) The percent (m/m) of solvent is 70%
(d) The percent (m/m) of solvent is 89%

Ans: (d)

10. Vinegar is a (m/m) 5% acetic acid solution. How many grams of acetic acid are there in 2g of vinegar:

(a) 10g
(b) 3g
(c) 0.01g
(d) 0.1 g

Ans: (d)

11. An HCl solution is prepared by mixing 4 moles of HCl with water reaching a volume of 250mL. Calculate the molarity of the solution

(a) 16M
(b) 0.016M
(c) 0.16M
(d) 0.014 M

Ans: (a)

12. How many mL of a 3M KCl solution contains 0.06 KCl moles.

Ans: 20mL

13. How many mL of a 4M NaCl (MW=58$g \cdot mol^{-1}$) solution contains 5 grans of NaCl.

Ans: 21mL

14. How many grams of solute are there in 100mL of a 0.01M HNO_3 (MW=63$g \cdot mol^{-1}$) solution.

Ans: 0.062g

15. How many mL of a 0.001M $Ca(OH)_2$ (MW=74$g \cdot mol^{-1}$) solution can be prepared from 5 mg of $Ca(OH)_2$.

Ans: 67.56mL

16. What is the final volume when 50mL of a 2M NaCl solution is diluted to a 1M.

(a) 25 mL
(b) 50 mL
(c) 125 mL
(d) 100 mL

Ans: d

17. What is the molarity of a solution prepared when 100mL a 4% HCl solution is diluted to a final volume of 500mL.

(a) 0.8 %
(b) 8 %
(c) 20 %
(d) 10 %

Ans: a

18. Describe how to prepare 50mL of a 0.5M H_2SO_4 solution, starting with a 1M stock H_2SO_4 solution.

Ans: 25mL

ELECTROLYTES AND INSOLUBLE COMPOUNDS

19. Indicate the composition of an aqueous solution of NaCl:

(a) Na^+ anions and Cl^- cations

(b) Ions and molecules
(c) Na^+ cations and Cl^- anions
(d) NaCl molecules

Ans: c

20. Indicate the composition of an aqueous solution of HCl:
    (a) Ions and molecules
    (b) H^+ anions and Cl^- cations
    (c) HCl molecules
    (d) H^+ cations and Cl^- anions

Ans: d

21. Indicate the composition of an aqueous solution of $CaCl_2$:
    (a) $CaCl_2$ molecules
    (b) Ca^{2+} cations and Cl^- anions
    (c) Ions and molecules
    (d) Ca^{+2} anions and Cl^- cations

Ans: b

22. Indicate the composition of an aqueous solution of H_2O:
    (a) H_2O molecules
    (b) H^+ cations and OH^- anions
    (c) Ions and molecules
    (d) H^+ anions and OH^- cations

Ans: c

23. Predict the soluble character of the following compound: $AgNO_3$.

Ans: soluble

24. Predict the soluble character of the following compound: AgBr.

Ans: insoluble

25. Predict the soluble character of the following compound: $CaCO_3$.

Ans: soluble

26. Predict the soluble character of the following compound: Na_2CO_3.

Ans: soluble

27. Break down the following chemical into ions if possible: $Ca(OH)_2$.

Ans: $Ca^{2+} + 2\,OH^-$

28. Break down the following chemical into ions if possible: NH_3.

Ans: $NH_3(aq)$

29. Break down the following chemical into ions if possible: K_2CrO_4.

Ans: $2\,K^+ + CrO_4{}^-$

30. Break down the following chemical into ions if possible: $Ca(NO_3)_2$.

Ans: $Ca^+ + 2\,NO_3{}^-$

Precipitation and acid-base reactions

31. Classify the following reaction as acid-base or precipitation:

$$NaCl_{(aq)}{}^{+}AgNO_{3\ (aq)} \longrightarrow NaNO_{3(aq)} + AgCl_{(s)}\uparrow$$

Ans: Precipitation

32. Classify the following reaction as acid-base or precipitation:

$$H_2SO_{4(aq)} + 2\,NaOH_{(aq)} \longrightarrow 2\,H_2O_{(l)} + Na_2SO_{4(aq)}$$

Ans: Acid-base

33. Classify the following reaction as acid-base or precipitation:

$$2\,Na_3PO_{4(aq)} + 3\,CaCl_{2(aq)} \longrightarrow 6\,NaCl_{(aq)} + Ca_3(PO_4)_{2(s)}\downarrow$$

Ans: Precipitation

34. Obtain the net ionic equation for the following reaction:

$$NaCl_{(aq)}{}^{+}AgNO_{3\ (aq)} \longrightarrow NaNO_{3(aq)}{}^{+}AgCl_{(s)}\downarrow$$

Ans: $Ag^+{}_{(aq)} + Cl^-{}_{(aq)} \longrightarrow AgCl_{(s)}\downarrow$

35. Obtain the net ionic equation for the following reaction:

$$H_2SO_{4(aq)} + 2\,NaOH_{(aq)} \longrightarrow 2\,H_2O_{(l)} + Na_2SO_{4(aq)}$$

Ans: $2\,H^+{}_{(aq)} + 2\,OH^-{}_{(aq)} \longrightarrow 2\,H_2O_{(l)}$

36. Obtain the net ionic equation for the following reaction:

$$2\,Na_3PO_{4(aq)} + 3\,CaCl_{2(aq)} \longrightarrow 6\,NaCl_{(aq)} + Ca_3(PO_4)_{2(s)}\downarrow$$

Ans: $3\,Ca^{2+}{}_{(aq)} + 2\,PO_4{}^{2-}{}_{(aq)} \longrightarrow Ca_3(PO_4)_{2(s)}\downarrow$

Redox

37. Balance the following redox reactions in acidic medium:

$$I^-_{(aq)} + MnO_4{}^-{}_{(aq)} \longrightarrow I_{2(s)} + Mn^{2+}_{(aq)}$$

Ans: $10\,I^-_{(aq)} + 2\,MnO_4{}^-{}_{(aq)}{}^{+}\,16\,H^+_{(aq)} \longrightarrow 5\,I_{2(s)} + 2\,Mn^{2+}_{(aq)} + 8\,H_2O_{(l)}$

38. Balance the following redox reactions in acidic medium:

$$MnO_4{}^-{}_{(aq)} + SO_3{}^{-2}{}_{(aq)} \longrightarrow MnO_{2(s)}{}^{+}SO_4{}^{-2}{}_{(aq)}$$

Ans: $2\,MnO_4{}^-{}_{(aq)} + 2\,H^+ + 3\,SO_3{}^{-2}{}_{(aq)} \longrightarrow H_2O_{(l)} + 2\,MnO_{2(s)}{}^{+}\,3\,SO_4{}^{-2}{}_{(aq)}$

39. Convert the following acidic redox into basic redox:

$$2\,MnO_4{}^{-}{}_{(aq)} + 2\,H^{+} + 3\,SO_3{}^{-2}{}_{(aq)} \longrightarrow H_2O_{(l)} + 2\,MnO_{2(s)} + 3\,SO_4{}^{-2}{}_{(aq)}$$

Ans: $2\,MnO_4{}^{-}{}_{(aq)} + H_2O_{(l)} + 3\,SO_3{}^{-2}{}_{(aq)} \longrightarrow 2\,OH^{-}_{(aq)} + 2\,MnO_{2(s)} + 3\,SO_4{}^{-2}{}_{(aq)}$

40. Balance the following redox reactions:

(a) $Fe_{(s)} + O_{2(g)} \longrightarrow Fe^{3+}_{(aq)} + O^{2-}_{(aq)}$

(b) $Cu_{(s)} + Ag^{+}_{(aq)} \longrightarrow Ag_{(s)} + Cu^{2+}_{(aq)}$

41. Balance the following redox reactions in acidic medium:

(a) $Zn_{(s)} + NO_3{}^{-}{}_{(aq)} \longrightarrow Zn^{2+}_{(aq)} + NO_{2(g)}$

(b) $O_{2(g)} + Fe^{2+}_{(aq)} \longrightarrow Fe^{3+}_{(aq)}$

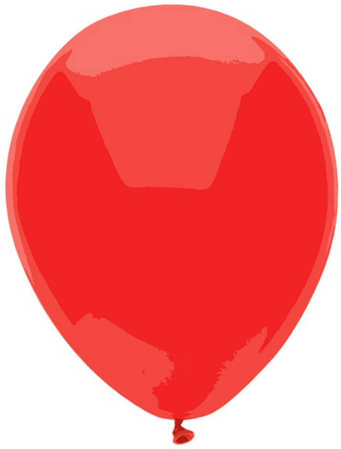
(c) $Sn^{2+}_{(aq)} + Cr_2O_7{}^{2-}{}_{(aq)} \longrightarrow Cr^{3+}_{(aq)} + Sn^{4+}_{(aq)}$

(d) $FeS_{(s)} + NO_3{}^{-}{}_{(aq)} \longrightarrow NO_{(g)} + Fe^{3+}_{(aq)} + SO_4{}^{2-}{}_{(aq)}$

(e) $Sn_{(s)} + NO_3{}^{-}{}_{(aq)} + Cl^{-}_{(aq)} \longrightarrow SnCl_6{}^{2-}{}_{(aq)} + NO_{2(g)}$

6

Reactions in gase phase



THE air we all breath contains numerous gases, such as oxygen, nitrogen or carbon dioxide. Some of these gases are indeed essential for life. As an example, living creature 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 the atmosphere, which contains numerous gases.

GOALS

1. Use the ideal gas law
2. Use the real gas law
3. Calculate partial pressures
4. Compute gas-property changes
5. Carry stoichiometric calculations with volumes

6.1 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 by means of weak forces. Solids on the other hand are packed materials and its 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 of the combination of two atoms of the same element. For example, molecular oxygen (O_2) is gas. Similarly, molecular nitrogen (N_2), molecular hydrogen (H_2), molecular chlorine (C_2), or molecular fluorine (F_2) are all diatomic gases–they contain two atoms of the same element. Other gases result of 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 nobel gases (Ne, He, Ar) also exist in gas state.

Figure 6.1: Oxygen is a flammable gas at room temperature.

`Volume` The volume of a gas (V) is the amount of space it occupies, and gases fully occupy the volume of its container. Liters (L) and milliliters (mL) are units of volume. Liter is a cubic unit and one litter equals to a cubic decimeter (dm^3). Similarly, one milliliter (mL) equals to a cubic centimeter (cm^3).

$$\boxed{\frac{1\ \text{L}}{1\ dm^3}} \quad \boxed{\frac{1\ \text{mL}}{1\ cm^3}}$$

In this chapter all formulas require the use of litters.

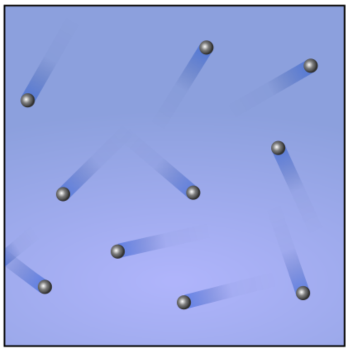


Figure 6.2: Gas pressure is due to the collisions of the gas particles with the walls of its container.

`Temperature` The temperature (T) of a gas is related to the speed (the average velocity) of gas 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 all formulas require the use of Kelvin temperature ($T_K = T_C + 273$).

`Amount of gas` The amount of a gas (n) refers to the quantity of gas particles. The bigger the amount of gas, the larger the amount of gas particles. The amount of gas is measured in moles or grams.

`Pressure, P` The particles of a gas are constantly moving. On its movement, they frequently hit the walls of its container, like the raindrops hitting the ceiling when it rains. When they hit the walls they exert a pressure, and pressure is defined as the force acting on certain area. The larger pressure the stronger the collisions with the walls and the higher the frequency of collision–the stronger the force applied on the walls. Imagine you are driving a motorcycle. While you drive you can feel the collision of the air's particle with your face. The faster you go the higher pressure. The value of air's 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 the 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 with respect to the sea level, the lower the air pressure. The weather also affects pressure, and in hot days the pressure of air is higher, whereas on a cold day pressure is lower.

Units of pressure are: atmospheres (atm), torr, pascals (Pa) or millimeters of mercury (mmHg). In order to covert pressure units, you can use the following conversion factors:

Add these conversions into your flashcard.

$$\boxed{\frac{1\ \text{atm}}{760\ \text{mmHg}}} \quad \boxed{\frac{1\ \text{torr}}{1\ \text{mmHg}}} \quad \boxed{\frac{1\ \text{atm}}{101325\ \text{Pa}}}$$

one millimeters of mercury (mmHg) is the same as 1 torr. As a note, the name torr knowledges the person who invented the barometer: Torricelli, an Italian physicist.

Sample Problem 52

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:

$$2\ \cancel{\text{atm}} \times \frac{101325\ \text{Pa}}{1\ \cancel{\text{atm}}} = 202650\text{mmHg}$$

STUDY CHECK

An oxygen sample has a pressure of 730 mmHg. Convert this value to atmospheres.

Answer: 0.96 atm.

6.2 Ideal gas law

Ideal gases are gases made of particles that do not interact with each other, differently real gases contain particles that interact, often times strongly, with each other. The temperature, pressure, volume and number of moles of a gas are not independent. They are related by a law: the ideal gas law. In this section we will introduce this law in two different forms.

`Ideal gas law in terms of moles` The ideal gas law says:

Add this formula to your flashcard.

$$\boxed{PV = nRT} \quad \text{Ideal Gas Law}$$

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 ballon with you mouth, introducing air particles into the ballon. While the number of particles inside the ballon grows, its volume will grow too. More particles will collide with the walls of the ballon and hence, the pressure inside the ballon will also increase.

Sample Problem 53

Helium gas is used to inflate blimps, scientific balloons and party balloons. What is the volume in litters of a 0.2 moles of Helium ballon at 300K and 2 atm.

SOLUTION

Analyze the Problem	Given	Asking
	$T = 300K$ $P = 2atm$ $n = 0.2mol$ $R = 0.082\frac{\text{atm}\cdot\text{L}}{\text{mol}\cdot\text{K}}$	V

Using now the ideal gases formula: $PV = nRT$, we have

$$2\cancel{atm}\cdot V = 0.2\cancel{mol}\cdot 0.082\frac{\cancel{\text{atm}}\cdot\text{L}}{\cancel{\text{mol}}\cdot\cancel{\text{K}}}\cdot 300\cancel{K}$$

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 He ballon containing 3 moles of Helium at 40C°.

Answer: 77.00 atm.

`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 $g \cdot mol^{-1}$

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 54

What is the density of Helium ballon 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:

Analyze the Problem	Given	Asking
	$T = 400K$ $P = 3atm$ $MW = 2g \cdot mol^{-1}$ $R = 0.082 \frac{\text{atm} \cdot \text{L}}{\text{mol} \cdot \text{K}}$	D

Using now the ideal gases formula in terms of density: $P \cdot MW = DRT$, we have

$$3\cancel{atm} \cdot 2 \frac{g}{\cancel{mol}} = D \cdot 0.082 \frac{\cancel{atm} \cdot \text{L}}{\cancel{mol} \cdot \cancel{K}} \cdot 400\cancel{K}$$

Solving for D we have 0.18 $g \cdot L^{-1}$.

STUDY CHECK

What is the molecular mass of a 4 $g \cdot L^{-1}$ density gas at 30C° and 5 atm.

Answer: 19.87 $g \cdot mol^{-1}$.

`STP conditions` STP conditions refer to standard pressure (1 atm) and temperature (273K) conditions. If we work at STP conditions it means pressure will be fixed at 1 atm and temperature at 273K.

Add this relation into your flashcard.

$$\boxed{\text{1 atm and 273K}} \quad \text{STP Conditions}$$

Sample Problem 55

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:

$$1\cancel{atm} \cdot V = 5\cancel{mol} \cdot 0.082\frac{\cancel{\text{atm}} \cdot \text{L}}{\text{mol} \cdot \cancel{\text{K}}} \cdot 273\cancel{K}$$

and solving for V we have a final volume of 112L.

STUDY CHECK

Calculate the grams in 4L of N_2 at STP conditions.

Answer: 5g.

6.3 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 ballon with your hand. The temperature and number of moles of the gas inside the ballon 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 ballon harder and with more frequency.

`Pressure-Volume change` If temperature and the number of moles of a 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:

$$\boxed{P_1 \cdot V_1 = P_2 \cdot V_2} \quad \text{Boyle's law}$$

where:

P_1, V_1 are the initial pressure and volume

P_2, V_2 are the final pressure and volume

`Pressure-Temperature change` Imagine you have a gas container with rigid walls, so that its volume and the number of particles inside will remain constant–think for example of an aerosol container. What will be the effect of warming up the container? You can guess that the container might explode. If we keep the number of moles and the volume of the container fixed, the pressure and temperature of the gas are related by means of Gay-Lussac's law:

$$\boxed{\frac{P_1}{T_1} = \frac{P_2}{T_2}} \quad \text{Gay-Lussac's law}$$

where:

P_1, T_1 are the initial pressure and temperature

P_2, T_2 are the final pressure and temperature

Figure 6.3: An hot air ballon will rise up as air's density decreases with temperature

`Volume-Moles change` Imagine a hot air ballon. Air comes in and out of the ballon as it is not a closed ballon. Hence the pressure on the ballon is the same as the atmospheric pressure. Also as the ballon is in contact with the air the temperature will be constant. If you inflate the ballon with hot air, the volume of the ballon and the number of moles are related by Avogadros's law:

$$\boxed{\frac{V_1}{n_1} = \frac{V_2}{n_2}} \quad \text{Gay-Lussac's law}$$

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

In order 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 the situation in which you have a 1L hot air ballon with 1 moles of a 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 moles of gas and the final estate corresponds to an unknown volume and 5 moles. Using the ideal gas formula twice you have:

$$\left.\begin{array}{r} PV_1 = n_1RT \\ \\ PV_1 = n_1RT \end{array}\right\} \quad \frac{PV_1}{PV_2} = \frac{n_1RT}{n_2RT} \tag{6.1}$$

as some of the variables of the cancel out:

$$\frac{\cancel{P}V_1}{\cancel{P}V_2} = \frac{n_1\cancel{RT}}{n_2\cancel{RT}} \tag{6.2}$$

and you end up with Avogadros' law. If you plug the numbers into the formula:

$$\frac{1L}{V_2} = \frac{1\text{ mol}}{5\text{ mol}} \tag{6.3}$$

and you get a final volume of 5L.

Sample Problem 56

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:

Analyze the Problem	Given	Asking
	$V_1 = 3L$ $P_1 = 5atm$ $P_2 = 10atm$	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:

$$\left.\begin{aligned} P_1V_1 &= nRT \\ P_2V_2 &= nRT \end{aligned}\right\} \quad \frac{P_1V_1}{P_2V_2} = \frac{nRT}{nRT} \tag{6.4}$$

Plugging the values:

$$\frac{P_1V_1}{P_2V_2} = \frac{nRT}{nRT} \tag{6.5}$$

and solving:

$$\frac{3 \cdot 5}{10 \cdot V_2} = 1 \tag{6.6}$$

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.

Answer: 2.66 atm.

Relating the different variables of a gas The questions is now, if we increase the pressure at 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 increase at fixed pressure and volume, will the temperature of the gas increase or perhaps decrease. We can answer these questions by means of the ideal gas law. If the variables that we need to relate are in the same side of the equation (e.g. P and V) then if one of the variables increase the other will decrease. Differently, If the gas variables to relate are located in 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 in 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 in opposite sides of the ideal gas law ($PV = nRT$), if P increases, T will increase as well.

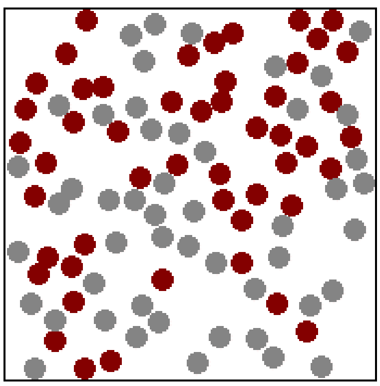


Figure 6.4: Each component in a gas mixture exert a partial pressure.

6.4 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.

Partial and total pressure Imagine you have a container with 1atm of Ar, and nother container of the same volume containing 1 atm of Ne. If you combine the containers into a single container (and 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 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:

$$\boxed{P_{Total} = p_1 + p_2 + \ldots} \quad \text{Dalton's Law}$$

Sample Problem 57

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.

Answer: 4 atm.

Discussion: Explain why a hot air balloon rises up. Furthermore, why a He-filled birthday balloon rises while if you fill it with air it does no?

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 with respect to all components

$$\boxed{X_A = \frac{n_A}{n_A + n_B + \ldots}}$$

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 58

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.

Answer: p_{Ar}=2.5 atm, p_{He}=2.5atm.

`Molar volume` If we work at STP conditions the volume of one mole of gas equals to 22.4L, and we refer to this relationship as the molar volume.

Add this relation into your flashcard.

$$\boxed{\frac{1mol}{22.4\text{ L at STP}}} \quad \text{Molar Volume}$$

This relationship allows us to carry stoichiometric calculations in chemical reaction involving gases.

`stoichiometry and gases` If you encounter chemical reactions with gases, the molar volume relation allows you to carry 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 by means of their pressure and is 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. In order 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 is 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 litters 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 59

Hydrogen gas reacts with oxygen gas to produce water vapor according to the following equation:

$$\underset{2\text{L}}{2\,\text{H}_2(\text{g}) + \text{O}_2(\text{g})} \longrightarrow \underset{\text{x mol}}{2\,\text{H}_2\text{O}(\text{g})}$$

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:

$$x = 2L\,\cancel{\text{L of O}_2} \times \frac{2\,\cancel{\text{L of H}_2\text{O}}}{2\,\cancel{\text{L of O}_2}} \times \frac{1 \text{ mol of H}_2\text{O}}{22.4\cancel{\text{L of H}_2\text{O}}} = 0.178 \text{ mol of H}_2\text{O}.$$

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 $g \cdot mol^{-1}$) gas to produce ammonia at STP conditions according to the following equation:

$$\underset{2\text{L}}{3\,\text{H}_2(\text{g}) + \text{N}_2(\text{g})} \longrightarrow \underset{\text{x g}}{2\,\text{NH}_3(\text{g})}$$

Calculate the number of grams of ammonia produced from 0.5L of nitrogen.

Answer: 5g.

6.5 Real gases and the kinetic molecular theory of gases

Until now we have discusses ideal gases. Ideal gases represent very diluted gases in which the gas particles do not see each other, as they are apart. The particles of an ideal gas are very minute without a volume. The collisions between the particles and the walls of the container are elastic–this means the gas molecules do not loose any energy. As you can imagine, no gas is an ideal gas, as this is just an ideal model. This section will cover the properties of real gases, in which the gas particles interact among themselves. We will also discuss the kinetic molecular theory with a particular emphasis on particle average velocity–this is technically called root mean square velocity, v_{RMS}.

`Van der Waals equation for real gases` When we take into account the fact that the particles of a gas interact with each other the formula of the ideal gases do not work anymore. Instead, we can use the Van der Waals equation for real gases that functions in a very similar way.

$$\boxed{P = \frac{nRT}{V - nb} - \frac{n^2 a}{V^2}}$$ Van der Waals equation

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 $0.082\frac{\text{atm}\cdot\text{L}}{\text{mol}\cdot\text{K}}$

a and b are the Van der Waals constants in units of $atm \cdot L^2 \cdot mol^{-2}$ and $L \cdot mol$

`Van der Waals constants` The Van der Waals constants a and b represent the degree of interaction between the molecules of a gas. The larger these values the more interactions exists between the gas particles. For example, for He we have $a = 0.0341 atm \cdot L^2 \cdot mol^{-2}$ and $b = 0.0237 L \cdot mol$, whereas for H_2O we have $a = 5.46 atm \cdot L^2 \cdot mol^{-2}$ and $b = 0.0305 L \cdot mol$. Comparing the values of a for both gases–or the values of b–we can conclude that the interaction between the particles of He are very weak but the interactions between the particles of H_2O are stronger.

Sample Problem 60

Calculate the pressure of 0.2 moles of water vapor at 500K and 6 atm occupying a volume of 0.1L, using: (a) the ideal gas formula and (b) the van der Waals formula $a = 5.46 atm \cdot L^2 \cdot mol^{-2}$ and $b = 0.0305 L \cdot mol$.

SOLUTION

We will use the ideal gas formula first, given the number of moles (n=0.2 mol), temperature (T=500K), pressure (p=6 atm) and the volume (V=0.1L).

$$P = \frac{nRT}{V} = \frac{0.2 \times 0.083 \times 500}{0.1} = 82 atm$$

Now, using the van der Waals formula:

$$P = \frac{nRT}{V - nb} - \frac{n^2 a}{V^2} = \frac{0.2 \times 0.083 \times 500}{0.1 - 0.2 \times 0.0305} - \frac{0.2^2 \times 5.46}{0.1^2} = 65 atm$$

Both values are very different and this is consistent with the fact that water vapor does not behave as an ideal gas.

STUDY CHECK

Calculate the pressure of 0.9 moles of ammonia gas at 900K and 9 atm occupying a volume of 0.1L, using: (a) the ideal gas formula and (b) the van der Waals formula $a = 4.17 atm \cdot L^2 \cdot mol^{-2}$ and $b = 0.0371 L \cdot mol$.

Answer: 664 atm;659 atm.

`Kinetic theory of gases` The kinetic theory of gases is a model that explain 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:

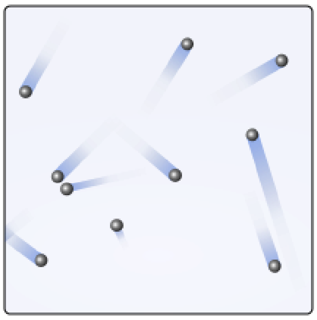


Figure 6.5: The particles of a gas move fast.

- 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 collide with the walls of its container exerting pressure.
- Gas particles are far apart from each other, barely interacting.
- The average kinetic energy of the particles of a gas (this is the energy of the particles due to movement) is proportional to the temperature of the gas.

By means of 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 its 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 explain 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 an experimental law; 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 describe 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 the root mean square velocity with temperature is:

$$\boxed{v_{RMS} = \sqrt{\frac{3000RT}{MW}}} \quad \text{root mean square velocity formula}$$

where:

MW is the molecular weight of the gas in $g \cdot mol^{-1}$

T is the temperature of the gas in K

R is the constant of the gas in energy units 8.314 $J \cdot K^{-1} \cdot mole^{-1}$

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 inverssely proportional to the molecular weight of the gas–the heavier the mass the lower velocity.

Sample Problem 61

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=20$g \cdot mol^{-1}$), CO_2(MW=44$g \cdot mol^{-1}$) and H_2O(MW=18$g \cdot mol^{-1}$). 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°.

Answer: 625m/s.

CHAPTER 6

Gases And Its Properties

1. The pressure of a gas is 2 atm. This value in mmHg is:

(a) 760 mmHg
(b) 1520 mmHg
(c) 1.52×10^{-3} mmHg
(d) 1.6×10^{3} mmHg

Ans: (b)

2. The pressure of a gas is 900 mmHg. This value in torr is:

(a) 900 torr
(b) 760 torr
(c) 1×10^{-3} torr
(d) 7×10^{5} torr

Ans: (d)

3. The pressure of a gas is 3000 Pa. This value in atm is:

(a) 3000 atm
(b) 2 atm
(c) 2.96×10^{-2} atm
(d) 3×10^{8} atm

Ans: (c)

4. The kinetic molecular theory of gases assumes that:

(a) gas particles interact with each other
(b) gas particles have large sizes
(c) particles move slowly
(d) gas particles move randomly

Ans: (d)

Ideal Gas Law

5. 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 talk.

(a) 45 moles
(b) 1 moles
(c) 5 moles
(d) 0.45 moles

Ans: (c)

6. A 4 moles sample of gas at 400K has a pressure of 10 atm. Calculate the volume of the sample:

(a) 10 L
(b) 3 L
(c) 13.12 L
(d) 1915 L

Ans: (d)

7. Eighteen moles of a gas in a 11L container at 400K exert a pressure of 3 atm. Calculate the molar mass of the gas.

(a) $1\ g \cdot mol^{-1}$
(b) $18\ g \cdot mol^{-1}$
(c) $161\ g \cdot mol^{-1}$
(d) $6\ g \cdot mol^{-1}$

Ans: (b)

8. Calculate the volume of a 4 moles of Ar at STP conditions.

Ans: 89.6L

9. Calculate the volume of a 4 moles of Ne at STP conditions.

Ans: 89.6L

10. Calculate the moles of gas in 3L of Ar at STP conditions.

Ans: 0.13mol

Change of Gas Properties

11. A sample of a gas at 400K and 12 atm is cooled in the same container to 200K. Calculate the new pressure

(a) 15 atm
(b) 6 atm
(c) 24 atm
(d) 0.2 atm

Ans: (b)

12. 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.

(a) 1.8 L
(b) 3.15 L
(c) 2.8 L
(d) 1 L

Ans: (c)

13. If the pressure of a gas increases, at fixed temperature and moles, its volume....

(a) increases
(b) decreases
(c) does not change

Ans: (b)

Mixture of Gases and Gas Stoichiometry

14. 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?

Ans: 1402torr

15. 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.

Ans: N_2 608 mmHg, O_2 171.6 mmHg

16. 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.

Ans: 0.017 mol

Real Gases and the Kinetic Molecular Theory of Gases

17. What is the pressure in atm of 1 mol of He at 600K in a 1L container: (a) Using the ideal gas law and (b) Using the real gas law given $a = 0.0342 atm \cdot L^2 \cdot mol^{-2}$ and $b = 0.0237 L \cdot mol$.

Ans: 49.2 atm; 50.36 atm

18. What is the rms speed of O2 at STP?.

Ans: Ans: 481.9 m/s

REVIEW PART C

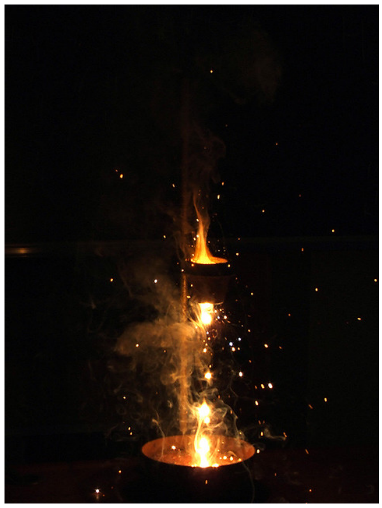
1. Indicate the redox number of: $H\underline{N}O_3$
   (a) V (c) III (e) II
   (b) I (d) IV
2. Indicate the redox number of: $\underline{N}O_2$
   (a) V (c) III (e) $-II$
   (b) I (d) IV
3. Indicate the redox number of: $\underline{N}_2H_5^+$
   (a) V (c) III (e) $-II$
   (b) I (d) IV
4. Classify the following chemical as: H_2SO_4
   (a) non electrolyte
   (b) strong electrolyte
   (c) weak electrolyte
5. Classify the following chemical as: CH_3COOH
   (a) non electrolyte
   (b) strong electrolyte
   (c) weak electrolyte
6. Classify the following chemical as: HF
   (a) non electrolyte
   (b) strong electrolyte
   (c) weak electrolyte
7. The volume of a gas with a pressure of 1.2 atm increases from 1.0 L to 4.0 L. What is the final pressure of the gas, assuming constant temperature?
   (a) 1.2 atm (c) 3.3 atm (e) 1.0 atm
   (b) 0.30 atm (d) 4.8 atm
8. What volume would a 0.250 mol sample of H_2 gas occupy, if it had a which has a pressure of 1.70 atm, and a temperature of 35 °C?
   (a) 0.269 L (c) 1.25 L (e) 283 L
   (b) 0.423 L (d) 3.72 L
9. At STP, what is the volume of 4.50 mols of nitrogen gas?
   (a) 167 L (c) 101 L (e) 1230 L
   (b) 3420 L (d) 60.7 L
10. How many L of O2 gas at STP, are needed to react with 15.0 g of Na?
$$4\,Na(s) + O_2(g) \longrightarrow 2\,Na_2O(s)$$
   (a) 3.65 L (c) 7.30 L (e) 32.0 L
   (b) 14.6 L (d) 22.4 L
11. When 3 L of HCl reacts, how many L of H_2 gas are formed at STP conditions?
$$Zn(s) + 2\,HCl(aq) \longrightarrow H_2(g) + ZnCl_2(aq)$$
   (a) 1.5 L (c) 8.32 L (e) 0.382 L
   (b) 0.120 L (d) 22.4 L
12. A 1.20-L container contains 1.10 g of an unknown gas at STP. What is the molcular weight of the unknown gas?
   (a) 1.32 g/mol (c) 22.4 g/mol (e) 1.10 g/mol
   (b) 0.917 g/mol (d) 20.5 g/mol
13. What is the concentration, in mass percent, of a solution prepared from 50.0 g NaCl and 150.0 g of water?
   (a) 0.250% (c) 40.0% (e) 3.00%
   (b) 33.3% (d) 25.0%
14. What is the molarity of a solution which contains 58.5 g of sodium chloride dissolved in 0.500 L of solution?
   (a) 0.500 M (c) 1.50 M (e) 4.00 M
   (b) 1.00 M (d) 2.00 M
15. What volume of 0.10 M NaOH can be prepared from 250. mL of 0.30 M NaOH?
   (a) 0.075 L (c) 0.75 L (e) 750 L
   (b) 0.25 L (d) 0.083 L
16. How many mL of a NaCl 1.7 M solution contains 4g of NaCl (MW=58.44 $g \cdot mol^{-1}$).
17. What is the final volume, in milliliters, when 5.00 mL of a 3% NaCl solution is diluted to provide a 1% NaCl solution.
18. Write the net-ionic equations for the following reaction:
$$HCl(l) + KOH(aq) \longrightarrow KCl(aq) + H_2O(l)$$
19. What is the pressure in atm of 3 mole of NH3 gas at 1600K in a 5L container? (a) Using the ideal gas law; (b) Using the real gas law (a(NH3)=4.17; b(NH3)=0.04).

Answers:

1. (a)	**2.** (d)	**3.** (c)	**4.** (b)
5. (c)	**6.** (c)	**7.** (b)	**8.** (d)
9. (c)	**10.** (c)	**11.** (a)	**12.** (d)
13. (d)	**14.** (d)	**15.** (c)	**16.** 40mL
17. 15mL	**18.** $H^+_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(l)}$	**19.** (a) 78atm; (b) 79.1atm	

PART D

7
Thermochemistry



ENERGY involves many aspects of our everyday life. The chemical reactions in our body consume or release energy as we walk, study, and even breath. We also use energy in our homes to warm our food or turn on the lights, and to 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 of energy. Similarly, the burning of fossil fuels such as oil, propane, or gasoline provides enough energy to maintain our homes. Some reactions produce energy whereas others release energy. On the other hand, how do we measure the energy release or consumed in a reaction? This chapter will answer this and other questions as it covers different aspects of thermochemistry that involves the interaction between chemistry and energy. You will learn about temperature, heat and how to compute the energy exchanged during a chemical reaction.

GOALS

1. Convert heat to temperature rise
2. Carry calorimetric calculations
3. Use the enthalpy table
4. Compute enthalpy changes
5. Apply Hess's Law to compute enthalpy

7.1 Energy & temperature

When you are running, walking, dancing, or thinking, you are using energy to do work. In fact, 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.

Figure 7.1: water on a dam has potential energy

`Potential & Kinetic Energy` Energy can be classified as potential energy or kinetic energy. Kinetic energy is the energy of motion and any object that is moving 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 releases when it collide 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 reservoir has potential energy. When the water goes over the dam, the potential energy is converted to kinetic energy.

Figure 7.2: a bullet has kinetic energy

`Heat` Heat refers to thermal energy, which is associated with the motion of particles in a substance. A frozen pizza feels cold because the particles in the pizza are moving very slowly. As the pizza receives heat, the motions of the particles increase, and the pizza becomes warm. Eventually the particles have enough energy to make the pizza hot and ready to eat. When a substance receives heat it gets warmer and it raises its temperature, whereas if it looses heat it gets cooler and its temperature decreases.

`Discussion:` What do you think about renewable energy? List three benefits and three inconvenients of renewable sources of energy.

`Energy units` Two different units of energy are often employed: calories (cal) and joules (J). One can transform calories to joules and joules to calories using the following conversion factor:

$$\boxed{1cal = 4.184J} \quad \text{or} \quad \boxed{\frac{1cal}{4.184J}} \quad \text{or} \quad \boxed{\frac{4.184J}{1cal}} \tag{7.1}$$

Add Equation 7.1 to your flashcard.

Sample Problem 62

Convert the following energy values:

(a) 50000 cal to Kcal (b) 48000 J to cal

SOLUTION

We will use the conversion factor for kilo and the relationship between calorie and joule:
(a) $50000cal \times \frac{Kcal}{1000cal} = 50Kcal$; (b) $48000J \times \frac{1cal}{4.184J} = 11472.2cal$.

STUDY CHECK

Convert the following energy units:

(a) 200 cal to Kcal (b) 7000 J to cal

Answer: (a) 0.2Kcal; (b) 1673 cal.

Figure 7.3: Some thermometers has different temperature scales.

`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), Fahrenheit (°F) or Kelvin (K). If you need to convert temperature units from Fahrenheit to celsius or from celsius to Fahrenheit you need to use the following formula on the left. Differently, if you need to convert between Kelvin and celsius you need to use the formula on the right:

$$\boxed{T_F = 1.8T_C + 32} \qquad \boxed{T_K = T_C + 273} \tag{7.2}$$

Add Equation 7.2 to your flashcard.

Sample Problem 63

Convert 25 °C to °F.

SOLUTION

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

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

2 **Step two:** use the formula $T_F = 1.8T_C + 32$ to convert from °C to °F.

$$T_F = 1.8\,T_c + 32$$

$25°$

3 **Step three:** solve for $T_F = 1.8 \times 25 + 32 = 77°F$.

STUDY CHECK

Convert 200°C to K.

Answer: 473K.

Figure 7.4: Metals have low specific heat, which means they heat up very fast.

7.2 From energy to temperature

Materials can absorb heat. Think about a pizza in your over, or a cup of milk in the microwaves. These substances receive heat from the oven or in form of microwaves and they become hot. Heat transforms in an increase of temperature. Some substances like metals are able to increase its temperature very quickly with a small amount of heat received, whereas others need a larger amount of heat to rise up its temperature. Think about why you use oil to deep fried food? Why not using water? First of all, oil can rise its temperature very quickly and on top of that it does not boil easily.

Material	Specific Heat $J/g°C$
H_2O (l)	4.184
H_2O (s)	2.010
ethyl alcohol	2.460
vegetable oil	2.010
Cu	0.385
Au	0.129
Fe (s)	0.444

`Specific Heat` Specific heat c_e is a property of each material that indicates the energy required to rise its temperature. 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 in order 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 rise 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 in order to increase the temperature of a gram of water in $1°C$, whereas we need to give 0.2 cal in order to increase the temperature of a gram of aluminum in $1°C$. Why are these two numbers so different? The answer is because 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. Mind the specific heat if water is a well know value that you need to be familiar with:

$$\boxed{c_e^{H_2O} = 4.184J/g°C} \quad \boxed{c_e^{H_2O} = 1cal/g°C} \tag{7.3}$$

Figure 7.5: Oil has low specific heat. This allows achieving hight temperatures with small heat intakes.

`From heat to temperature` When a materials receives heat, that heat normally becomes temperature and the temperature of the material increases. For example, if you warm milk in a microwaves, that milk warms up from room temperature until a higher temperature. How to estimate the temperature increase given the heat received? We can use the following formula:

Add Equation 7.4 to your flashcard.

$$\boxed{Q = m \cdot c_e \cdot (T_{Final} - T_{Initial})} \tag{7.4}$$

where m is the mass of material in grams, c_e is the specific heat of the material which can either be given in $cal/g°C$ or $J/g°C$, $T_{Initial}$ is the initial temperature and T_{Final} is the final temperature in $°C$. Q is the amount of heat received, either in cal or J.

Sample Problem 64

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

SOLUTION

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

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

2 **Step two:** use the formula $Q = m \cdot c_e \cdot (T_{Final} - T_{Initial})$ to temperature increase to heat absorbed:

$25°C$

$$Q = m \cdot c_e \cdot (T_{Final} - T_{Initial})$$

$45.2g$ $\quad$ $0.214\frac{cal}{g \cdot °C}$ $\quad$ $50°C$

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 ($c_e = 0.0308\frac{cal}{g \cdot °}C$) if its temperature rises from 25°C to 100°C.

Answer: $Q = 231cal$.

1370

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 65

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.

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

$$\underbrace{Q}_{100cal} = \underbrace{m}_{50g} \cdot \underbrace{c_e}_{0.214\frac{cal}{g\cdot°C}} \cdot (T_{Final} - \overbrace{T_{Initial}}^{25°C})$$

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.

Answer: $T_{Final} = 65°C$.

7.3 Calorimetry

How do we measure the heat exchanged in chemical reactions? And more importantly, how do we know if a chemical reaction produces or consumes energy? The answer to these questions is: by means of a tool called calorimeter. Some calorimeters are very fancy and expensive, whereas others are as simple as a coffee cup. And both are used to measure the energy exchanges–sometimes produced, other consumed–in chemical reactions. This section will show you how to carry calorimetric calculations.

`The calorimeter` A calorimeter is device used to measure the energy exchanged in chemical reactions. Calorimetry is the science that measures heat exchange by using calorimeters. In essence is a closed system such as a coffee cup that does not let the heat come though its walls. There are tow different types of calorimeters. A constant-pressure calorimeter is the simplest of all calorimeters and is called constant-pressure as the pressure inside the calorimeter is constant and equal to the atmospheric pressure. Think of a double coffee cup covered with a lit. Inside this cup a chemical reaction occurs in a liquid phase. If the reaction produces any gases as the cup if just covered with a lit, the pressure will always be equal to the atmospheric pressure. Differently, a constant-volume calorimeter–also know as a bomb calorimeter–is a more complex and costly calorimeter in which normally gas phase reactions occur. Examples of the use of a constant-volume calorimeter are the calculation of the energy value of food–that is how they know how many calories are there in a bag of chips. In constant-volume calorimeter, the pressure is not constant but the volume of the calorimeter is.

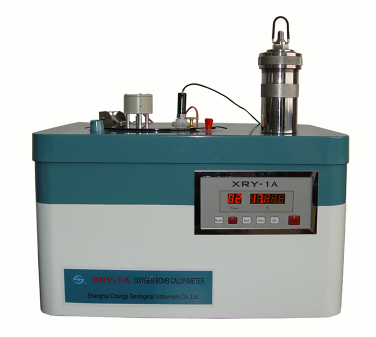


Figure 7.6: A constant-volume calorimeter is also called bomb calorimeter.

`Calorimetry formula` The overall use of a calorimeter is to calculate the energy exchanges in chemical reactions–this is called the heat of reaction $\Delta Q_{reaction}$. In essence, there is a single formula to cary calorimetric calculations and all calorimeters function in the same way. Let us use a reaction that produced heat as an example. Inside the calorimeter, you introduced a given sample of a compound and a reaction happens producing heat (first contribution), the heat generated goes from the reaction to a container with water that warms up (second contribution). At the same time, the walls of the calorimeter also absorb certain amount of heat (third contribution). The formula used in calorimetry is the following:

$$0 = n \cdot \Delta Q_{reaction} + \Delta Q_{water} + \Delta Q_{walls}$$

where:

$n \cdot \Delta Q_{reaction}$ is the hear exchanged due to a chemical reaction

ΔQ_{water} is the heat received or released by water

ΔQ_{walls} is the heat absorbed by the walls

The water contribution is given by the heat formula given above and the walls contribution is the result of the effective heat capacity of the calorimeter, $c_{e,Cal}$. The larger the effective heat capacity of the calorimeter the more heat it absorbs from the reaction. Heat capacity is the same idea as specific heat but in different units. After we plug these two contribution into the formula above we arrive to the calorimetry formula:

Add this formula to your flashcard.

$$\boxed{0 = n \cdot \Delta Q_{reaction} + m \cdot c_{e,Water} \cdot (T_F - T_I) + c_{e,Cal} \cdot (T_F - T_I)}$$ Calorimetry formula

where:

$\Delta Q_{reaction}$ is the hear exchanged due to a chemical reaction in J/mol

n is the number of moles reacted inside the calorimeter

m is the mas of water contained in the calorimeter

$c_{e,Water}$ is the specific heat absorbed of water:$4.184 J/g°C$

$c_{e,Cal}$ is the heat capacity of the calorimeter also known as calorimeter factor

T_F is the final temperature of water in the calorimeter

T_I is the initial temperature of water in the calorimeter

Figure 7.7: A constant-pressure calorimeter is also referred as a coffee-cup calorimeter and it consists simply in two cups with a lit and a thermometer.

Exothermic and endothermic reactions Some reactions release heat and are called exothermic. Others absorb heat and are called endothermic. Think for example the combustion of the gas in a cooking stove, it produces gas and hence the chemical reaction happening is exothermic. Differently, if you cook bread, it needs heat to rise. Of if you melt an ice cube you need to give energy to the cube so that it becomes water. These are examples of endothermic reactions. Endothermic reactions have positive $\Delta Q_{reaction}$ whereas exothermic reactions have negative $\Delta Q_{reaction}$.

Sample Problem 66

A 3 mol-sample of a chemical is combusted in a calorimeter with 10g of water and a heat capacity of $10KJ/°C$. Calculate the heat of reaction knowing that the initial temperature of the water inside the calorimeter is $25°C$ and the final $40°C$.

SOLUTION

Analyze the Problem	Given	Asking
	$n = 3mol$ $m = 10g$ $T_F = 40°C$ $T_I = 25°C$ $c_{e,C} = 10000J/°C$ $c_{e,W} = 4.184J/g°C$	$\Delta Q_{reaction}$

We have all date needed to solve the calorimetry formula. We have the moles of chemical inside the calorimeter, the heat capacity of the calorimeter, the initial and final temperature of water, and the amount of water. Mind that the specific heat of water is always given and you need to remember the value. Also and more importantly mind that the units of the heat capacity of the calorimeter are $KJ/°C$, whereas the units of the specific heat of water are $J/g°C$ and hence, we need to convert KJ into J; that is the reason we use $10000J/°C$ as the heat capacity of the calorimeter. Plugging all values into the calorimetry formula we have:

$$0 = 3mol \cdot \Delta Q_{reaction} + 10g \cdot 4.184J/g°C \cdot (40°C - 25°C) \\ +10000J/°C \cdot (40°C - 25°C)$$

Solving for $\Delta Q_{reaction}$ we obtain $-50209J/mol$ that is the same as $-50.209KJ/mol$. As the value is negative, it means that the reaction produced energy and hence is exothermic.

STUDY CHECK

A 2 mol-sample of a chemical reacts in a bomb calorimeter with 20g of water and a heat capacity of $11KJ/°C$. Calculate the heat of reaction knowing that the temperature of water inside the calorimeter rises $10°C$.

Answer: $-55KJ/mol$.

7.4 Enthalpy

Standard state	ΔH_f° (KJ/mol)
H_2(g)	0
O_2(g)	0
N_2(g)	0
Cl_2(g)	0
Fe(s)	0
Al(s)	0
C(graphite)	0
P_4(s)	0

In the last section we have seen that when a chemical reaction proceeds it exchanges energy with the surroundings. This energy can me measured in many different conditions. When it is measured at constant pressure–these are regular conditions in chemistry, think about a reaction happening at a beaker–this energy change has a different name: it is called enthalpy and is represented with the symbol ΔH_f°. In this section we will cover the different types of enthalpies depending on the type of reaction–formation or reaction–and we will find out how to compute the enthalpy change for a reaction using tables of standard enthalpies.

`Table of standard enthalpies` Enthalpies are tabulated in tables of standard enthalpies. The term standard refers to standard pressure conditions (1 atm) and is indicates by a degree sign on the top right side. Let us learn how to use this table. If you look for the standard enthalpy of C–an element–from the table you might find several values. The values of graphite carbon is $\Delta H_f^\circ = 0 KJ/mol$. Differently, the values for diamond carbon is different than zero, being $\Delta H_f^\circ = 1.0 KJ/mol$. Similarly, the value for gas carbon is not zero also, being $\Delta H_f^\circ = 716.67 KJ/mol$. This is because the nature state of carbon is in the form of graphite. That is, the most common way in which we find carbon in nature is in the form of graphite and not diamond or gas. Let us find the standard enthalpy for molecular nitrogen, N_2(g)–another element. If you look into the table you will find a value of $\Delta H_f^\circ = 0 KJ/mol$, again because the natural state of nitrogen is in the form of gas N_2. What is the standard enthalpy of gas hydrogen, H_2? If you look in the table, the value is also $\Delta H_f^\circ = 0 KJ/mol$. The rule of thumb is: elements on its natural state have zero H_f°. Below we will explain more about the meaning of natural state. Now, look for the standard entropy of carbon monoxide gas, CO(g). The value should not be zero, as carbon dioxide is not an element and is made of two different types of atoms. Indeed, in the table we find $\Delta H_f^\circ(\mathrm{CO}) = -110.5 KJ/mol$.

`Natural state of an element` The natural state of an element is the most stable state in which we find this element in nature. For example, Aluminum, it's natural state is not as a liquid or as a gas, is a s metallic solid. That is the reason why $\Delta H_f^\circ(\mathrm{Al(g)}) = 314 KJ/mol$, whereas $\Delta H_f^\circ(\mathrm{Al(s)}) = 0 KJ/mol$. In general, for metals, its natural state is solid. For non-metallic elements, such as hydrogen, oxygen, nitrogen, fluorine, chlorine, its natural state is in the form of a diatomic gas molecule. For example, $\Delta H_f^\circ(\mathrm{H_2(s)}) = 0 KJ/mol$, $\Delta H_f^\circ(\mathrm{N_2(s)}) = 0 KJ/mol$ or $\Delta H_f^\circ(\mathrm{O_2(s)}) = 0 KJ/mol$. For the case of carbon, its natural state is graphite, $\Delta H_f^\circ(\mathrm{C(s)_{graphite}}) = 0 KJ/mol$.

`Standard enthalpy of molecules` Molecules such as H_2O or NO have standard entropy different than zero. Mind that molecules are not elements, and hence are made of different elements.

`Standard enthalpy change for a reaction` In order to calculate the standard enthalpy for a reaction you need to use the following formula:

Add this formula to your flashcard.

$$\boxed{\Delta H_R^\circ = \Delta H_{products}^\circ - \Delta H_{reactants}^\circ}$$ Enthalpy change

where:

ΔH°_R is the standard enthalpy change of the reaction

$\Delta H^\circ_{products}$ is the standard enthalpy of all products

$\Delta H^\circ_{reactants}$ is the standard enthalpy of all reactants

Now, imagine we need to calculate the change of standard enthalpy for the following reaction:

$$N_2(g) + O_2(g) \longrightarrow 2\,NO(g)$$

In the case of the reaction above, we need to locate three enthalpies from the table: $\Delta H^\circ_f(N_2(g))$, $\Delta H^\circ_f(O_2(g))$, $\Delta H^\circ_f(NO(g))$. If we look in the tables, you will find the values: $\Delta H^\circ_f(N_2(g)) = 0KJ/mol$ and $\Delta H^\circ_f(O_2(g)) = 0KJ/mol$. This makes sense, as these are the natural states of nitrogen and oxygen. Differently, $\Delta H^\circ_f(NO(g)) = 90.29KJ/mol$. Now, in order to compute ΔH°_R we need to take into account that the reaction involves two NO molecules. Let us set up the formula:

$$\begin{aligned}\Delta H^\circ_R = \Delta H^\circ_{products} - \Delta H^\circ_{reactants} &= \Big(2 \cdot \Delta H^\circ_f(NO(g))\Big) - \Big(\Delta H^\circ_f(N_2(g)) + \Delta H^\circ_f(O_2(g))\Big) \\ &= \Big(2 \cdot 90.29\Big) - \Big(0 + 0\Big) = 181KJ\end{aligned}$$

This reaction is endothermic, that means it consumes energy. Let's work on another example and calculate ΔH°_R for the reaction:

$$2\,NO(g) + O_2(g) \longrightarrow 2\,NO_2(g)$$

We need to locate three enthalpies from the table: $\Delta H^\circ_f(NO(g))$, $\Delta H^\circ_f(O_2(g))$, $\Delta H^\circ_f(NO_2(g))$. If you locate these values in the table you will see $\Delta H^\circ_f(O_2(g)) = 0KJ/mol$, whereas $\Delta H^\circ_f(NO(g)) = 90.29KJ/mol$ and $\Delta H^\circ_f(NO_2(g)) = 33.2KJ/mol$. Using the formula for ΔH°_R we have:

$$\begin{aligned}\Delta H^\circ_R = \Delta H^\circ_{products} - \Delta H^\circ_{reactants} &= \Big(2 \cdot \Delta H^\circ_f(NO_2(g))\Big) - \Big(2 \cdot \Delta H^\circ_f(NO(g)) + \Delta H^\circ_f(O_2(g))\Big) \\ &= \Big(2 \cdot 33.2\Big) - \Big(2 \cdot 90.29 + 0\Big) = -114KJ\end{aligned}$$

This reaction is exothermic and releases heat.

How to indicate ΔH°_R in a reaction Normally in chemical reaction the value of ΔH°_R can be written in two different ways. You can see the enthalpy added in the reaction as a reactant or product or you can find the entropy written on the right side of the reaction. For example, in the first example above–the endothermic reaction–you can write: $N_2(g) + O_2(g) \longrightarrow 2\,NO(g) + 181$ KJ or you can find: $N_2(g) + O_2(g) \longrightarrow 2\,NO(g)$ $\Delta H^\circ_R = 181KJ$. For the second example above, the exothermic reaction, you can indicate the enthalpy like: $2\,NO(g) + O_2(g) + 114$ KJ $\longrightarrow 2\,NO_2(g)$ or you can find:

$$N_2(g) + O_2(g) \longrightarrow 2\,NO(g) \qquad \Delta H^\circ_R = -114KJ.$$

Note that in exothermic reaction, the enthalpy value is indicated as a product as the reaction produces heat, whereas in endothermic reaction it is indicated as a reactant, as these reactions consume heat.

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 related energy and the moles of reactants and products.

For the exothermic reaction:

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(g) \qquad \Delta H = -572KJ.$$

the moles of hydrogen are related to heat as:

$$\boxed{\frac{\text{2 moles of } H_2}{\text{-572 KJ}} \text{ or } \frac{\text{-572 KJ}}{\text{2 moles of } H_2}}$$

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 reactant or products into heat.

Sample Problem 67

Hydrogen reacts with nitrogen to produce ammonia (NH_3) according to the following reaction

$$3\,H_2(g) + N_2(g) \longrightarrow 2\,NH_3(g) + 92\,KJ$$

Calculate: (a) the enthalpy of reaction; (b) indicate whether the reaction is endo or exothermic; (c) calculate the heat produced 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

Calculate the number of hydrogen moles needed to generate -200KJ.

Answer: 6.5 moles.

Sample Problem 68

Using the enthalpy table, calculate ΔH°_R for the following reactions: (a) $4\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(l)$; (b) $3\,H_2(g) + N_2(g) \longrightarrow 2\,NH_3(g)$; (c) $2\,Al(s) + 3\,Cl_2(g) \longrightarrow 2\,AlCl_3(s)$.

SOLUTION

In order to answer all questions, we need a set of ΔH°_f values: $\Delta H^\circ_f(H_2(g))$, $\Delta H^\circ_f(O_2(g))$, $\Delta H^\circ_f(N_2(g))$, $\Delta H^\circ_f(Al(s))$ are all zero, whereas $\Delta H^\circ_f(H_2O(l)) = -285.8KJ/mol$, $\Delta H^\circ_f(NH_3(g)) = -45.0KJ/mol$ and $\Delta H^\circ_f(AlCl_3(s)) = -705.63KJ/mol$. For the first example, we have:

$$\begin{aligned}\Delta H^\circ_R &= \Big(2\cdot\Delta H^\circ_f(H_2O(l))\Big) - \Big(4\cdot\Delta H^\circ_f(H_2(g)) + \Delta H^\circ_f(O_2(g))\Big)\\ &= \Big(2\cdot -285.8\Big) - \Big(4\cdot 0 + 0\Big) = -572KJ\end{aligned}$$

For the second example:

$$\begin{aligned}\Delta H^\circ_R &= \Big(2\cdot\Delta H^\circ_f(NH_3(g))\Big) - \Big(2\cdot\Delta H^\circ_f(Al(s)) + 3\cdot\Delta H^\circ_f(Cl_2(g))\Big)\\ &= \Big(2\cdot -45\Big) - \Big(2\cdot 0 + 3\cdot 0\Big) = -90KJ\end{aligned}$$

Finally, for the last reaction:

$$\begin{aligned}\Delta H^\circ_R &= \Big(2\cdot\Delta H^\circ_f(AlCl_3(s))\Big) - \Big(3\cdot\Delta H^\circ_f(H_2(g)) + \Delta H^\circ_f(N_2(g))\Big)\\ &= \Big(2\cdot -705.63\Big) - \Big(3\cdot 0 + 0\Big) = -1411KJ\end{aligned}$$

STUDY CHECK

Using the enthalpy table, calculate ΔH°_R for the following reaction: $Fe_2O_3(s) + 3\,CO(g) \longrightarrow 2\,Fe(s) + 3\,CO_2(g)$.

Answer: $-25KJ$.

ΔH°_R and ΔH°_f Consider the following two reactions:

$$N_2(g) + O_2(g) \longrightarrow 2\,NO(g) \qquad \Delta H^\circ_f = -114KJ.$$

$$Fe_2O_3(s) + 3\,CO(g) \longrightarrow 2\,Fe(s) + 3\,CO_2(g) \qquad \Delta H^\circ_R = -25KJ.$$

The first example represents a formation reaction and thus the enthalpy is labeled as ΔH°_f. This is called standard enthalpy of formation. A formation reaction always have the elements on its standard state as reactants. Think of $N_2(g)$ and $H_2(g)$. The enthalpy of formation for both is zero as they represent nitrogen and hydrogen on their natural states. Differently, the second reaction is not a formation reaction, as the reactants are not elements on its natural state. Think of $CO(g)$ or $Fe_2O_3(s)$, their enthalpy is not zero as they do not represent any element on its natural state. That is the reason, the second enthalpy is labeled as ΔH°_R. This is called standard enthalpy of reaction.

7.5 Hess's Law: Manipulating reaction enthalpies

In the previous section we relied on a table el standard enthalpies of formation in order to compute enthalpy changes in general reaction. This enthalpy change ΔH_R° is related to the heat exchanged in the reaction. In this section we will not use the tables of enthalpy anymore. Imagine you do not have access to this table. And we will find alternative ways to predict ΔH_f° given a series of reactions with know enthalpies. In short you will have to identify the enthalpies that are zero–the enthalpies corresponding to an element on its natural state– and set up an equation that helps you find out the missing enthalpy.

`Reverting reactions` Imagine they give you the following reaction:

$$N_2(g) + O_2(g) \longrightarrow 2\,NO(g) \qquad \Delta H_1^\circ = -114KJ$$

and you need to calculate the enthalpy change for this other reaction:

$$2\,NO(g) \longrightarrow N_2(g) + O_2(g) \qquad \Delta H_2^\circ$$

If you compare both reaction you will see the second reaction equals to the first reaction but reverted. If you revert a reaction, the enthalpy change changes sign. Therefore, $\Delta H_2^\circ = 114KJ$.

`Timing reactions by a number` Imagine they give you the following reaction:

$$N_2(g) + O_2(g) \longrightarrow 2\,NO(g) \qquad \Delta H_1^\circ = -114KJ$$

and you need to calculate the enthalpy change for this other reaction:

$$2\,N_2(g) + 2\,O_2(g) \longrightarrow 4\,NO(g) \qquad \Delta H_2^\circ$$

If you compare both reaction you will see the second reaction equals to the first reaction timed by two. If you time a reaction by two, the enthalpy change should also be timed by two. Therefore, $\Delta H_2^\circ = 2 \cdot -114 = -228KJ$.

`Combining reactions` Imagine they give you the following two reactions:

$$\begin{array}{rcll} C(s) + O_2(g) & \longrightarrow & CO_2(g) & \Delta H_1 \\ H_2(g) + \frac{1}{2}\,O_2(g) & \longrightarrow & H_2O(l) & \Delta H_2 \end{array}$$

and the ask the enthalpy change for the following reaction:

$$C(s) + H_2(g) + \tfrac{3}{2}\,O_2(g) \longrightarrow CO_2(g) + H_2O(l) \qquad \Delta H_3^\circ$$

If you look closely to the last reaction, you will see it results from adding the first two reactions, so that:

$$\begin{array}{rcll} C(s) + O_2(g) & \longrightarrow & CO_2(g) & \Delta H_1 = -393KJ \\ H_2(g) + \frac{1}{2}\,O_2(g) & \longrightarrow & H_2O(l) & \Delta H_2 = -286KJ \\ \hline \text{Sum: } C(s) + H_2(g) + \frac{3}{2}\,O_2(g) & \longrightarrow & CO_2(g) + H_2O(l) & \Delta H_3 \end{array}$$

Therefore, $\Delta H_3 = \Delta H_1 + \Delta H_2 = -679KJ$. When adding two chemical reactions the resulting enthalpy is the result of adding the enthalpy of both reactions.

Sample Problem 69

Calculate the enthalpy for this reaction:

$$2\,C(s) + H_2(g) \longrightarrow C_2H_2(g) \qquad \Delta H_4^\circ$$

Given the following thermochemical equations:

$$\begin{aligned} C_2H_2(g) + \tfrac{5}{2}\,O_2(g) &\longrightarrow 2\,CO_2(g) + H_2O(l) & \Delta H_1 &= -1299.5KJ \\ C(s) + O_2(g) &\longrightarrow CO_2(g) & \Delta H_2 &= -393.5KJ \\ H_2(g) + \tfrac{1}{2}\,O_2(g) &\longrightarrow H_2O(l) & \Delta H_3 &= -285.8KJ \end{aligned}$$

SOLUTION

In order to get the enthalpy for reaction (4) we will have to combine reactions (1), (2) and (3), by adding, subtracting, or multiplying by a number so that the results adds up to reaction (4). A trick to do this is compare molecule by molecule in reaction (4) and see in which reaction we can find the same one. For example, reaction (4) contains 2C(s) in the reactant side. C(s) can also be found in (2) also as reactant. However, in (2) C(s) is not timed by 2. There we will use two times reaction (2):

$$2\,C(s) + 2\,O_2(g) \longrightarrow 2\,CO_2(g) \qquad 2 \cdot \Delta H_2^\circ = -787$$

Reaction (4) also contains H_2(g), which can be found in (3). There we will use (3) as it is:

$$H_2(g) + \tfrac{1}{2}\,O_2(g) \longrightarrow H_2O(l) \qquad \Delta H_3^\circ = -285.8KJ$$

Reaction (4) also contains C_2H_2(g) as a product. We can find the same chemical in (1) but as a reactant. There we will have to invert (1):

$$2\,CO_2(g) + H_2O(l) \longrightarrow C_2H_2(g) + \tfrac{5}{2}\,O_2(g) \qquad -1 \cdot \Delta H_1 = 1299.5KJ$$

If we add the three previous reactions, we have:

$$\begin{aligned} 2\,C(s) + 2\,O_2(g) &\longrightarrow 2\,CO_2(g) & 2 \cdot \Delta H_2 &= -787KJ \\ H_2(g) + \tfrac{1}{2}\,O_2(g) &\longrightarrow H_2O(l) & \Delta H_3 &= -285.8KJ \\ 2\,CO_2(g) + H_2O(l) &\longrightarrow C_2H_2(g) + \tfrac{5}{2}\,O_2(g) & -1 \cdot \Delta H_1 &= 1299.5KJ \\ \hline \text{Sum: } 2\,C(s) + H_2(g) &\longrightarrow C_2H_2(g) & \Delta H_4 & \end{aligned}$$

Therefore in the enthalpy for the reaction (4) will be:

$$\Delta H_4^\circ = 2 \cdot \Delta H_2 + \Delta H_3 - 1 \cdot \Delta H_1 = 226.7KJ$$

STUDY CHECK

Calculate the enthalpy for this reaction:

$$CS_2(l) + 3\,O_2(g) \longrightarrow CO_2(g) + 2\,SO_2(g)$$

Given the following thermochemical equations:

$$\begin{aligned} C(s) + O_2(g) &\longrightarrow CO_2(g) & \Delta H_1 &= -393.5KJ \\ S(s) + O_2(g) &\longrightarrow SO_2(g) & \Delta H_2 &= -296.8KJ \\ C(s) + 2\,S(s) &\longrightarrow CS_2(l) & \Delta H_3 &= 87.9KJ \end{aligned}$$

Answer: $-1075KJ$.

Standard entropy table at 1atm and 298K in KJ/mol.

Substance	ΔH_f°	Substance	ΔH_f°	Substance	ΔH_f°
Aluminum					
Al(s)	0	$AlCl_3$(s)	-705.63	Al_2O_3(s)	-1675.5
$Al(OH)_3$(s)	-1277	$Al_2(SO_4)_3$(s)	-3440	NH_3(Aq)	-80.8
NH_3(g)	-46.1	NH_4NO_3(s)	-365.6	Al(g)	314
Barium					
$BaCl_2$(s)	-858.6	$BaCO_3$(s)	-1213	$Ba(OH)_2$(s)	-944.7
BaO(s)	-548.1	$BaSO_4$(s)	-1473.2	$BaSO_4$(s)	-1473.2
Boron					
Solid BCl_3(s)	-402.96				
Bromine					
Br_2(l)	0	Aqueous Br^-	-121	Br(g)	111.884
Br_2(g)	30.91	BrF_3 (g)	-255.60	HBr(g)	-36.29
Cadmium					
Cd(s)	0	CdO(s)	-258	$Cd(OH)_2$(s)	-561
CdS(s)	-162	$CdSO_4$(s)	-935		
Calcium					
Ca(s)	0	Ca(g)	178.2	$Ca_2{}^+$(g)	1925.90
CaC_2(s)	-59.8	$CaCO_3$(s)	-1206.9	$CaCl_2$(s)	-795.8
$CaCl_2$(aq)	-877.3	$Ca_3(PO_4)_2$(s)	-4132	CaF_2(s)	-1219.6
CaH_2(s)	-186.2	$Ca(OH)_2$(s)	-986.09	$Ca(OH)_2$(aq)	-1002.82
CaO(s)	-635.09	$CaSO_4$(s)	-1434.52	CaS(s)	-482.4
$CaSiO_3$(s)	-1630				
Caesium					
Cs(s)	0	Cs(g)	76.50	Cs(l)	2.09
Cs^+(g)	457.964	CsCl(s)	-443.04		
Carbon					
$C_{graphite}$(s)	0	d $C_{diamond}$(s)	1.9	C(g)	716.67
CO_2(g)	-393.509	CS_2(l)	89.41	CS_2(g)	116.7
CO(g)	-110.525	$COCl_2$(g)	-218.8	CO_2(aq)	-419.26
$HCO_3{}^-$(aq)	-689.93	$CO_3{}^{2-}$(aq)	-675.23		
Chlorine					
Cl(g)	121.7	Cl^-(aq)	-167.2	Cl_2(g)	0
Chromium					
Cr(s)	0				
Copper					
Cu(s)	0	CuO(s)	-155.2	$CuSO_4$(aq)	-769.98
Fluorine					
F_2(g)	0				
Hydrogen					
H(g)	218	H_2(g)	0	H_2O(g)	-241.818
H_2O(l)	-285.8	H^+(aq)	0	OH^-(aq)	-230
H_2O_2	-187.8	H_3PO_4(l)	-1288	HCN(g)	130.5
HBr(l)	-36.3	HCl(g)	-92.30	HCl(aq)	-167.2
HF(g)	-273.3	HI(g)	26.5		
Iodine					
I_2(s)	0	I_2(g)	62.438	I_2(aq)	23
I^-(aq)	-55				

(cont.) Standard entropy table at 1atm and 298K.

Substance	ΔH_f°	Substance	ΔH_f°	Substance	ΔH_f°
Iron					
Fe	0	Fe_3C(s)	5.4	$FeCO_3$(s)	-750.6
$FeCl_3$(s)	-399.4	FeO(s)	-272	Fe_3O_4(s)	-1118.4
Fe_2O_3(s)	-824.2	$FeSO_4$(s)	-929	$Fe_2(SO_4)_3$(s)	-2583
FeS(s)	-102	FeS_2(s)	-178		
Lead					
Pb(s)	0	PbO_2(s)	-277	PbS(s)	-100
$PbSO_4$(s)	-920	$Pb(NO_3)_2$(s)	-452	PbO(s)	-276.6
Magnesium					
Mg(s)	0	$Mg_2{}^+$(aq)	-466.85	$MgCO_3$(s)	-1095.7
MgO(s)	-601.6	$MgSO_4$(s)	-1278.2	$MgCl_2$(s)	-641.8
Manganese					
Mn(s)	0	MnO (s)	-384.9	MnO_2(s)	-519.7
Mn_2O_3(s)	-971	Mn_3O_4(s)	-1387	$MnO_4{}^-$(aq)	-543
Mercury					
HgO(s)	-90.83	HgS(s)	-58.2		
Nitrogen					
N_2(g)	0	NH_3(aq)	-80.8	NH_3(g)	-45.90
NH_4Cl	-314.55	NO_2(g)	33.2	N_2O(g)	82.05
NO(g)	90.29	N_2O_4(g)	9.16	N_2O_5(s)	-43.1
Oxygen					
O(g)	249	O_2(g)	0	O_3(g)	143
Phosphorus					
P_4(s)	0	P_{red}(s)	-17.4	P_{black}(s)	-39.3
PCl_3(l)	-319.7	PCl_3(g)	-278	PCl_5(s)	-440
PCl_5(g)	-321	P_2O_5(s)	-1505.5		
Potassium					
KBr(s)	-392.2	K_2CO_3(s)	-1150	$KClO_3$(s)	-391.4
KCl(s)	-436.68	KF(s)	-562.6	K_2O(s)	-363
$KClO_4$(s)	-430.12				
Silicon					
Si(g)	368.2	SiC(s)	-74.4	$SiCl_4$(l)	-640.1
SiO_2(s)	-910.86				
Silver					
AgBr(s)	-99.5	AgCl(s)	-127.01	AgI(s)	-62.4
Ag_2O(s)	-31.1	Ag_2S(s)	-31.8		
Sodium					
Na(s)	0	Na(g)	+107.5	$NaHCO_3$(s)	-950.8
Na_2CO_3(s)	-1130.77	NaCl(aq)	-407.27	NaCl(s)	-411.12
NaF(s)	-569.0	NaOH(aq)	-469.15	NaOH(s)	-425.93
Na_2O(s)	-414.2				
Sulfur					
$S_{8monoclinic}$(s)	0.3	$S_{8rhombic}$(s)	0	H_2S(g)	-20.63
SO_2(g)	-296.84	SO_3(g)	-395.7	H_2SO_4(l)	-814
Titanium					
Ti(s)	0	Ti(g)	468	$TiCl_4$(g)	-763.2
$TiCl_4$(l)	-804.2	TiO_2(s)	-944.7		
Zinc					
Zn(g)	130.7	$ZnCl_2$(s)	-415.1	ZnO(s)	-348.0

CHAPTER 7

Energy and temperature

1. The energy associated with the motion of particles in a substance is called

   (a) kinetic energy
   (b) potential energy
   (c) heat
   (d) chemical energy
   (e) none of the above

   Ans: (a)

2. The energy stored in the chemical bonds of a carbohydrate molecule is

   (a) kinetic energy
   (b) potential energy
   (c) heat
   (d) chemical energy
   (e) none of the above

   Ans: (d)

3. The energy stored in heigh is

   (a) kinetic energy
   (b) potential energy
   (c) heat
   (d) chemical energy
   (e) none of the above

   Ans: (b)

4. 650J is the same amount of energy as

   (a) 155 cal
   (b) 2720 cal
   (c) 650 cal
   (d) 1550 cal
   (e) 2.72 cal

   Ans: (a)

5. 3.25 kcal is the same amount of energy as

   (a) 3.25 J
   (b) 0.777 J
   (c) 777 J
   (d) 13598 J
   (e) 13.6 J

   Ans: (d)

6. A temperature of 41$°F$ is the same as

   (a) 5$°C$
   (b) 310$°C$
   (c) -9$°C$
   (d) 16$°C$
   (e) 42$°C$

   Ans: (a)

7. A temperature of 20$°C$ is the same as

   (a) -22$°F$
   (b) 68$°F$
   (c) 43$°F$
   (d) 239$°F$
   (e) 94$°F$

   Ans: (b)

8. A temperature of 300K is the same as

   (a) 27$°C$
   (b) 20$°C$
   (c) 45$°C$
   (d) 50$°C$
   (e) 90$°C$

   Ans: (a)

From Energy To Temperature

9. The specific heat of a substance is the amount of heat needed to

   (a) change 1 g of the substance from the solid to the liquid state.
   (b) raise the temperature of 1 g of the substance by 1$°C$.
   (c) change 1 g of the substance from the liquid to the solid state.
   (d) convert 1 g of a liquid to gas.
   (e) convert 1 g of a solid to a gas.

   Ans: (b)

10. 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 $0.108 cal/g°C$.

    (a) 38 cal
    (b) 1.1 cal
    (c) 3.8 cal
    (d) 93 cal
    (e) 130 cal

    Ans: (a)

11. What is the final temperature of a 35 g sample of iron at 25$°C$ after receiving 50cal? Iron has a specific heat of $0.108 cal/g°C$.

    (a) 25$°C$
    (b) 35$°C$
    (c) 38$°C$
    (d) 50$°C$
    (e) 27$°C$

    Ans: (c)

12. 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 $0.2 cal/g°C$.

    (a) 55$°C$
    (b) 25$°C$
    (c) 40$°C$
    (d) 45$°C$
    (e) 50$°C$

    Ans: (d)

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

    (a) $0.04 cal/g°C$
    (b) $0.14 cal/g°C$
    (c) $1.04 cal/g°C$
    (d) $2.0 cal/g°C$
    (e) $0.34 cal/g°C$

    Ans: (a)

Calorimetry

14. A 3 moles sample of C(s) is burned in a constant-volume calorimeter containing 40g of water. The temperature inside the calorimeter increases from 25.0$°C$ to 25.89 $°C$. The calorimeter constant is 9.90 $KJ/°C$. Calculate the molar enthalpy of combustion of the sample.

    Ans: -3KJ/mol

15. A 10 grams sample of fructose (MW=180$g \cdot mol^{-1}$) is burned in a constant-volume calorimeter containing 50g of water. The temperature inside the calorimeter increases 7$°C$. The calorimeter constant is 10.8 $KJ/°C$. Calculate the molar enthalpy of combustion of the sample.

Ans: -1387KJ/mol

16. When a 0.09-g sample of trinitrotoluene (TNT, MW=213$g \cdot mol^{-1}$), is burned in a bomb calorimeter, the temperature increases from 23.5 $°C$ to 27.1$°C$. The heat capacity of the calorimeter is 400 J/$°C$, and it contains 100 mL of water. How much heat was produced by the combustion of the TNT sample?

Ans: -6973KJ/mol

17. You warm up 10 g of a solid to 200$°C$ and add it to 60 g of water in a coffee-cup calorimeter. The water temperature changes from 24$°C$ to 29$°C$. Calculate the specific heat of the solid. You can consider negligible the heat capacity of the coffee-cup calorimeter.

Ans: 0.17J/$°C$

ENTHALPY

18. The following reaction:

$$C_6H_{12}O_6(s) + 6\,O_2(g) \longrightarrow 6\,CO_2(g) + 6\,H_2O(g) + 2800KJ$$

(a) is endothermic
(b) is exothermic
(c) uses energy
(d) Consumes energy
(e) None of the above

Ans: (b)

19. The following reaction:

$$B_2O_3(s) + 3\,H_2O(g) + 2035KJ \longrightarrow 3\,O_2(g) + B_2H_6(g)$$

(a) is endothermic
(b) is exothermic
(c) produces energy
(d) None of the above

Ans: (a)

20. The following reaction:

$$C_6H_{12}O_6(s) + 6\,O_2(g) \longrightarrow 6\,CO_2(g) + 6\,H_2O(g) + 2800KJ$$

Fill the conversion factor:

$$\frac{\_\_\_\_\_\ \text{moles of } O_2}{\text{-2800 KJ}}$$

Ans: 6

21. In the following combustion reaction:

$$C_6H_{12}O_6(s) + 6\,O_2(g) \longrightarrow 6\,CO_2(g) + 6\,H_2O(g) + 2800KJ$$

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.

Ans: 8400KJ

22. Calculate the enthalpy of reaction for:

$$2\,OF_2(g) \longrightarrow O_2(g) + 2\,F_2(g)$$

given:

$$\Delta H^\circ_f(OF2(g)) = 24.5KJ$$

Ans: -49KJ/mol

23. Calculate the enthalpy of reaction for:

$$2\,ClF(g) + O_2(g) \longrightarrow Cl_2O(g) + OF_2(g)$$

given:

$$\Delta H^\circ_f(ClF(g)) = -56KJ$$
$$\Delta H^\circ_f(Cl_2O(g)) = 88KJ$$
$$\Delta H^\circ_f(OF_2(g)) = 25KJ$$

Ans: 225KJ/mol

24. Calculate the enthalpy of reaction for:

$$ClF_3(g) + O_2(g) \longrightarrow Cl_2O(g) + \tfrac{3}{2}\,OF_2(g)$$

given:

$$\Delta H^\circ_f(ClF_3(g)) = -156KJ$$
$$\Delta H^\circ_f(Cl_2O(g)) = 88KJ$$
$$\Delta H^\circ_f(OF_2(g)) = 25KJ$$

Ans: 281KJ/mol

HESS'S LAW

25. Using the following reactions:

$$2\,OF_2(g) \longrightarrow O_2(g) + 2\,F_2(g) \qquad \Delta H_1 = -49KJ$$
$$2\,ClF(g) + O_2(g) \longrightarrow Cl_2O(g) + OF_2(g) \qquad \Delta H_2 = 225KJ$$
$$ClF_3(g) + O_2(g) \longrightarrow \tfrac{1}{2}\,Cl_2O(g) + \tfrac{3}{2}\,OF_2(g) \qquad \Delta H_3 = 324KJ$$

Determine the enthalpy change for:

$$ClF(g) + F_2(g) \longrightarrow ClF_3(g)$$

Ans: -187KJ/mol

26. Using the following reactions:

$$N_2(g) + 3\,H_2(g) \longrightarrow 2\,NH_3(g) \qquad \Delta H_1 = -92KJ$$
$$C(s) + 2\,H_2(g) \longrightarrow CH_4(g) \qquad \Delta H_2 = -75KJ$$
$$H_2(g) + 2\,C(s) + N_2(g) \longrightarrow 2\,HCN(g) \qquad \Delta H_3 = 270KJ$$

Determine the enthalpy change for:

$$CH_4(g) + NH_3(g) \longrightarrow HCN(g) + 3\,H_2(g)$$

Ans: 256KJ/mol

27. Using the following reactions:

$$3\,C(s) + 3\,H_2(g) + \tfrac{1}{2}\,O_2(g) \longrightarrow C_3H_6O(l) \qquad \Delta H_1 = -285KJ$$
$$C(s) + O_2(g) \longrightarrow CO_2(g) \qquad \Delta H_2 = -394KJ$$
$$H_2(g) + \tfrac{1}{2}\,O_2(g) \longrightarrow H_2O(l) \qquad \Delta H_3 = -286KJ$$

Determine the enthalpy change for:

$$C_3H_6O(l) + 4\,O_2(g) \longrightarrow 3\,CO_2(g) + 3\,H_2O(l)$$

Ans: -1755KJ/mol

8

Solids and liquids

THERE are three different states of the matter: solid, liquid and gas. At this point, we have studied the properties of gases and liquid solutions. We have not encountered yet solids or pure liquids. This chapter fully deals with the properties of solids and liquids. Liquids have indeed very peculiar properties and this chapter will cover–among other–the vapor pressure. Liquids are not isolated; they are normally in contact with the atmosphere. The liquid molecules which are closer to the air can scape forming a vapor; this vapor exerts certain pressure. This vapor is what you feel, for example, when the weather is very humid. Finally, this chapter covers the idea of intermolecular forces. The molecules of an ideal gas are independent from each other. This means they do not see each other at all–they do not interact with each other. Differently, the molecules of liquids and solids interact with each other by means of stronger force that act between molecules–these are called intermolecular forces. The properties of this forces will help you understand why some liquids boil at higher temperature than others or some solids have higher melting point.

GOALS

1. Identify intermolecular forces
2. Identify different types of solids
3. Identify units cells
4. Calculate density of solids
5. Calculate vapor pressure

8.1 Intermolecular forces

The atoms of solids or liquids are connected by means of chemical bonds, forming molecules. Bonds are forces within molecules. This bonds can be ionic or covalent depending on the nature of the elements that form the compound. At the same time, the molecules of a liquid or solid compound interact with each other by means of intermolecular forces. The word intermolecular means between molecules. This section describe the three different types of intermolecular forces existing, its nature and intensity.

Dispersion forces All molecules are made of atoms and atoms contain electrons. The electrons of an atom are distributed homogeneously in the atom. This means that there are no negative or positive regions in an atom. On the other hand when two atoms get close together, the presence of each other alternated the electron distribution creation what is known as a temporary dipole. This temporary dipole results in dispersion forces. These forces exist in any chemical, as all chemicals contain atoms that can polarize temporarily. The larger the atomic number the stronger these forces, as in general the more electrons the stronger will be the temporary dipole.

Dipole-Dipole forces Dipole-dipole forces exist only in polar compounds, being the result of the permanent dipole moments existing in polar molecules. Examples of polar compounds are: HCl or H_2O. This forces are in general stronger than dispersion forces.

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 molecule with hydrogen bonds is HF or NH_3.

Sample Problem 70

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.

Answer: all have dispersion, only NH_3, HF has dipole and only NH_3, HF has H-bonds.

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. In order 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 by means of weak dispersion forces. Imagine now boiling water. Water is polar and water has O–H bonds, hence water molecules interact by means of 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 71

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	✓	✓
Dipole-Dipole	✓	✓
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.

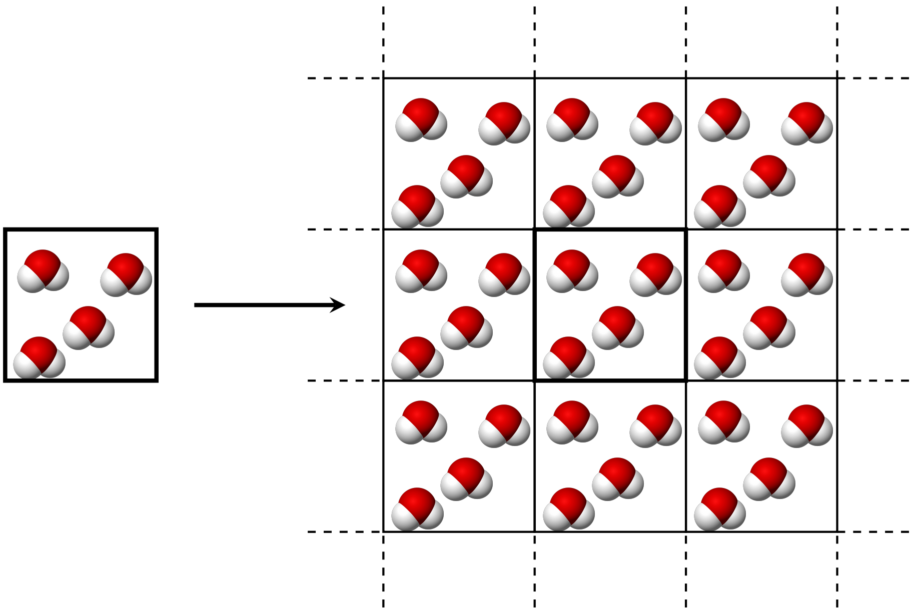
Answer: BP(CH_3F)>BP(CH_4).

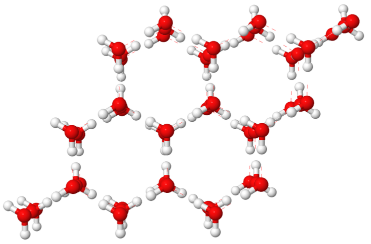
8.2 The solid state

What makes solids unique in comparison to liquids and gases? They answer is their structure. There are two main different types of solids: crystalline solids and amorphous solids. Crystalline solids are made of atoms or molecules periodically, regularly, arranged in the three dimensions of the space. Examples of a crystalline solid are table salt or sugar. Amorphous solids have disordered structures. An example of an amorphous solid is window glass. This section will focus on the properties of crystalline solids as their periodicity makes their properties easier to study.

Figure 8.1: Cooking salt, NaCl is a crystalline solid, whereas glass, SiO_4 is an amorphous solid.

`Crystalline lattice: the unit cell` The structure of crystalline solids is periodic. The term does not refer to periodic in time, but in periodic in space. Hence, the structure of crystalline solids is the result of the repetition of a small piece of the structure in the space. The overall structure is called *crystalline lattice*. Here an example of a very simple two dimensional lattice. In this lattice, the central box is repeated infinitely in two directions of the space generating a lattice.





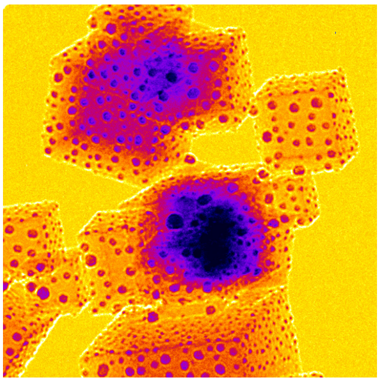


Figure 8.2: ice, $H_2O(s)$ is a molecular solid made of water molecules, whereas magnesium oxide MgO is an ionic solid made of Mg^{2+} cations and O^{2-} anions. Gold is an atomic solid made of gold atoms.

As the lattice is made of repetition, the smallest repeating unit is called the *unit cell*. Therefore, simply with the unit cell one can generate the whole crystal lattice by repeating the unit cell in the tree dimensions. Therefore, it is unnecessary to study the crystalline whole lattice as the unit cell is enough to understand many properties of crystalline solids such as their density. In the following we will study in more detail the properties of crystalline solids and some of the most common unit cells.

`Types of crystalline solids` Examples of crystalline solids are: sugar and table salt. These two solids have very different constitutions. Table salt is made of ions: Na^+ and Cl^-. Sugar is made of molecules. We say NaCl is an ionic solid, whereas sugar is a *molecular solid*. Other examples of *ionic solids*: MgO, CaF_2. Other examples of *molecular solids*: ice which is made of water molecules. A third type of crystalline solids are called *atomic solids*, as they are made of atoms. Think of metallic iron or graphite. Both are atomic solids made of atoms, Fe and C. Overall, molecular solids are made or molecules–often times covalent molecules–whereas ionic solids are made of ions and result from ionic compounds. Finally, atomic solids are made of atoms. In the following we will study more about a specific type of atomic solids: metallic solids.

Sample Problem 72

Classify the following solids as ionic, molecular or atomic: diamond, dry ice (CO_2), iron and CaF_2.

	diamond	CO_2	Fe	CaF_2
Molecular				
Ionic				
Atomic				

SOLUTION

In general ionic solids correspond to ionic compounds and molecular solids correspond to covalent compounds. Therefore, dry ice should be a molecular solid and CaF_2 and ionic solid. Iron and diamond are both made of atoms and hence they are atomic compounds.

	diamond	CO_2	Fe	CaF_2
Molecular	✗	✓	✗	✗
Ionic	✗	✗	✗	✓
Atomic	✓	✗	✓	✗

STUDY CHECK

Classify the following solids as ionic, molecular or atomic: silver, graphite, $CaCO_3$ and $NH_3(s)$.

Answer: atomic, atomic, ionic and molecular.

8.3 Metals and ionic solids

Among the different types of crystalline solids, metals and ionic solids are very important. This section will cover the structure of metallic solids like gold or iron and ionic solids like sodium chloride.

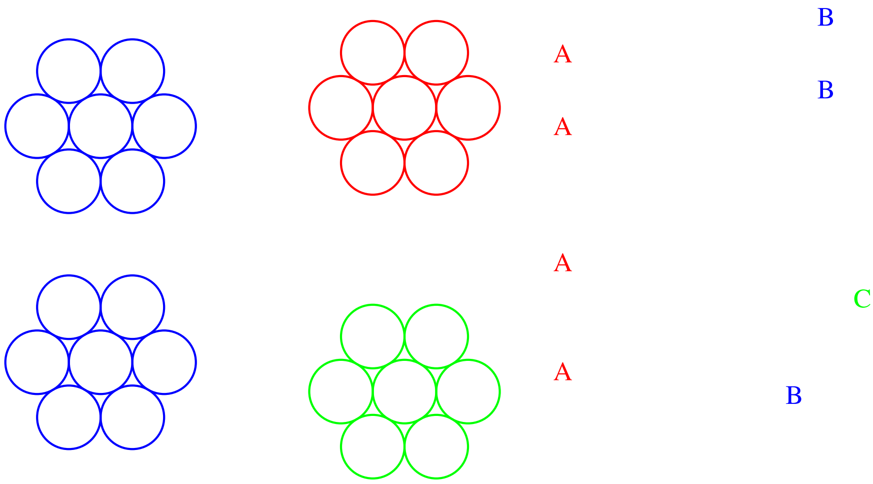


Figure 8.3: Different unit cells–such as sc, fcc or hcc–result from the different packing of metal layers. The simplest cubic structure results of placing two compact metal layer one on top of the other. When we place the second layer on the holes of the first layer we achieve a AB layer distribution and the resulting unit cell is called face centered unit cell. If we use a AB packing and now add a new layer on the holes of the second layer we obtain a ABC packing and the resulting unit cell is called hexagonal close cell.

`Closed packing of metals` Metallic solids are the results of the packing of metal atoms the space. Picture a single layer of spheres all packed together. The most compact way to pack a layer of spheres is the situation in which one sphere is surrounded by six other spheres. This is called the closest packing. In this situation, each three spheres are connected by means of an indentation or dimple. Now the question is how do we pack a second layer on top of the first later. We can simply place the second layer just on top of the same positions of the first layer. This would lead to a simple cubic packing and this type of packing is not the most compact packing. The unit cell resulting from this packing is called *simple cubic*. We could also pack the second layer on the indentations of the second layer. As this second later would be located at different locations as the second layer, we call this second layer B and the first layer A. Now let us think about adding a third layer. There are two possible locations for this new layer; you can locate this third layer on top of the first layer layer–with this I mean on the same location as the first layer–this would lead to an ABAB packing, as the first layer is the same as the third layer. The unit cell resulting from this packing is called *hexagonal close cell, (hcp)*. Differently, you can either locate the third layer on the indentations of the second layer leading to a ABC layer packing. In this packing the third layer is now now the same as the first layer. The unit cell resulting from this packing is called *face centered cubic, (fcc)*.

Crystal structure	Atoms per cell
Inside	1
sc	1
bcc	2
fcc	4

Figure 8.4: Atoms per unit cell

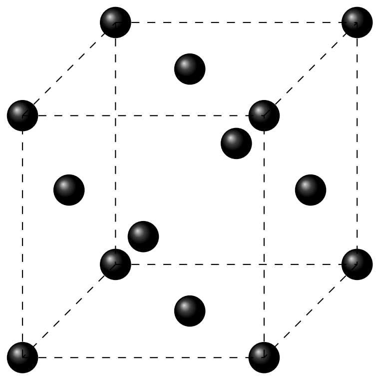
Location	Sharing factor, f
Inside	1
Vertex	$\frac{1}{8}$
Faces	$\frac{1}{2}$
Edges	$\frac{1}{4}$

Figure 8.5: Sharing factors for different locations of a cubic unit cell

`Atom sharing in unit cells` Before we cover the different metallic units cells let us talk about atom sharing. Think about a cubit unit cell, that is a cube with one sphere (atom) in every corner of the cube. The whole lattice is produced by repeating the unit cell on the three dimensions. Hence, every corner of the cube is shared among other corders. This means, every corner–containing an atom–shares that atom with all units cells connected to that corner. Therefore, those atoms in the corner are not whole part of a single unit cell and they are shares. Every corner of a cube is shared among eight other cubes. Imagine pilling numerous boxes in layers. Every corner of each box is shared by three other boxes in the same plane and by four boxes on the plane on top–that is a total of eight boxes. They way you need to think of the different atoms in a single unit cell, is that they are shared depending on their location. As we discussed, corners of a cubic unit cell are shared by a total of 8 others unit cells. Atoms that belong to a face of a unit cell are shared by two unit cells. Atoms that are inside a unit cell fully belong to a single unit cell and they are not shared. Atoms that belong to a edge of the cube–an edge is the line that connects two vertexes of a cube–are shared by four units cells.

Sample Problem 73

The following structure is called face centered unit cell. This is a cubit unit cell with one atom in each corner of the cell and atoms also in the facets of the cell. Calculate the number of atoms in the unit cell:



SOLUTION

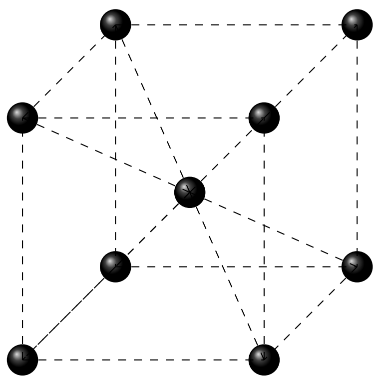
If you count the number of spheres in the drawing you might think the cell contains fourteen atoms. However, this is not true, as each sphere is shared by other unit cells. Remember each location of the unit cell counts as a fraction. If an atom is fully inside in the cell–not in the vertexes, neither in the faces or sides–the sharing factor is one. If an atom belongs to a vertex, the sharing factor is $\frac{1}{8}$. Atoms in a face has a sharing factor of $\frac{1}{2}$ and atoms in the edges have a sharing factor of $\frac{1}{4}$.

Location	Sharing Factor, f	Number of atoms, N	$f \times N$
Corner	$\frac{1}{8}$	8	1
Faces	$\frac{1}{2}$	6	3

By multiplying the number of atoms in each location by the sharing factor and adding we obtain the total number of atoms in the cell. Overall, this unit cell has four atoms:

STUDY CHECK

The following structure is called simple body centered unit cell. This is a cubit unit cell with one atom in each corner of the cell and an atom also in the center of cell. Calculate the number of atoms in the unit cell:



Answer: 2.

`Metal unit cells` Here we will cover three different metal unit cells, all cubic cells. First, the simple cubic unit cell, with an atom each of the vertexes of the cell. This is the less compact unit cell with one atom per unit cell. Second, the body-centered unit cell is a cubic unit cell with atoms in the vertex of the cell and a single atom in the center of the cell. This cell has two atoms per unit cell. Third, the face-centered unit cell, with atoms in the vertex of the cell and also on the faces of the cell, on the sides of the cube. This is the most compact unit cell, with four atoms per cell. In the following image you can manipulate a face-centered cubic cell.

Crystal structure	relation between c and r
sc	$c = r$
bcc	$c = \frac{4}{\sqrt{3}r}$
fcc	$c = \sqrt{8}r$

Figure 8.6: Relationship between cell parameter c and atomic size r for different types of unit cells

`Cell parameter` Cubic unit cells have the shape of a cube and hence all side of the cube have the same length. This length is called cell parameter c. Unit cells with large cell parameter have more spacing between atoms. The opposite is true for cells with smaller cell parameter. The cell parameter of a unit cell is related to the atomic radius. Let us analyze the case of a face-centered unit cell. In each side of the cell, in each face, we have four atoms in the vertexes and one in the center of the face. Of course these atoms do not belong only to this unit cell. However, if we symbolically cut the atoms in the face we can see the relation between the radius of the atom and the unit cell. The edges of the cell does not correspond to any cell parameter. However, the line that connect the bottom part with the opposite top part corresponds to a specific number of cell parameters, as the atoms are touching in this direction. In particular this distance is $4r$. Using Pythagoras theorem we have: $c^2 + c^2 = (4r)^2$. Therefore, $c = \sqrt{8}r$.

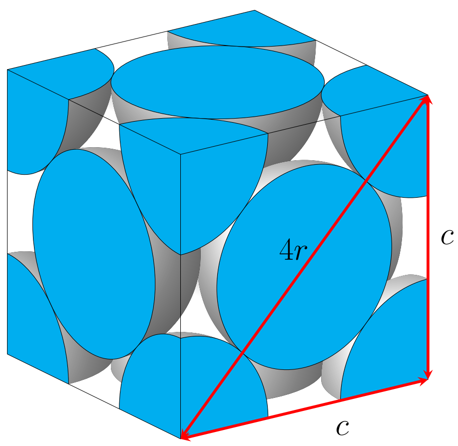
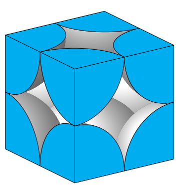


Figure 8.7: This is space-filling representation of a fcc unit cell.

Sample Problem 74

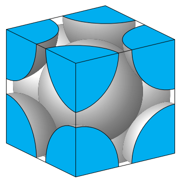
For the following unit cell, calculate the relationship between the cell parameter and the atomic radius.



SOLUTION

For this unit cell, the atoms in the bottom part are touching. Hence, the cell parameter should be related to the atomic radius. In particular, two half atoms occupy the same distance as the cell parameter, so $c = r$.

STUDY CHECK



Answer: $c = \frac{4}{\sqrt{3}c}$.

Metal density Different metals have different density. The value for density will depend on the cell parameter but also on the compacity of the unit cell, the more compact the unit cell the more atoms per cell and hence the more density. The formula that relates density with cell parameter and atoms per cell is:

Add this formula to your flashcard.

$$\boxed{d = \frac{N \cdot AW}{c^3 \cdot 6.023 \times 10^{-7}}}$$ Metallic density formula

where:

d is the density in $g \cdot ml^{-1}$

N is the number of atoms per unit cell

6.023×10^{-7} is related to the conversion between atoms and grams

AW is the atomic weight of the metal

c is the cell parameter in pm

Sample Problem 75

Calculate density of iron ($AW = 55.845 g \cdot mol^{-1}$) knowing this is a bcc metal with cell parameter is 286pm.

SOLUTION

We know that iron is a bcc metal and hence it has two atoms per unit cell. Also we know its atomic weight $AW = 55.845 g \cdot mol^{-1}$ and the cell parameter $c = 286$pm. Using the metallic density formula:

$$d = \frac{2 \cdot 55.845}{286^3 \cdot 6.023 \times 10^{-7}} = \frac{111.69}{14.09} = 7.93 g \cdot ml^{-1}$$

STUDY CHECK

Calculate density of gold (AW=$196.96 g \cdot mol^{-1}$) knowing this is a fcc metal with cell parameter is 406pm.

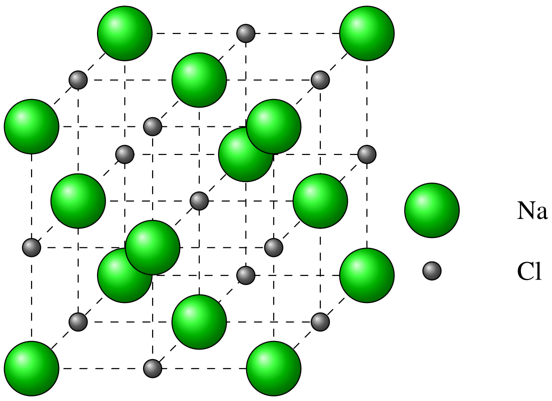
Answer: $19.54 g \cdot ml^{-1}$.

Ionic solids Ionic solids have high melting point and they are typically hard. They also do not conduct the electricity in solid form. An example of an ionic solid is NaCl.

The structure of NaCl and many other ionic solids results from the superposition of two different compact lattices–this is the reason these are called binary solids as they are made of two units–and each lattice is superimposed. Normally, the largest ion (Na^+) forms a packed arrangement such as fcc or ccp, and the smallest ion (Cl^-) resides on the holes of the lattice. Here we will care about constructing the formula of the unit cell, such as NaCl by counting the atoms in the unit cell.

Sample Problem 76

Calculate the formula for the following unit cell



SOLUTION

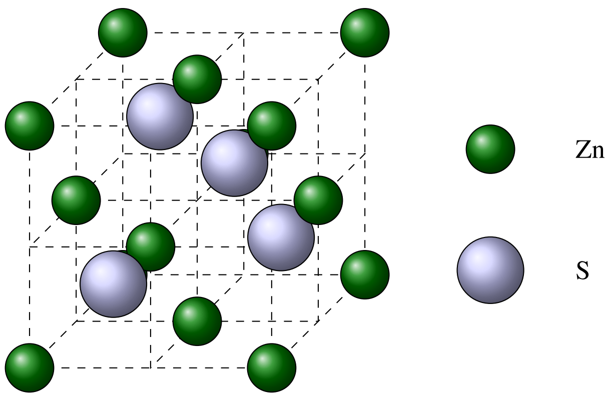
The unit cell contains Cl^- and Na^+. Remember every location in the unit cell has different sharing factor. We will compute the number of atoms in each location and multiply by the sharing factor to calculate the number of Cl and Na in the cell:

Location	Sharing Factor, f	Number of atoms, N	$f \times N$
Corner	$\frac{1}{8}$	$8Na^+$	1
Faces	$\frac{1}{2}$	$6Na^+$	3
sides	$\frac{1}{4}$	$12Cl^-$	3
Inside	1	$1Cl^-$	1

Overall, we have Na_4Cl_4 which corresponds with the formula NaCl.

STUDY CHECK

Calculate the formula for the following unit cell:



Answer: Zn_4S_4.

8.4 Liquid state

This section will cover the liquid state. In particular the importance of the vapor pressure of a liquid and the role of the enthalpy of vaporization–remember enthalpy represents heat.

`Vapor pressure of a liquid` The molecules of a liquid that are in contact with the atmosphere are more likely to be able to scape into the gas phase forming what we call a vapor pressure. This effect is responsible for the humidity the air and the smell of liquid chemicals. Chemicals with high vapor pressure would vaporize readily and of they have a smell would be able to smell them readily also.

`Enthalpy of vaporization,` ΔH_{vap} The enthalpy of vaporization is the energy needed to vaporize a liquid. Think about the smell of a perfume you like. Now, think about the smell of water. Why a perfume smells and water does not. The enthalpy of vaporization of a perfume is small whereas ΔH_{vap} for water is larger ($41 KJ \cdot mol^{-1}$). This means it is easier for the perfume molecules to scape into the gas phase and hence produce a smell. Another example is acetone–nail polish remover. This chemical has a very distinctive smell. ΔH_{vap} for acetone is $31 KJ \cdot mol^{-1}$. If you compare this value with the value of water you can see acetone is more likely to have a smell. Mind that ΔH_{vap} values are normally positive. This corresponds to the fact that we have to give energy to the liquid in order to create a vapor, and hence the process is endothermic.

Sample Problem 77

Order the following compounds from high to low vapor pressure: C_6H_6 (ΔH_{vap}=31$KJ \cdot mol^{-1}$), C_6H_5OH (ΔH_{vap}=39$KJ \cdot mol^{-1}$), H_2O (ΔH_{vap}=41$KJ \cdot mol^{-1}$)

SOLUTION

The larger ΔH_{vap} the harder it is to vaporize a liquid and hence the lower the vapor pressure of the liquid. If we compare the liquids in this example, water has the lowest vapor pressure, whereas cyclohexane (C_6H_6) has the highest vapor pressure.

STUDY CHECK

Order the following compounds from high to low vapor pressure (P_{vap}): NH_3 (ΔH_{vap}=23$KJ \cdot mol^{-1}$), CH_4 (ΔH_{vap}=8$KJ \cdot mol^{-1}$), C_4H_{10} (ΔH_{vap}=15$KJ \cdot mol^{-1}$)

Answer: $P_{vap}(CH_4) > P_{vap}(C_4H_{10}) > P_{vap}(NH_3)$.

`Vapor pressure change with temperature` This pressure strongly depends on temperature. That is the reason why summer days can also be humid days if you live near the seaside. The following formula gives the relation between vapor pressure and temperature. Mind that for every temperature we will have a vapor pressure value. In the formula you will to pairs of temperatures and hence two pairs of vapor pressures:

Add this formula to your flashcard.

$$\boxed{Ln\left(\frac{P_{vap,T_1}}{P_{vap,T_2}}\right) = \frac{\Delta H_{vap}}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)}$$ Clausius-Clapeyron relation

where:

P_{vap,T_1} is the vapor pressure at temperature T_1 in Kelvin

ΔH_{vap} is the enthalpy of vaporization in $J \cdot mol^{-1}$

R=8.314J/K/mol is the constant of the gases in energy units

Sample Problem 78

The vapor pressure of water at 298K is 0.03 atm. Calculate the vapor pressure of water at 323K given $\Delta H_{vap} = 43.9 KJ \cdot mol^{-1}$.

SOLUTION

In order to use the Clausius-Clapeyron relation we need two pairs of (T, P_{vap}) values. In this problem, we have the value of the vapor pressure at 298K, hence we have (298K, 0.03 atm) and they ask the pressure at 323K. Therefore the second pair is (298K, x atm), where X is the vapor pressure at 298–what they are asking in the problem. We can call (298K, 0.03 atm) as (T_1, P_{vap,T_1}) and (298K, X atm) as (T_2, P_{vap,T_2}). At this point we have $T_1 = 298K$ and $P_{vap,T_1} = 0.03atm$ and $T_2 = 323K$ and $P_{vap,T_2} = x$. We also have the enthalpy of vaporization. Minds that this value has to be given in $J \cdot mol^{-1}$ and hence, we will use $\Delta H_{vap} = 43.9 \times 10^3 J \cdot mol^{-1}$. Now we can plug these values into the formula:

$$Ln\left(\frac{0.03}{x}\right) = \frac{43.9 \times 10^3}{8.314}\left(\frac{1}{323} - \frac{1}{298}\right)$$

Let us solve this step by step. First we solve the part on the right:

$$Ln\left(\frac{0.03}{x}\right) = -1.37$$

Now, in order to eliminate the logarithm we should use the exponential function in both sides:

$$\frac{0.03}{x} = e^{-1.37}$$

Calculating the exponential of -1.37 we have:

$$\frac{0.03}{x} = 0.25$$

That leads to a x value of 0.11 atm.

STUDY CHECK

Using the data below, calculate ΔH_{vap} for HNO_3.

T (K)	P_{vap} (mmHg)
10	26.6
20	47.9
30	81.3

Answer: 07.80 mmHg.

8.5 Phase diagrams

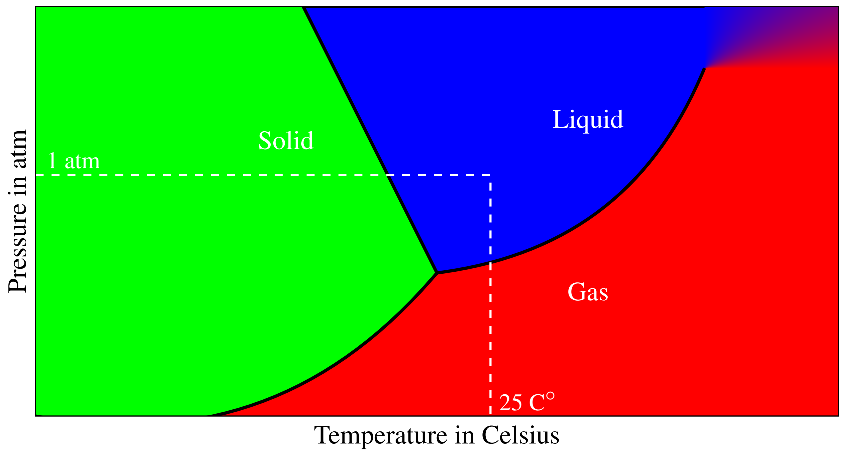


Figure 8.8: The phase diagram of water with pressure in the Y axis and temperature in the X axis. This diagram displays the different states of matter of water for different pressure and temperature conditions. The coordinated of the triple point are (0.0098 C°,0.0060 atm). This means that at this low pressure and temperature conditions we have three phases in contact: water, ice and steam. The coordinates of the critical point are (374 C°, 218 atm). This means that for temperature beyond 374 C° it is not possible to liquify steam.

Water can be found at different states: liquid, solid and gas. We know at room temperature–and atmospheric pressure–water is a liquid. However, what if we warm up a sample of water? When does it become vapor? And more importantly, what if the working pressure is not one atmosphere? Would water boil the same near the sea or on top of a mountain? The answer to all these questions can be found in the phase diagram of water. This section will cover phase diagrams. You will learn how to read phase diagrams in order to predict the state of matter at any temperature and pressure conditions. You will also learn how to identify critical and triple points.

Phase diagram of water A phase diagram is just a diagram with temperature in the X axis and pressure in the Y axis. It tells you whether you have gas, liquid or gas at a large range of pressure and temperature conditions. For example, the figure on the side of the page presents the phase diagram of water and the line indicates the phase present at (Temperature, Pressure) conditions of (25 C°,1 atm). Obviously, this phase is liquid water. *Normal conditions* refer to pressure conditions of 1 atm. Hence, we say that the normal boiling point of water–this means at 1 atm–is 100 C°. In the following we will analyze a set of experiments represented as vertical and horizontal lines in the diagram. Horizontal lines are cooling/heating experiments in which pressure is kept fixed and temperature changes. Vertical lines represent compression/decompression experiments in which pressure changes at constant temperature.

Heating experiments The figure on the side diplays a set of cooling/heating experiments. In the first experiment, we start by having a solid that we heat up to obtain first a mixture between liquid and solid and then a pure liquid. In this experiment we just transitioned between solid into a liquid. Experiment 2 is different. We also start by having a solid. The difference is that this time we reach a point called *tripe point* in this point the three phase coexist at a single pressure and temperature. Therefore, in this experiment, we go from a solid into a mixture of solid, liquid and gas. After that we transition directly into a gas. Experiment number three is called sublimation. In this experiment we start by having a solid that transitions into a gas by means of a mixture of solid and gas.

Compression experiments The figure on the side also diplays a set of compression/decompression experiments. The first experiment is a compression experiment in which we start from a gas and we end up having a liquid by means of a mixture of both. The second experiment start beyond the *critical point* and hence even if you compress the gas you will never reach a liquid state. The critical point is the point beyond which one cannot liquify a gas or gasify a liquid.

Important points in a phase diagram There are two important points in a phase diagram. One is the triple point in which three phases coexist. Another important point is the critical point belong which one cannot liquify or condense the chemical. Remember also the lines in a phase diagram represent phase transitions and hence two phases are present on these lines.

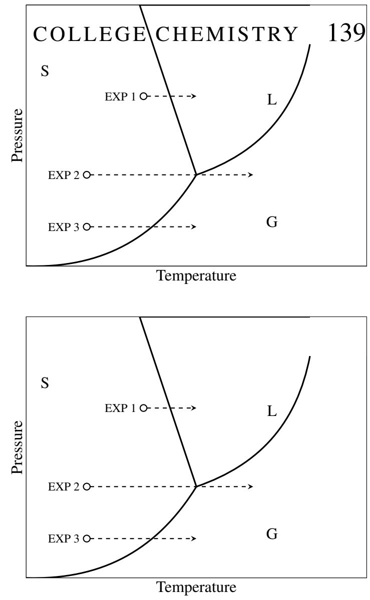


Figure 8.9: Some heating and compressing experiments.

CHAPTER 8

INTERMOLECULAR FORCES

1. What is the strongest Intermolecular force existing between the molecules of CH_3OH:
   (a) dispersion forces (c) hydrogen bonds
   (b) dipole forces

   Ans: (c)

2. What is the strongest Intermolecular force existing between the molecules of H_2:
   (a) dispersion forces (c) hydrogen bonds
   (b) dipole forces

   Ans: (a)

3. What is the strongest Intermolecular force existing between the molecules of CCl_4:
   (a) dispersion forces (c) hydrogen bonds
   (b) dipole forces

   Ans: (a)

4. What is the strongest Intermolecular force existing between the molecules of CH_4:
   (a) dispersion forces (c) hydrogen bonds
   (b) dipole forces

   Ans: (a)

5. What is the strongest Intermolecular force existing between the molecules of CCl_3H:
   (a) dispersion forces (c) hydrogen bonds
   (b) dipole forces

   Ans: (b)

6. What is the strongest Intermolecular force existing between the molecules of HF:
   (a) dispersion forces (c) hydrogen bonds
   (b) dipole forces

   Ans: (c)

7. What is the strongest Intermolecular force existing between the molecules of HCl:
   (a) dispersion forces (c) hydrogen bonds
   (b) dipole forces

   Ans: (b)

8. Which molecule forms intermolecular H bonds?
   (a) HF (b) H_2

   Ans: (a)

9. Which molecule forms intermolecular H bonds?
   (a) NH_3 (b) CH_4

   Ans: (a)

10. Which molecule forms intermolecular H bonds?
    (a) CH_3-O-CH_3 (b) H_2O

    Ans: (b)

11. Which molecule or ion has stronger dipole forces?
    (a) HCl (b) HF

    Ans: (b)

12. Which molecule or ion has stronger dispersion forces?
    (a) CH_3CH_3 (b) CH_4

    Ans: (a)

13. Which substance has higher boiling point?
    (a) CH_3CH_3 (b) CH_4

    Ans: (a)

14. Which substance has higher boiling point?
    (a) CO_2 (b) H_2O

    Ans: (b)

SOLID STATE

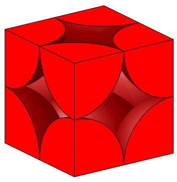
15. Indicate the number of atoms contained in the body-centered (bcc) cubic unit cell, for structures with the same type of atoms?
    (a) 1 (c) 4
    (b) 2

    Ans: (b)

16. Indicate the number of atoms contained in the simple cubic (sc) unit cell, for structures with the same type of atoms?
    (a) 1 (c) 4
    (b) 2

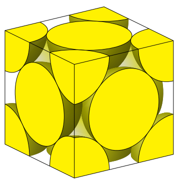
    Ans: (b)

17. The image displays the structure of Polonium. What is the number of atoms per unit cell for this metal?



    Ans: 1

18. The image displays the structure of Gold. What is the number of atoms per unit cell for this metal?



    Ans: 4

19. Classify the following crystalline solid: cesium chloride.

(a) atomic solid
(b) molecular solid
(c) ionic solid

Ans: (c)

20. Classify the following crystalline solid: tungsten.

(a) atomic solid
(b) molecular solid
(c) ionic solid

Ans: (a)

21. Classify the following crystalline solid: acetic acid.

(a) atomic solid
(b) molecular solid
(c) ionic solid

Ans: (b)

22. Classify the following crystalline solid: hydrogen sulfide.

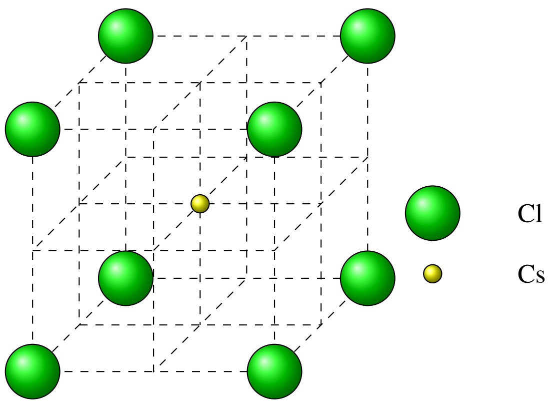
(a) atomic solid
(b) molecular solid
(c) ionic solid

Ans: (b)

23. An element crystallizes in a face-centered cubic lattice and has a density of 1.5 $g \cdot mL^{-1}$ and a cell parameter of 452pm. Calculate the approximate mass of the element.

Ans: 20 $g \cdot mol^{-1}$

24. Calculate the formula for the following unit cell:



Ans: Cs_1Cl_1

Liquid state

25. A liquid has a enthalpy of vaporization of 30 kJ/mol and a boiling point of 122C° at 1.00 atm. Calculate its vapor pressure at 200C°.

Ans: 4.51atm

26. What is the enthalpy of vaporization of a liquid that has a vapor pressure of 500 torr at 100C° and a boiling point of 90C° at 460 torr?

Ans: 9.3 kJ/mol

REVIEW PART D

1. A 2.5 g sample of french fries is placed in a calorimeter with 500.0 g of water at an initial temperature of 21 °C. After combustion of the french fries, the water has a temperature of 48 °C. What is the combustion energy for the process if the calorimeter factor is negligible?
   (a) -23 KJ/g (c) -0.14 KJ/g (e) -5.4 KJ/g
   (b) -11 KJ/g (d) -4.2 KJ/g
2. An unknown metal with mass of 100 g absorbs 6 KJ of heat, and its temperature increases from 22 °C to 23 °C. Determine the specific heat of this metal in $J/g°C$.
   (a) 60 (c) 40 (e) 10
   (b) -60 (d) 160
3. The specific heat of copper is 0.0920 $cal/g°C$, and the specific heat of silver is 0.0562 $cal/g°C$. If 100 cal of heat is added to one g of each metal at 25 °C, what is the expected result?
   (a) The copper will reach a higher temperature.
   (b) The silver will reach a higher temperature.
   (c) The two samples will reach the same temperature.
   (d) The copper will reach a temperature lower than 25 °C.
   (e) The silver will soften.
4. which of the following has a non-zero ΔH_f^0
   (a) S(s) (c) NaCl(s) (e) Cl_2(g)
   (b) O_2(s) (d) Na(s)
5. At constant temperature and pressure, the heats of formation of H_2O(g), CO_2(g), and C_2H_6(g) are given below. Calculate ΔH_f^0 for 1 mol of C_2H_6 gas to oxidize to carbon dioxide gas and water vapor?

$$C_2H_6\,(g) + O_2(g) \longrightarrow 2\,CO_2(g) + 3\,H_2\,O(l)$$

   $\Delta H_f^0(H_2O_{(g)})$=-251KJ/mol; $\Delta H_f^0(CO_{2(g)})$=-393KJ/mol; $\Delta H_f^0(C_2H_{6(g)})$=-84KJ/mol.
   (a) -8730KJ (c) -1455KJ (e) +2910KJ
   (b) -2910KJ (d) +1455KJ
6. Given these two standard enthalpies of formation:

$$S(s) + O_2(g) \longrightarrow SO_2(g)\ \Delta H_1 = -295KJ$$
$$S(s) + \tfrac{2}{3}\,O_2(g) \longrightarrow SO_3(g)\ \Delta H_2 = -395KJ$$

   What is ΔH_f^o for this reaction:

$$O_2(g) + 2\,SO_2(g) \longrightarrow 2\,SO_3(g)$$

   (a) -1380 KJ/mol (c) -295KJ/mol (e) -100KJ/mol
   (b) -690KJ/mol (d) -200KJ/mol
7. If ΔH_f^o for a reaction is positive?
   (a) the reaction rate is generally very fast.
   (b) H^o(products) is smaller than H^o(reactants).
   (c) the reaction rate is generally very slow.
   (d) the process is endothermic
   (e) H^o(reactants) is bigger than H^o(products).
8. Calculate ΔH_f^o for the reaction given the following information:

$$N_2(g) + 3\,H_2(g) \longrightarrow 2\,NH_3\,(g)$$

   $\Delta H_f^o(NH_3(g))$= -46KJ/mol
   (a) +100KJ (c) -920KJ (e) 10KJ
   (b) -92KJ (d) -120KJ
9. Calculate ΔH_f^o for the reaction given the following information:

$$N_2(g) + 3\,H_2(g) \longrightarrow 2\,NH_3\,(g)$$

   $\Delta H_f^o(NH_3(g))$= -46KJ/mol
   (a) +100KJ (c) -920KJ (e) 10KJ
   (b) -92KJ (d) -120KJ
10. The strongest interactions between molecules of ammonia (NH_3) are
    (a) ionic bonds.
    (b) H bonds.
    (c) polar covalent.
    (d) dipole-dipole.
11. The strongest interactions between molecules of hydrogen (H_2) are
    (a) ionic bonds.
    (b) H bonds.
    (c) polar covalent.
    (d) dipole-dipole.
    (e) dispersion forces.
12. The strongest interactions between molecules of hydrogen chloride are
    (a) ionic bonds.
    (b) covalent bonds.
    (c) H bonds.
    (d) dipole-dipole interactions.
    (e) dispersion forces.
13. Which of the following boils at the highest temperature?
    (a) CH_4 (c) C_3H_8 (e) C_5H_{12}
    (b) C_2H_6 (d) C_4H_{10}
14. Which one of the following classifications is incorrect?
    (a) H_2O(s), molecular solid
    (b) C_4H_{10}(s), molecular solid
    (c) KF(s), ionic solid
    (d) SiC(s), covalent solid
    (e) S(s), metallic solid

Answers:

1. (a)	**2.** (a)	**3.** (b)	**4.** (c)	**5.** (c)	**6.** (d)	**7.** (d)	**8.** (b)
9. (b)	**10.** (b)	**11.** (e)	**12.** (d)	**13.** (e)	**14.** (e)		

PART E

9

Atomic structure

MATTER is everywhere around you, from the water you drink to the air you inhale. Matter is made of elements and elements are made of atoms. In the world we can also find light, that somehow at first sight seems different than matter. Light is warm and it has color. This chapter covers the structure of the atoms, with a focus on the structure of the many electrons than atoms are made of. It also focuses on light and the interaction of light with matter. You will be able to understand differences in electronic configuration.

GOALS

1. Compute frequency, wavelength and the energy of light
2. Compute energy the levels for H
3. Compute transition energies for H
4. Obtain electronic configurations
5. Compare periodic properties

9.1 Light

What is light? Light–also called electromagnetic radiation–is a form or energy. There are many different types of radiation. Think about the light coming from a bulb, or the radiation that warms up your food in a microwave, or even when you heat a pizza in the over. Radiation is characterized by frequency and by the wavelength. This section will cover these two properties of light.

Discussion: Look around your apartment and find four different types of radiation.

`Frequency and energy` Light travels in time. That is the reason you can hear a whistle from afar. The *frequency* of a radiation–the frequency of a specific type of light–characterizes how this radiation oscilates in time. The unit of frequency is the hertz and frequency is represented by the symbol γ. At the same time, frequency is connected to the energy of radiation. High frequency radiation are very energetic. Think for example of gamma rays; these type of radiation produced in nuclear plant has very high frequency and hence is very energetic. The formula that related frequency with energy is:

$$\boxed{E = 6.6 \times 10^{-34}\gamma}$$ Frequency formula

Add this formula to your flashcard.

where:

E is the energy in joules

6.6×10^{-34} is called Plank's constant, h

γ is frequency in hertzs (Hz)

As you can see in the previous formula, the frequency is directly proportional to frequency.

`Wavelength and energy` Light also travels in space. As it moves, it oscillates in space. Think about dropping a stone into a lake. As you drop the pebble, the energy from the pebble propagates in the surface in water. The energy of light also propagates in space and the *wavelength* of a radiation is the distance between two consecutive peaks. As such, wavelength, represented by the letter λ and with units of nm is also related to energy by means of the formula:

Add this formula to your flashcard.

$$\boxed{E = \frac{1.98 \times 10^{-16}}{\lambda}} \quad \text{wavelength formula}$$

where:

E is the energy in joules

λ is wavelength in nm

$1.98 \times 10^{-16} = h \cdot c \cdot 10^9$ was adjusted to be able to use λ in nm

Mind that wavelength is inversely related to energy. That means, the larger wavelength the smaller energy. Also mind that wavelength refers to the movement of light in space and frequency refers to the movement in time.

`Relationship between wavelength and frequency` Wavelength and frequency are related by means of the speed of light c that is 3×10^8. However, if we want λ to be in nm we can use the following formula:

Add this formula to your flashcard.

$$\boxed{\gamma = \frac{3 \times 10^{17}}{\lambda}}$$

where:

γ is frequency in Hz

λ is wavelength in nm

$3 \times 10^{17} = c \cdot 10^9$ was adjusted to be able to use λ in nm

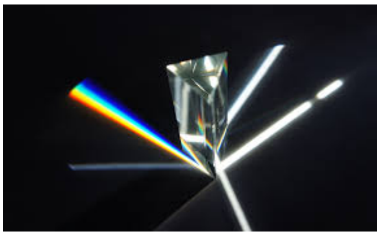


Figure 9.1: White light contains many colors.

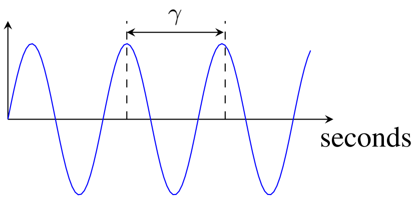


Figure 9.2: Frequency refers to time

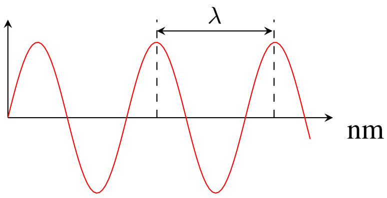


Figure 9.3: Wavelength refers to space

All radiation always travels at the speed of light. At the same time, this speed is the maximum speed allowed for any object.

Sample Problem 79

Calculate: (a) the energy of a radiation with wavelength of 300nm; (b) the energy of a radiation with frequency of 10^{19} Hz; (c) the frequency of a radiation with wavelength of 300nm.

SOLUTION

(a) To answer the first question we will use the wavelength formula, as wavelength is given ($\lambda = 300nm$) and we need to calculate the energy (E), in Joules:

$$E = \frac{1.98 \times 10^{-16}}{\lambda} = \frac{1.98 \times 10^{-16}}{300} = 6.6 \times 10^{-19} J$$

(b) To answer the second question we will use the frequency formula, as frequency is given ($\gamma = 10^{19} Hz$) and we need to calculate the energy (E), in Joules:

$$E = 6.6 \times 10^{-34} \gamma = 6.6 \times 10^{-34} \cdot 10^{19} = 6.6 \times 10^{-15} J$$

(c) To answer the last question we will use the formula that related frequency with wavelength–through the speed of light–as frequency is asked and wavelength is given ($\lambda = 300nm$); mind the units of frequency are herts:

$$\gamma = \frac{3 \times 10^{17}}{\lambda} = \frac{3 \times 10^{17}}{300} = 1 \times 10^{15} Hz$$

STUDY CHECK

Calculate: (a) the wavelength of radiation with energy of $5.6 \times 10^{-19} J$; (b) the frequency of a radiation with frequency of $4.8 \times 10^{-18} J$; (c) the wavelength of a radiation with frequency of $2 \times 10^{15} Hz$.

Answer: (a) $353nm$; (b) $7.2 \times 10^{10} Hz$; (c) $4.1 \times 10^{6} nm$

Radiation	γ (Hz)
Gamma	>3×10^{19}
X-rays	$3 \times 10^{19} - 3 \times 10^{16}$
UV	$3 \times 10^{16} - 8 \times 10^{14}$
IR	$4 \times 10^{14} - 4 \times 10^{11}$
MicroW	$3 \times 10^{11} - 3 \times 10^{8}$
RadioW	$3 \times 10^{8} - 3 \times 10^{3}$

Types 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 radiowaves.

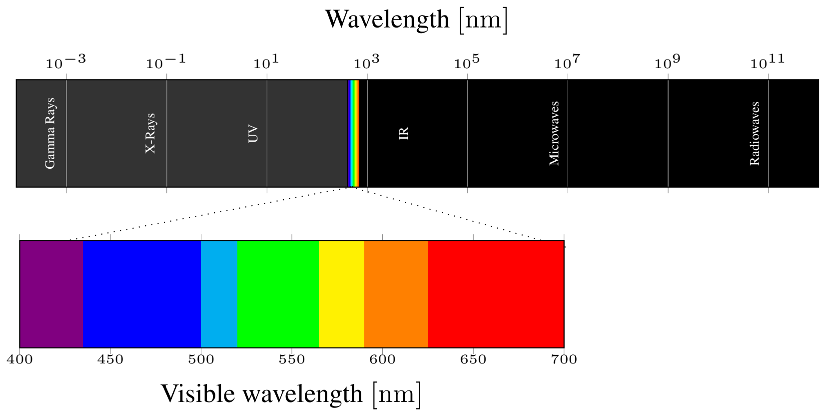


Figure 9.4: Spectrum of the electromagnetic radiation

Color	λ (nm)
Violet	380-450
Blue	450-485
Cyan	485-500
Green	500-565
Yellow	565-590
Orange	590-625
Red	625-740

For example, radiation with wavelength of 10^{-2} nm belongs to gamma rays radiation, whereas radiation with wavelength of 10^{4} nm belongs to the Infrared. Only a small range of wavelengths belong to visible radiation. This means you will not be able to see IR radiation. 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.

Sample Problem 80

Indicate: (a) the color of a radiation with $\lambda = 650nm$; (b) the type of a radiation with $\lambda = 10^{5}$ nm; (c) the type of a radiation with $\gamma = 10^{16}$ Hz.

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. Finally, in order to answer the last question we need to convert frequency into wavelength, as the figure above only indicates wavelength. $\gamma = 10^{16}$ Hz corresponds to $\lambda = 30$ nm. Hence, this frequency corresponds to the UV.

STUDY CHECK

Indicate: (a) the color of a radiation with $\gamma = 7.5 \times 10^{14}$ Hz; (b) the type of a radiation with $\gamma = 10^{8}$ Hz.

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

9.2 The atomic spectrum of hydrogen

This section will explain the atomic spectrum of hydrogen. Atoms emit light, but not any type of light. They emit specific frequencies of radiation. In the case of hydrogen the atomic spectrum of hydrogen is just an image that show the different wavelengths of the radiation emitted by atoms of hydrogen. This section will also cover the reasons for this specific emission of light and will introduce the Bohr model that explains the lines in the atomic spectrum of hydrogen.

`Spectrum of hydrogen` The atomic spectrum of hydrogen presents a set of lines in a black background. Those lines correspond to radiation emitted by the atoms of hydrogen. 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. Why is this spectrum important? this spectrum historically was used to understand the structure of the atom and in particular the structure of the electrons inside the atom. That is why we go over this section in this class.

Figure 9.5: The four visible lines in the Balmer series.

`The Bohr model` The Bohr model is a model that explains the electronic structure of the atom and in particular explains very well the atomic spectrum of hydrogen. This model is based on the idea that the electron of hydrogen moves around the nucleus in only in certain allowed circular orbits. Each orbit is called energy level. Each orbit is also characterized by a energy E_n and a number n. For example, the first level energy is E_1 and the third level energy is E_3; the higher the n the larger–more positive–is the energy of the level. The following formula gives you the energy value for each level:

Add this formula to your flashcard.

$$\boxed{E_n = -2.178 \times 10^{-18}\frac{1}{n^2}} \quad \text{Bohr formula in J}$$

where:

E_n is the energy of the level n in joules

n is the number of the level

$-2.178 \times 10^{-18} J = R_H$ is called the Rydberg

For example, the energy of the first level is $E_1 = -2.178 \times 10^{-18} J$, whereas the energy of the third level is $E_3 = -5.44 \times 10^{-19} J$.

`Electron-Volt a new unit of energy` The energy values for the energy levels in J are somehow very small. Sometimes, it is convenient to use another energy unit that makes these values have more reasonable values.

Add this formula to your flashcard.

$$\boxed{1eV = 1.60218 \times 10^{-19} J} \quad \text{or} \quad \boxed{\frac{1eV}{1.60218 \times 10^{-19} J}} \quad \text{or} \quad \boxed{\frac{1.60218 \times 10^{-19} J}{1eV}}$$

Now let's see how is the energy in eV of the first level is $E_1 = -13.6eV$, whereas the energy of the third level is $E_3 = -1.5J$.

`Energy levels of hydrogen` Let's understand a bit more the idea behind energy levels. These are just numbers that represent the location–in energy units–of the different locations where the electrons can be attached to the hydrogen atom. The first level is E_1 and is the most negative energy value. This is also the most stable level, that is, the electrons in these level are tightly bonded to the nucleus. The following levels have more and more positive energy. Is we compare two energy levels, for example $E_2 = -3.40eV$ and $E_4 = -0.85eV$, an electron on the level number four is less stable than on level two. Hence it would be easier to remove an electron from level number four than from level two. For small n values the levels are spear from spread from each other. When n increases, the energy levels are more and more closer to each other. Finally, there are infinite number of levels.

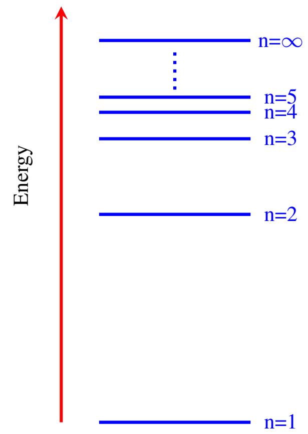


Figure 9.6: The energy levels of hydrogen

`Transition energies` Bohr's models is able to explain the atomic spectrum of hydrogen. Each line in the spectrum represents a transition between two levels of energy. For example the line at 102nm represents the transition of an electron between the level three and the level one. The atomic spectrum of hydrogen is obtained by means of exciting hydrogen atoms with energy, so that the electron jumps from a lower level into a higher level. When the atom relaxes, it emits light as the electrons move back from high every levels into lower–more stable–levels.

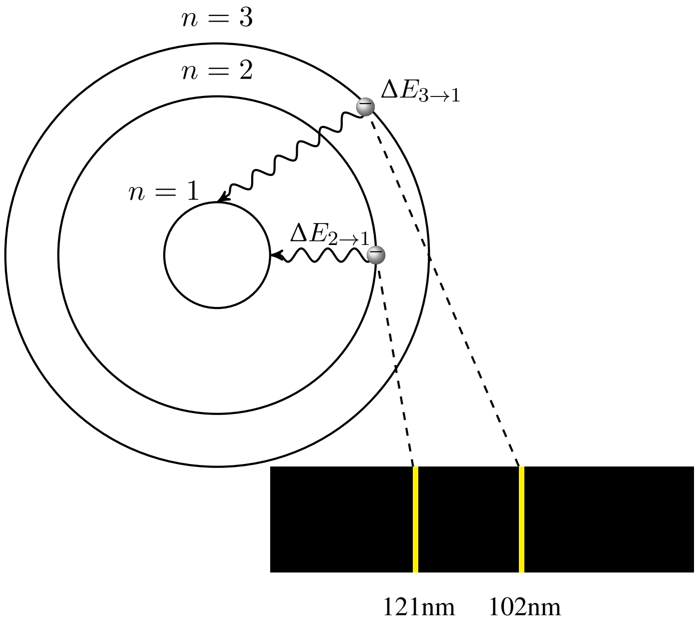


Figure 9.7: Representation of two energy transitions

The energy for an electron transition between two energy levels, from n_1 to n_1 is given by:

$$\boxed{\Delta E_{n_2\to n_1} = -2.178 \times 10^{-18}\left(\frac{1}{n_2^2} - \frac{1}{n_1^2}\right)}$$ Energy transition formula

Add this formula to your flashcard.

where:

$\Delta E_{n_2\to n_1}$ is the energy in joules for the transition, the line in the spectra

n_2 is the number of the final level

1860 n_1 is the number of the initial level

$-2.178 \times 10^{-18} J = R_H$ is called the Rydberg ($-13.59 eV = R_H$)

Sample Problem 81

Calculate the following transition energies:

(a) $\Delta E_{4 \to 3}$ in J

(b) $\Delta E_{4 \to 3}$ in eV

(c) Calculate the final energy level for a transition with energy 1.34eV knowing the first energy level involved in the transition is $n = 3$

SOLUTION

(a) We will use the energy transition formula to calculate the energy needed to move one electron from $n_1 = 4$ to $n_2 = 3$:

$$\Delta E_{n_2 \to n_1} = -13.59 \left(\frac{1}{n_2^2} - \frac{1}{n_1^2} \right) = -13.59 \left(\frac{1}{3^2} - \frac{1}{4^2} \right)$$
$$= -13.59 \left(0.111 - 0.0625 \right) = -0.66 eV$$

The transition energy is negative, this means the atom releases energy when transitioning between these two levels.
(b) We will use the energy transition formula this time in eV:

$$\Delta E_{n_2 \to n_1} = -2.178 \times 10^{-18} \left(\frac{1}{n_2^2} - \frac{1}{n_1^2} \right) = -2.178 \times 10^{-18} \left(\frac{1}{3^2} - \frac{1}{4^2} \right)$$
$$= -2.178 \times 10^{-18} \left(0.111 - 0.0625 \right) = -1.058 \times 10^{-19} J$$

(c) In this case, we know $\Delta E_{n_2 \to 3}$ and we know the initial level is $n_1 = 3$. We can certainly solve for n_2:

$$1.34 = -13.59 \left(\frac{1}{n_2^2} - \frac{1}{3^2} \right) = -13.59 \left(\frac{1}{n_2^2} - \frac{1}{9} \right)$$

Solving for n_2 we have: $n_2 = 9$. Mind you need to square root n_2^2 to get the final value of n_2.

STUDY CHECK

Calculate the following transition energies: (a) $\Delta E_{9 \to 3}$ in J and (b) $\Delta E_{5 \to 4}$ in eV.

Answer: (a) $-2.15 \times 10^{-19} J$; (b) -0.30eV.

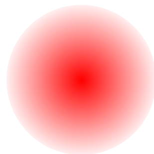
9.3 Quantum mechanics and electronic structure

Bohr's model was a simplistic model. Still, it was able to correctly predict the fine structure of hydrogen–the atomic spectrum of hydrogen–with its energy transitions. The downside of this model resulted from considering that the electron moves in different orbits. The correct assumption of the model was that the energy levels of the atom were quantized–electrons can only exist in specific energy levels characterized by the number n and not in a continuum of energy. This section will cover a more realistic theory that describe the structure of the atom: quantum mechanics. The outcome of this section is the existence of orbitals and quantum numbers.

> "The present is the only things that has no end."
> *Schrödinger*

`Quantized energy and continuum energy` Quantum mechanics is a theory used to do modeling in chemistry. Think of an engineer designing a plane. Before building and selling the plane engineers have to model it with computer software to make sure the plane will work properly. In a similar way, chemists carry modeling. The theory behind normal engineering modeling is classical mechanics, based in Newton's law.The theory behind chemistry modeling is quantum mechanics, based in Schrödinger equation. Classical mechanics is based on the idea that the energy of a system, a plane or a car, is continuum, that is, the car or a plane can have any possible energy, starting from zero to any number you can think of. Differently, quantum mechanics is based in the idea that the energy of a system, an atom or a molecule, is quantized, that is the energy of an atom or a molecule can only be certain specific values. For example, a hydrogen atom can only have E_1, or E_4 or E_5. It would never have a energy corresponding to $E_{\frac{1}{2}}$. That is the meaning of quantized.

Representation of an electron



Representation of a tennis ball

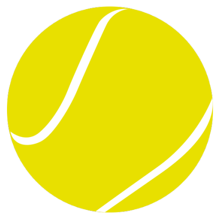


Figure 9.8: An electron is a quantum object whereas a tennis ball is a classical object

`The wave function` In quantum mechanics the wave function of an atom or a molecule Ψ is a complex function that contains all information of this atom or this molecule. By means of this function, we can simulate the behavior and properties of the atom. You want to think of Ψ as a box that contains information, in particular all information of the system you want to simulate, such as a large molecule or a small atom. Ψ depends on several variables, in particular the position and two angles; as such Ψ results from the multiplication of an angular and radial function.This function has no real meaning and only its square value has a real physical interpretation. Ψ^2 represents the probability of finding an electron near a particular point in space. In order to calculate any property of the system of interest, we need to solve Schrödinger's equation:

$$\hat{H}\Psi = E\Psi$$

$\hat{H}$ is called Hamiltonian operator and depends on the kinetic and potential energy of the system.

`Orbitals` An orbital is a single-electron wave function. In another words is a wave function that contains information of a single electron. The square value of an orbital represents the probability of finding an electron at a specific location. Electrons are very different than larger objects such as for example a tennis ball. Larger objects are localized, that is they are located at a specific point in space. Differently, electrons are delocalized, that means they are not located at a single point in space and time, and therefore we can only guess the probability of finding the electron an a specific point.

`Different types of orbitals` There are four different types of orbitals: *s* orbitals, *p* orbitals, *d* orbitals and *f* orbitals. At the same time, there are one type of *s* orbitals, three types of *p* orbitals (p_x, p_y and p_z) and five types of *d* orbitals (d_{xy}, d_{xz},

d_{yz}, d_{z^2} and $d_{x^2-y^2}$). For example, *s* orbitals are very easy to visualize and they look like spheres. Differently, *p* orbitals are more complex. To understand their shape, image you hold a ballon with both of your hands and you twist it in different directions.

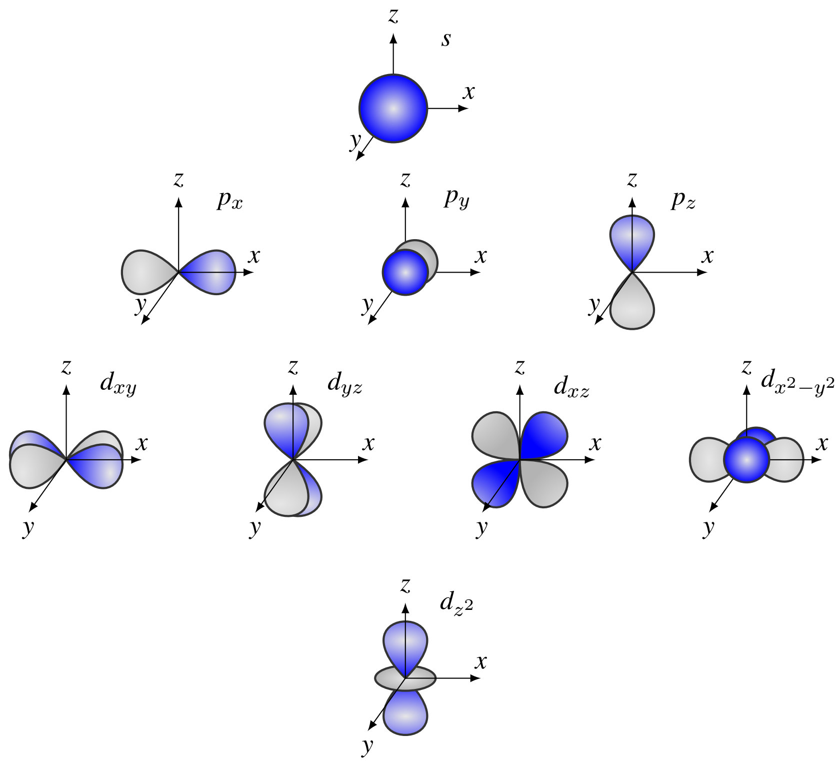


Figure 9.9: s, p and d orbitals. s orbitals have spherical symmetry, whereas p orbitals have two lobes with dumbbell shape, a positive and negative lobe. The name for the different p orbitals depend on the axis that it crosses. For example, p_x orbitals cross the x axis. The positive part of this orbital corresponds with the positive part of the axis–the right side. d orbitals have more complex shapes. Most of d orbitals have four lobes, two positive and two negative. For example, orbital d_{xy} is placed in the xy plane without crossing any axis. Differently, orbital $d_{x^2-y^2}$ is located in the xy plane and cross both axis. The blue part of the orbital represent the positive side of the orbital, whereas the grey part is negative.

9.4 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 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. On the other hand, 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 display all orbitals in an atom, whereas the abbreviated only display the valence electrons–these electrons are less-tied to the nucleus–and the core has nobel. At the same time, every orbital is characterized by a set of numbers–these are called quantum numbers. These numbers are not independent from each other and there are constrictions that relate the possible values of the quantum numbers. This section will show you how to construct electron configurations and how to extract quantum numbers from it.

`Electron energy levels` 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 have the same energy. The energy levels are labeled with a number n that equals to 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

$$2n^2 \tag{9.1}$$

in which n is the energy level. You can see by using this formula that for example, the third level can accommodate 27 electrons.

l value	Orbital label
0	s
1	p
2	d
3	f

Quantum number	Values
n	1, 2, 3...
l	0, 1, ..., n-1
m_l	-l, -l+1, ..., 0, ..., l-1, l
m_s	$+^1/_2$ or $-^1/_2$

Orbital	Maximum Number of electrons
s	2
p	6
d	10
f	14

Sample Problem 82

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: $n = 9$.

Sublevels 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.

Orbital Filling 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 put a maximum of two electrons. That is why you will find s^1 orbitals and s^2, with the latest being completely filled with electrons. In a p orbital you can put a maximum of six electrons and in a d orbital a maximum of ten. Finally, in a d orbital you can put fourteen or less electrons. For example, the orbital notation p^2 is correct as in p orbitals you can put six or less electrons. In this case, this orbital still have space to accept more electrons. Differently, the notation $d^1$2 is incorrect, as in d orbitals you can fit ten or less electrons and never twelve.

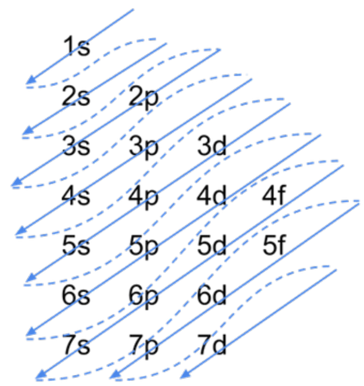


Figure 9.10: Table indicating the orbital filling order. In order to use this table you need to start by filling the orbital $1s$, after that you follow the arrow until next level and start from the begining of next arrow. For example, after $3d$ you need to fill $6s$.

In order to fill the orbitals you should follow Figure 9.10. You start from the top of the table and follow the arrows that indicates the orbitals ordering. For examples the first orbital to be filled will be $1s$. After that you should fill $2s$ 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 electrons 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 $1s^2$ and the one for Li is $1s^2 2s^1$. 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 what orbitals to fill? Figure 9.10 shows the orbital order. You need to start from the top of the image, from orbital $1s$ and proceed to 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 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 83

Obtain the 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 9.10 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: $1s^2 2s^2 2p^2$. The s orbitals are all filled, whereas the p orbital is only occupied with two electrons.

STUDY CHECK

Obtain the electronic configuration of Ni.

Answer: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$.

Abbreviated Electron Configuration The electronic configuration of Nickel (Ni) is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$, whereas the electronic configuration of Argon (Ar) is $1s^2 2s^2 2p^6 3s^2 3p^6$. If you look closely, the electronic configuration of Argon is part of the electronic configuration of Nickel. We called both of these configuration the *full electronic configuration*. We can make the configuration of Nickel shorter by writing the electronic configuration of Nickel as: $[Ar]4s^2 3d^8$. We call this last configuration as 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 the 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 by the orbital $4s$, according to the 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, it has 18 core electrons in the Ar core and 10 electrons in the valence.

Let us work another example: Cd. It has 48 electrons, and is located in group 5. The novel gas on 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 belong to period five. In the valence electrons, we will place 12 electrons. Hence, the abbreviated configuration will be: $[Kr]5s^2 4d^{10}$.

Sample Problem 84

Obtain the full and abbreviated electronic configuration of Silver (Ag, Z=47).

SOLUTION

The atomic number of Ag is Z=47 and that means Nb has 47 electrons. The orbital order from Figure 9.10 is: $1s$,$2s$, $2p$, $3s$, etc. Each s orbital can fit two electrons, whereas the occupancy of the p orbitals is six electrons. d orbitals can fit 10 electrons. 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 elenctrons.

STUDY CHECK

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

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

Orbital diagrams orbital diagrams are a visual way to represent the electron configuration of an atom. Every s orbital is represented by a box that can fix two electrons. Differently, p orbitals are composed of three boxes–p_x, p_y and p_z–with a total capacity of six electrons, with two electrons per box. d orbitals are represented by five boxes with a total capacity of ten electrons. Only the valence electrons are normally represented.

s $\boxed{\uparrow\downarrow}$

p $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$

d $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$

Sample Problem 85

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

SOLUTION

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

O $[He]2s^2 2p^4$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow}$ $\boxed{\uparrow}$

2s 2p

STUDY CHECK

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

Answer: $\boxed{\uparrow}$

Quantum numbers Every electron in an orbital is characterized by four different quantum numbers. Each electron in an atom has unique values for these quantum numbers. The first quantum number n is called *principal quantum number* and has integral values such as 1, 2 or 4. n cannot be zero. The second quantum number l is called *angular quantum number* and has integral values such as 1, 0 or 4. l can be zero and the values of l depend on the values of n. In particular l goes from 0 until $n-1$. For example if $n=3$, therefore l can vary be: $0,1,2$. The third quantum number m_l is called *magnetic quantum number* and has integral values that vary from $-l$, $-l-1$, ..., 0, ..., $l+1$, l. For example, if $l=3$, m_l can be any of these values: -3, -2, -1, 0, 1, 2, or 3. The fourth quantum number m_s is called *spin quantum number* and can only be either $+^1/_2$ or $-^1/_2$. Two of the quantum numbers can be extracted from the orbital label. For example

Sample Problem 86

Indicate if the following combination of quantum numbers is allowed:

n	l	m_1	m_s	**Allowed?**
1	1	0	$+^1/_2$	
2	0	0	$+^1/_2$	
3	3	-1	$-^1/_2$	

SOLUTION

The four quantum numbers are not independent. The quantum number n is related to the quantum number l and the number l is related to m_l. The only quantum number that is independent is the spin, m_s which can be $+^1/_2$ or $-^1/_2$. The first combination is not possible, as if $n=1$, l can only be $n-1$, that is zero. The second combination is correct, as if $n=2$, the number l can be: 0 or 1. At the same time if $l=0$, then m_l can also be zero. And finally, the spin value of $+^1/_2$ is allowed. The last combination is not allowed, as n and l cannot be the same number.

n	l	m_1	m_s	**Allowed?**
1	1	0	$+^1/_2$	No
2	0	0	$+^1/_2$	Yes
3	3	-1	$-^1/_2$	No

STUDY CHECK

Indicate if the following combination of quantum numbers is allowed:

n	l	m_1	m_s	**Allowed?**
4	3	3	$+^1/_2$	
4	3	3	0	
2	1	-1	$+^1/_2$	

Answer: yes, no, yes.

9.5 Periodic Trends

Some atomic properties are periodic, which means that they follow a certain trends in the periodic table. These properties increases or decreases across a period and then the trend is repeated again in each successive period or group. We will consider here four atomic properties: the atomic size, the ionization energy, the metallic character and the electronegativity.

`Atomic Size` The *atomic radius* is the distance from the nucleus to the valence electrons. The atomic radius increases when going down a group in the periodic table (the columns of the table) and when going from right to left on the periodic table (in a period). For example, if we compare the atomic size of H and He, as He is further right it has a smaller radius ($r_{He} < r_H$). Similarly, if we compare H and Li, as Li is further down it will have a larger radius ($r_H < r_{Li}$).

`Ionization Energy` The *ionization energy* (IE) is the energy needed to remove an electron from an atom in gas state. 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 the metals and high for the 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 ($IE_H > IE_{Li}$). The ionization energy of hydrogen is 313Kcal:

$$H_{(g)} \longrightarrow H_{(g)}{}^{+} + e^{-} + 313\,\text{Kcal}$$

`Metallic character` All elements in the periodic table have somehow a certain *metallic character* as they all can loose 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 in 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 ($MC_K > MC_{Ca}$). For the case of F and Cl, Cl is more metallic as it is located further down a group ($MC_{Cl} > MC_F$).

`Electronegativity` Atoms combine with each other and by sharing or give away electrons. The *electronegativity* of an atom is its ability to attract the shared electrons when combined. Imagine two people holding a pillow with their hands, with the pillow representing a pair of electrons. If one of the person is more electronegative it will pull the pillow closer to them–it will attract more the electrons. Electronegativity (EN) increase (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. When comparing the EN of I and F, we found that F is more electronegative as it is further up in the same group ($EN_F > EN_I$).

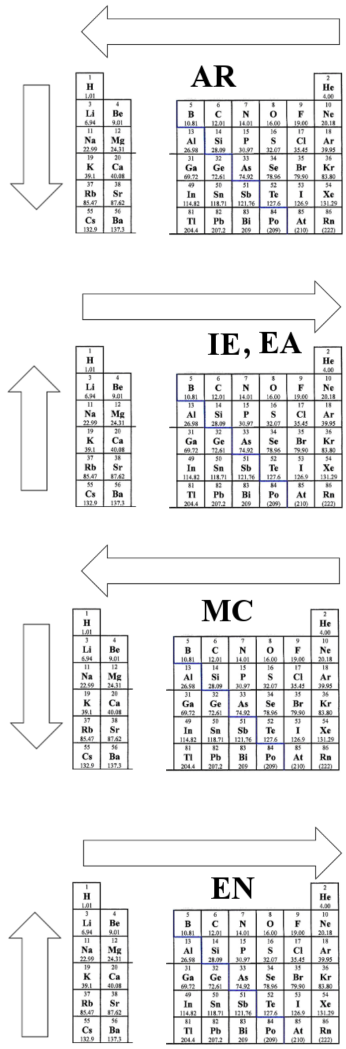


Figure 9.11: Periodic properties across the table

Sample Problem 87

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.

Electron affinity The electron affinity (EA) is the tendency of an atom to receive one or more electrons becoming negatively charged. Electron affinities are positive values and the larger this value the larger the tendency of the atom to accept electrons. For example, the electron affinity of hydrogen is 17Kcal:

$$H_{(g)} + e^- + 17\,\text{Kcal} \longrightarrow H^-_{(g)}$$

EA increases from left to right in a given period, that means the elements from the right of the table are more willing to gain electrons becoming negatively charged.

CHAPTER 9

LIGHT

1. Calculate the energy in joules of a radiation with frequency 2.0×10^{18} Hz?

Ans: 1.32×10^{-15}J

2. Calculate the frequency of a radiation with energy 5.6×10^{-20} J?

Ans: 8.5×10^{13}Hz

3. Calculate the energy in joules of a radiation with wavelength 653 nm?

Ans: 3.03×10^{-19}J

4. Calculate the wavelength of a radiation with energy 5.34×10^{-16}J?

Ans: 0.37 nm

5. Calculate the wavelength of a radiation with frequency of 3.4×10^{14} Hz?

Ans: 882 nm

6. Indicate the color of a radiation with λ= 510nm.

Ans: Green

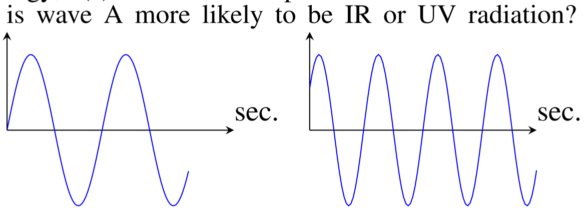
7. Indicate the color of a radiation with λ= 580nm.

Ans: Yellow

8. Classify the nature of a radiation with γ= 3.4×10^{8}Hz

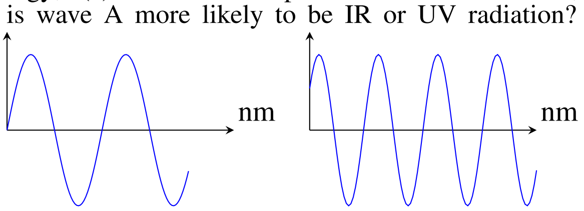
Ans: Microwaves

9. Sections of two electromagnetic waves A and B are represented below. Rank them in order of (a) increasing frequency; (b) increasing energy; (c) If wave B represents visible radiation, is wave A more likely to be IR or UV radiation?



Ans: (a) γ_A<γ_B; (b) E_A<E_B; (c) IR

10. Sections of two electromagnetic waves A and B are represented below. Rank them in order of (a) increasing wavelength; (b) increasing energy; (c) If wave B represents visible radiation, is wave A more likely to be IR or UV radiation?



Ans: (a) λ_B<λ_A; (b) E_A<E_B; (c) UV

THE ATOMIC SPECTRUM OF HYDROGEN

11. Which of these electron transitions correspond to absorption of energy and which to emission?

(a) $\Delta E_{1\to2}$
(b) $\Delta E_{2\to1}$
(c) $\Delta E_{3\to1}$
(d) $\Delta E_{3\to5}$
(e) $\Delta E_{3\to5}$

Ans: absorption: (a), (d), (e); Emission: (b), (c)

12. Use the Bohr equation to find the energy of the photon emitted when an H atom undergoes a transition from $n = 1$ to $n = 4$.

Ans: 2.04×10^{-18} J

13. Use the Bohr equation to find the wavelength (in nm) of the photon emitted when an H atom undergoes a transition from $n = 2$ to $n = 4$.

Ans: 485 nm

14. Use the Bohr equation to find the frequency (in Hz) of the photon emitted when an H atom undergoes a transition from $n = 1$ to $n = 5$.

Ans: 3.16×10^{15} Hz

15. An electron in the lowest energy level of H atom absorbs a photon of wavelength 96.97 nm. Indicate the final energy level of the electron moved.

Ans: $n = 5$

QUANTUM MECHANICS AND ELECTRONIC STRUCTURE

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

Ans: (a) two orbitals; (b) 3 orbitals; (c) none; (d) 16 orbitals.

17. Indicate if the following combination of quantum numbers are allowed:

n	l	m_1	m_s	**Allowed?**
4	4	1	$+^1/_2$	
2	1	4	$+^1/_2$	
4	2	-2	$-^1/_2$	

Ans: no, no, yes.

18. For each of the following sublevels, give the n and l values and the number of orbitals: (a) 6s; (b) 4d; (c) 2p.

Ans: (a) $n = 6; l = 0$; (b) $n = 4; l = 2$; (c) $n = 2; l = 1$.

19. What is the element with the electron configuration $1s^2 2s^2 2p^6 3s^2 3p^5$?

(a) Be
(b) F
(c) Cl
(d) S
(e) Ar

Ans: (c)

20. What is the element with the electron configuration $1s^2 2s^2 2p^6 3s^2 3p^4$?

    (a) Be (c) Cl (e) Ar
    (b) F (d) S

Ans: (d)

21. What is the element with the abbreviated electron configuration $[Kr]5s^2 4d^8$?

    (a) Ni (c) Pt (e) Xe
    (b) Pd (d) Kr

Ans: (b)

PERIODIC PROPERTIES

22. Of the elements: B, C, F, Li, and Na, the element with the largest atomic radius is

    (a) B (c) F (e) Na
    (b) C (d) Li

Ans: (e)

23. Of the elements: B, C, F, Li, and Na, the element with the largest ionization energy is

    (a) B (c) F (e) Na
    (b) C (d) Li

Ans: (e)

24. Of the elements: B, C, F, Li, and Na, the element with the largest electronegativity is

    (a) B (c) F (e) Na
    (b) C (d) Li

Ans: (c)

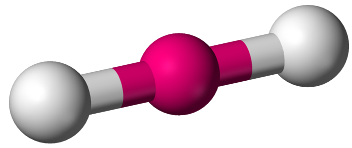
25. Of the elements: B, C, F, Li, and Na, the element with the largest metallic character is

    (a) B (c) F (e) Na
    (b) C (d) Li

Ans: (d)

10

Chemical Bond and Geometry



MOLECULES can found in a myriad of different shapes in nature. Some like carbon dioxide are one-dimensional or linear, other like methane have a three-dimensional shape. This chapter is devoted to the analysis of the molecular shape and to study more advanced models of the chemical bond. More importantly, this chapter is devoted to the reason why different molecules have different shapes: the molecular bond and the existence of lone pairs of electrons. You will be able to draw the connections between the atoms of a molecule. The chemical bonds and the molecular geometry affect the properties of a molecule. At the end of the chapter we will address the polarity of molecules and you will understand the reasons why you use soap to get rid of oil while doing dished.

GOALS

1. Construct electron-dot structures
2. Identify the geometry of a molecule
3. Identify the polar character of a molecule
4. Calculate the bond hybridization
5. Interpret molecular orbital diagrams

10.1 Electron-dot structures of atoms

Atoms are made of protons, neutrons and electrons. Electrons–in particular valence electrons–are responsible for the main chemical properties of an atom. These electrons are weakly tied to the nucleus in comparison with the core electrons and hence they can be exchanged. Atoms in a molecule, with a few exception, tend to be surrounded by eight electrons so that its electron configuration resembles a noble gas. This is known as the octet rule, and this is the reason why atomic F ($[He]2s^22p^5$) can easily receive an extra electron producing ionic F^- ($[He]2s^22p^6$=[Ne]) or atomic Na ($[Ne]3s^1$) prefers to loose an electrons producing ionic Na^+ ($[He]2s^22p^6 = [Ne]$). The electron-dot structure of an atom or a molecule is a visual representation of the electronic arrangement around an atom or in a molecule.

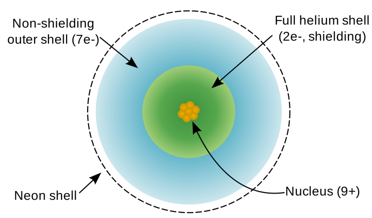


Figure 10.1: Valence electrons in a Ne atom

`Discussion:` oil spills on your shirt on diner. List three chemicals than can remove the stain

`Valence electrons` The electrons of an atom can be divided in 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 in an atoms is the same as the group number. As an example, hydrogen H belong to the group IA and hence has one valence electrons. Similarly, oxygen O belongs to the group VIA and therefore it has six valence electrons.

Sample Problem 88

Indicate the number of valence electrons for the following atoms: N, O, C and S.

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

STUDY CHECK

Indicate the number of valence electrons for the following atoms: Cl and B.

Answer: Cl ($7e^-$), B ($3e^-$).

The octet rule Atoms gain or loose electrons when they combine to form molecules. The octet rule says that each atom in a molecule is surrounded by eight electrons. There are two important exceptions to this rule as H is only surrounded by two electrons, and B by six.

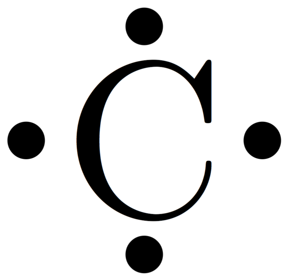


Figure 10.2: Lewis structure of C

Electron-dot structure of an atom In order to write the electron-dot structure of an atom, you just need to write down the symbol of the atom surrounded by the number of valence electrons one by one four directions, and if you have more than four electrons then add remaining electrons as pairs. For example, oxygen has six valence electrons and hence, the electron-dot structure would be $\cdot\ddot{\underset{\cdot}{\mathrm{O}}}\colon$ Similarly, for the case of fluorine the the electron-dot structure would be $\colon\ddot{\underset{\cdot}{\mathrm{F}}}\colon$ In the case of an ion, you need to add (if its an anion) or substract (if its a cation) electrons, and for example the electron-dot structure of O^{2-} is $\colon\ddot{\underset{\cdot\cdot}{\mathrm{O}}}\colon^{2-}$

Sample Problem 89

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: $\cdot\ddot{\underset{\cdot}{\mathrm{N}}}\cdot$ and $\cdot\dot{\underset{\cdot}{\mathrm{C}}}\cdot$. Cl^- has eight valence electrons, that is seven plus one, and hence its electron-dot structure will be $\colon\ddot{\underset{\cdot\cdot}{\mathrm{Cl}}}\colon^-$.

STUDY CHECK

Write down the electron-dot structure for N^{3-}

Answer: $\colon\ddot{\underset{\cdot\cdot}{\mathrm{N}}}\colon^{3-}$.

Electron-dot structure of diatomic molecules Now we will address how to build up electron-dot structures of molecules. The first step is (a) to set up the atoms in the molecule in the form of a line. After that, (b) you need to count the total number of valence electrons in the molecule, by adding the valence electrons of each atoms. Then you (c) calculate the pairs of electrons–the total number of valence electrons divided by two; pairs of electrons are represented by lines. In the following (d) you need to start distributing the pairs in the molecule in a very specific way: first connecting the surrounding atoms to the central atom, after placing pairs on top of the surrounding atoms and finally by placing the remaining pairs in the central atom. Each atom should be surrounded by four pairs with the exception of H and B.

Sample Problem 90

Construct the electron-dot structure of HCl.

SOLUTION

1. **Step one:** we first arrange the atoms in the molecule as H Cl.
2. **Step two:** now we count the number of valence electrons: H(1) and Cl(7)

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. $\mathrm{H{:}\ddot{\underset{\cdot\cdot}{Cl}}{:}}$, using lines instead of pairs (this is not necessary but makes the electron-dot structure look better) we obtain $\mathrm{H-\overline{\underline{Cl}}|}$.

STUDY CHECK

Construct the electron-dot structure of HF.

Answer: $\mathrm{H-\overline{\underline{F}}|}$.

Electron-dot structure of general molecules Now we will address how to build up electron-dot structures of more complex molecules. The first step is (a) 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, (b) you need to count the total number of valence electrons in the molecule, by adding the valence electrons of each atoms. Then you (c) calculate the pairs of electrons–the total number of valence electrons divided by two; pairs of electrons are represented by lines. In the following (d) you need to start distributing the pairs in the molecule in a very specific way (this is the key to building good electron-dot structures): first connecting the surrounding atoms to the central atom, after placing pairs on top of the surrounding atoms and finally by placing the remaining pairs in the central atom. Each atom should be surrounded by four pairs (this is the octet rule) with the exception of H and B as they do not follow the octet rule. When you have the final electron-dot structure, the pairs of electrons (or lines) that connect two atoms are called *bonds*, whereas the pairs not involved in an connection are called *lone pairs*. A very important note is that, at this point, is not that 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 in the surroundings) is correct.

Sample Problem 91

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. $\mathrm{H{:}\ddot{\underset{..}{O}}{:}H{:}}$ and using lines instead of pairs (this is not necessary but makes the electron-dot structure look better) we obtain $\mathrm{H-\overline{\underline{O}}-H}$. 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: $\mathrm{H-\overline{N}-H}$; three bonds and one lone pair.
$\mathrm{\quad\ \ |}$
$\mathrm{\quad\ \ H}$

Atomic charges in a molecule In order to build up the electron-dot structures of a molecule you needs to count the number of valence electrons. Each atom bring a different number of valence electrons to a molecule. For example, H brings one whereas O brings two. When you arrange the electron pairs in the molecule, each atoms should have the number of electrons that they bring. For example in this electron-dot structure $\mathrm{H{:}\ddot{\underset{..}{Cl}}{:}}$ the hydrogen atom brings one electrons to the molecule, and in the molecule the atom owns one electrons, as in $\mathrm{H{:}}$ one of the dots belong to H and the other belongs to the Cl–the electrons are shared in a covalent bond. In the same way, the Cl atom brings seven electrons and in the molecule it owns seven electrons, and has six electrons here $\mathrm{\ddot{\underset{..}{Cl}}{:}}$ plus the one that shared with the hydrogen atom. In another words, the $\mathrm{H{:}\ddot{\underset{..}{Cl}}{:}}$ electron-dot structure is the combination of $\mathrm{H\cdot}$ and $\mathrm{\cdot\ddot{\underset{..}{Cl}}{:}}$. We say that the charges on each atom are zero, as each atom in the molecule owns the same number of electrons that it originally brings.

Sample Problem 92

Indicate the atomic charges on each of the atoms of $\mathrm{H-\overline{C}-H}$
$\mathrm{\qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad\quad\ |}$
$\mathrm{\qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad\quad\ H}$

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).

STUDY CHECK

Indicate the atomic charges on each of the atoms of $\mathrm{H-\overline{\underline{N}}-H}$

Answer: H(0) and N(-1). $\mathrm{H-\overline{\underline{N}}^{\ominus}-H}$

Multiple bonds Often times you are going to encounter electron-dot structures like $\mathrm{{:}N\equiv N{:}}$ in which the atoms are connected by means of a triple bond. You will also encounter double bonds $\mathrm{\ddot{O}=\ddot{O}}$ and single bonds. These are called multiple bonds and are formed while constructing electron-dot structures in order to avoid having atoms with atomic charges different than zero. Charged atoms in a molecule are in general not favorable. When you build electron-dot structures and you end up having large atomic charges, you can avoid that by moving electrons from the atom into the bond, hence creating multiple bonds.

Sample Problem 93

Construct the electron-dot structure of O_2.

SOLUTION

1. **Step one:** we first arrange the atoms in the molecule as O O .
2. **Step two:** now we count the total number of valence electrons, including all atoms: 2xO(6) that gives a total of twelve electrons.
3. **Step three:** let us count the pairs of electrons; we have twelve electrons and that is six pairs.
4. **Step four:** now we distribute the pairs, fist connecting the atoms $O-O$ (we have five extra pairs to distribute at this point).
5. **Step five:** we place the remaining pairs on top of the oxygen atoms $|\overline{O}-O\rangle$
6. **Step six:** now we calculate the charge on each atom $|\overline{O}^{\ominus}\cdot{}^{\oplus}O\rangle$
7. **Step seven:** in order to eliminate the charges we move lone pairs into the bond $\langle O=O\rangle$ Now the charges are zero and this is more important than imposing the octet rule.

STUDY CHECK

Construct the electron-dot structure of CO_2.

Answer: $\langle O=C=O\rangle$

10.2 Molecular shape

Molecules are arrangements of atoms, and these arrangements can have different form. Think about an H_2O molecule, which contains two hydrogen atoms and one oxygen. Knowing that both hydrogens are connected to oxygen by means of a covalent bond, one can envision several molecular geometries such as $H-\overline{O}-H$ or maybe $\overset{\overline{O}}{H\diagup\;\diagdown H}$. The goal of this section is to identify the geometry of a given molecule. In order to do this, the electron-dot structure of the molecule are the key.

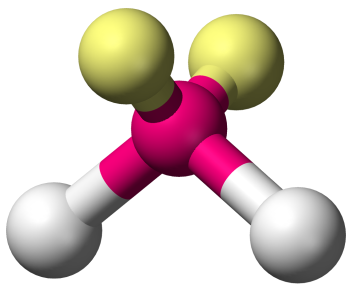


Figure 10.3: tetrahedral molecule

`ABE Molecular code` If the molecule contains two atoms, there is only a possible geometry these two atoms can exhibit, and this is a linear arrangement. For the case of more complex molecules, in order to identify the geometry you need to figure out the ABE code of the molecule. In this code B refers to the number of atoms connected to the central atom in a molecule, and E is the number of lone pairs on the central atom. For example, the electron-dot $H-\overline{O}-H$ structure has two bonds with the central atom B_2 and two lone pairs on top of the central atom E_2 and hence the ABE code of the molecule would be AB_2E_2. Another example the ABE code for ammonia $\mathrm{H-\overline{N}-H}$ (with H below N) would be AB_3E, as the molecule has three atoms connected to the central nitrogen and and N has a single lone pair. You can find a list of the equivalence between ABE codes and the molecular geometry in Table 10.4. In order to predict the geometry of a molecule, once you have the ABE code, Table 10.4 will give you the geometry. For example, an AB_2 molecule will be linear, whereas an AB_2E_2 is bent. The bond angles are also indicated in the table, and for example a CO_2 molecule, which will be linear will have a 180° angle. This means both C-O bonds will form a line.

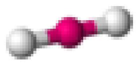
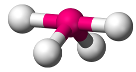
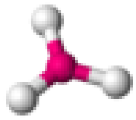
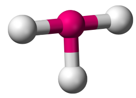
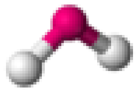
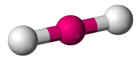
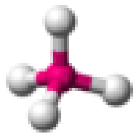
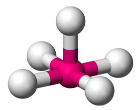
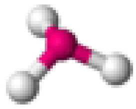
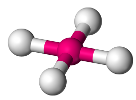
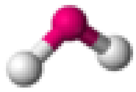
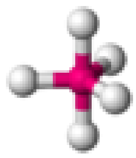
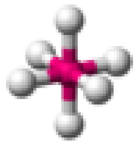
ABE Code	Molecular shape	Bond Angle	3D model	ABE Code	Molecular shape	Bond Angle	3D model
AB_2	Linear	180°		A_4B	seesaw	180°,120°, 90°	
AB_3	Trigonal Planar	120°		A_3B_2	T-shaped	90°, 180°	
AB_2E	Bent	120°		A_2B_3	Linear	180°	
AB_4	Tetrahedral	109°		A_5B_1	square pyramidal	90°	
AB_3E	Trigonal pyramidal	109°		A_4B_2	square planar	90°, 180°	
AB_2E_2	Bent	109°					
AB_5	trigonal bipyramidal	90°, 120°,180°					
AB_6	octahedral	90°, 180°,180°					

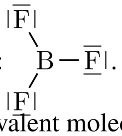
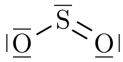
Figure 10.4: Molecular geometries

Add Table 10.4 into your flashcard.

Sample Problem 94

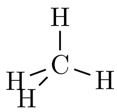
Identify the geometry of the following molecules: BF_3 and SO_2.

SOLUTION

We need first the electron-dot structure of both molecules. For BF_3 $|\underline{\overline{F}}-B-\underline{\overline{F}}|$ (with $|\overline{F}|$ above B). 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 $|\overline{O}-\overline{S}=\underline{O}|$ and its class is AB_2E. Hence the molecular geometry is linear and the molecule should be drawn as 

STUDY CHECK

Identify and draw the geometry of methane (CH_4).

Answer: Tetrahedral 

10.3 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 the bond. Polar bonds results in the existence of a permanent dipole moment that makes a molecule polar. Polar molecules can interact with polar molecules and mix.

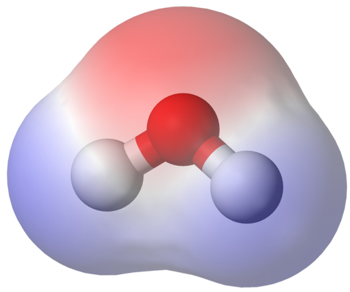
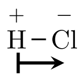
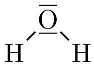


Figure 10.5: Electrostatic potential for the H_2O molecule showing molecular regions of positive and negative charge

`Bond polarity` Let us compare two different molecules: H_2 and HCl. We say H_2 is a non-polar molecule. The reason for this is that each atom in the covalent H-H bond equally share the electrons. Differently, HCl is a polar molecule, as H is an electropositive atom and Cl is electronegative. That implies that in the H-Cl covalent bond each atoms shares the electrons in the bond differently. H will be less prone to attract the electrons and Cl would tend to attract the bond electrons more than H. The result would be that the electrons in the bond would belong more to Cl than to H. Another consequence is that the molecule would have a permanent dipole–a permanent charge distribution–result of a uneven charge distribution in the chemical bond. We represent excess of charge as on Cl as $Cl^{\delta-}$ and electron deficiency in H as $H^{\delta+}$. The polarity of the bond is represented as:



`Polarity of diatomic molecules` Molecules can either be polar or non-polar. The polarity of diatomic molecules only depends on the nature of the atoms that forms the molecule. If the atoms in the molecule are the same (e.g. H_2 or O_2), then the molecule would be non-polar. If the atoms are different then the molecule would be polar. Examples are H_2 a nonpolar molecule and HCl or HBr, both polar molecules. You can apply the same concept to a bond inside a molecule. The C-O bond in a CO_2 molecule is a polar covalent bond, as C and O have different electronegativities.

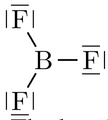
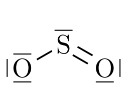
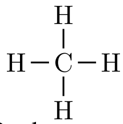
`Polarity of larger molecules` The polarity of larger molecules would depend on the molecular geometry. Let us analyze the case of CO_2. Each of the C-O bonds on the molecule are polar bonds. However, CO_2 is a linear molecule $\langle O=C=O \rangle$ and the polarity of each C-O bonds compensate so that at the end the molecule is polar. For the H_2O case, again, the H-O bond is polar. However, the molecule is bent and looks just like  . Both H-O bonds do not compensate as they point in different direction and the directions do not cancel out what makes the water molecule to be a polar molecule.

Sample Problem 95

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:

  and 

The bonds on BF_3 and SO_2 do not cancel out, as they do not point in opposite directions. Hence these two molecules are polar. On the other hand, the bonds on methane cancel each other out and hence even when the C-H bond is polar, the molecule 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 When you mix two different liquids or even gases, polarity is the key for the mixing process. If the molecules have the same polar character they will be able to mix, whereas they will not mix when the polar character is different. This section will cover several examples of mixing an polarity.

Molecules with the same polarity Water (H_2O) is a polar molecule. The ABE type of water of AB_2E_2 and hence its geometry is bent. That means both H-O bonds, which are polar, do not compensate with each other. Hence, the molecule will have a dipole moment and hence will be polar. Methanol (CH_3OH) is a polar molecule as well. The central atom of the molecule (C) is connected to three hydrogens and a 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 the same polarity CCl_4 is a nonpolar molecule, and H_2O is a polar molecule. As both have different polar character they will not mix together. If you mix water and CCl_4, tow 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 together. Soap has a polar and non-polar part. In order to remove oil from water, soap helps mixing both polar water and nonpolar oil.

10.4 Hybrid orbitals

Lewis structures are just representations of the bonds in a molecule. These representations are based on the localized electron bonding model, that assumes that molecules are made of atoms that share pairs of electrons. This section will addressed more advanced models. In particular hybrid orbitals are mixtures of atomic orbitals in a molecule. Atoms contains electrons and these electrons are located in atomic orbitals. We call these atomic orbitals as they belong to atoms. When a few atomic orbitals mix together, they hybridize, that is they mingle forming combinations of orbitals called hybrid. For example, when a s orbital hybridizes with a a single p orbitals, we obtain a hybrid sp orbital. At the same time, the s orbital can hybridize with two different p orbitals forming a sp^2 hybrid orbital. Hybrid orbitals are just a continuation of lewis structures and one can only obtain the hybrid orbitals of an atom in a molecule by means of the lewis structure. Therefore, at this point is very important you master the construction of lewis structures.

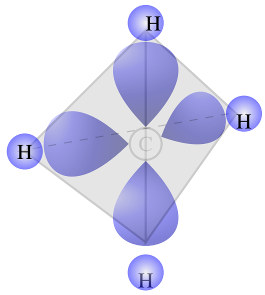
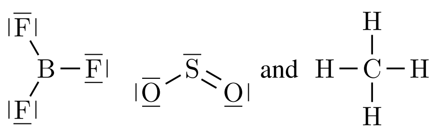


Figure 10.6: The molecule CH_4 with class AB_4 has hybridization sp^3. Four atomic orbitals are combined.

`From ABE code to hybridization` In order to obtain the hybridization of an atomic center in a molecule we just need the ABE code and Table 10.7. For example, if the code of a molecule is AB_4, the hybridization of the molecule will be sp^3. Similarly, if the class is AB_3 the hybridization will be sp^2 and in this case an empty p orbital will remain in the bond–mind there are three different p orbitals: p_x, p_y and p_z. Another example, would be a molecule with class AB. In this case, the hybridization will be sp and two empty p orbital will remain in the bond. A final example would be a molecule with class AB_4E. This time, the hybridization would be sp^3d^2. Mind that in general the number of hybrid orbitals correspond to adding the E and B from the class. For example, AB_4E_2, we have two E and four B with a total of six orbitals, hence we will need a s , three p's and two d's.

Sample Problem 96

Given the following Lewis structures, identify the hybridization of the central atom:



SOLUTION

In order to identify the hybrid character of the central atom, we first need to obtain the ABE code. For BF_3 the class is AB_3, for SO_2 is AB_2E and finally for CH_4 is AB_4. The number of electron regions for BF_3 is three. Therefore we would need three hybrid orbitals: $sp2$. An empty p orbital will remain unused in the bond. For SO_2 we need three electron regions and hence the hybridization of the central atom will also be $sp2$. For the case of methane, the hybridization will be $sp3$, as the molecule has four electron regions.

STUDY CHECK

Identify the hybridization of the central atom for the following molecules: O_2 and NH_3.

Answer: sp and $sp3$.

Add Table 10.7 into your flashcard.

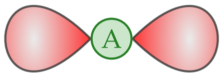
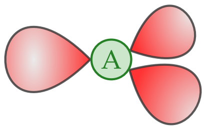
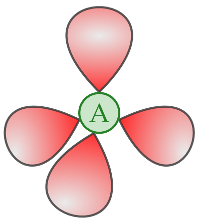
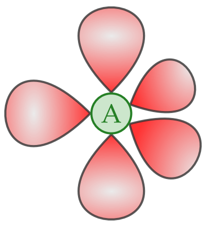
ABE Code	Electron Regions	Hybrid	Shape	Bond Angle
AB_2, ABE	2	sp		180°
AB_3, AB_2E,ABE_3	3	sp^2		120°
AB_4, AB_3E,AB_2E_2,ABE_3	4	sp^3		109.5°
AB_5, AB_4E,AB_3E_2,AB_2E_3	5	sp^3d		90° and 120°

Figure 10.7: Equivalency between the ABE code and the orbital hybridization.

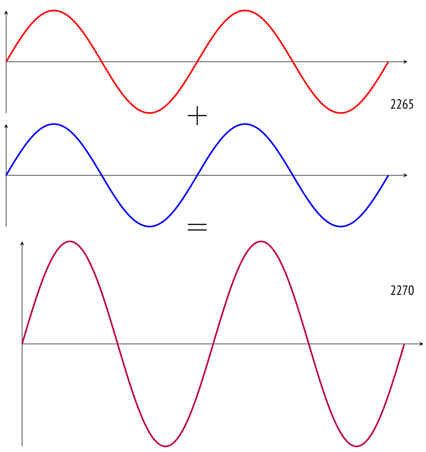


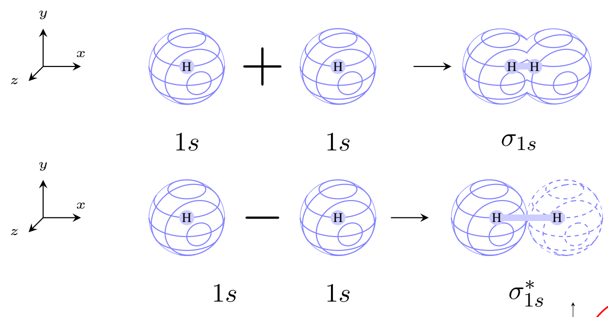
Figure 10.8: Constructive combination of two waves. The combination results in a higher intensity wave.

10.5 Molecular orbital theory

The molecular orbital theory is the most advanced bonding theory able to describe bond energies and bond lengths. Atomic orbitals are waves. When combining two waves one can obtain two possible results: a constructive combination and destructive combination. The molecular orbital theory assumes that atomic orbitals combine to form molecular orbitals. For every two atomic orbitals you can obtain two possible molecular orbitals: one is called bonding orbital, result from the constructive combination, and an other one called antibonding orbital, resulting from the destructive combination. In this section we will learn how to interpret molecular orbital diagrams.

`Bonding and antibonding orbitals` Atomic orbitals (AOs) combine to produce molecular orbitals (MOs). The combination of two atomic orbitals results in two new molecular orbitals: a bonding orbital and a antibonding orbital. Bonding MOs are more stable than the corresponding atomic orbitals. Antibonding MOs are less stable–they have a higher more positive energy–than the corresponding AOs. Antibonding orbitals are normally labeled with a * sign. Let us analyze both combinations of a $1s$ orbital. We can add both $1s$ orbital and the result is a bonding orbital, or we can substract both $1s$ orbitals and the result is an antibonding orbital, as the electron density cancels.

Figure 10.9: Bonding and antibonding σ orbitals resulting of combining two $1s$ atomic orbitals of Hydrogen.



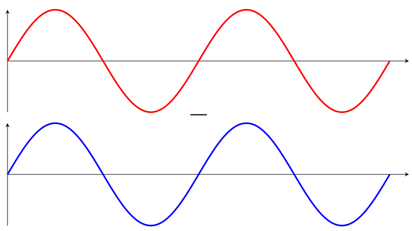
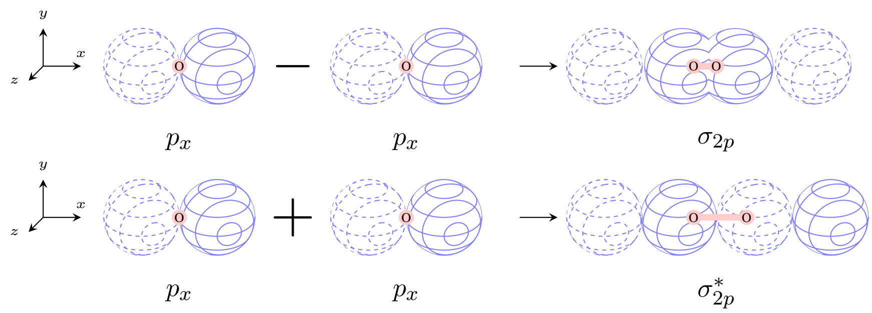


Figure 10.10: Destructive combination of two waves.

`Sigma and pi orbitals` Let's analyze now the mixing of two $2p_x$ orbitals of two oxygen atoms in order to make an O_2 molecule. Mind that p orbitals look like dumbbells and each side of the dumbbell is called lobe. In the p_x orbital the positive lobe is in the right side and the negative love on the left side. When combining both $2p_x$ if we add both orbitals we obtain a bonding orbital and if we subtract them we obtain an antibonding orbital. Both of these orbitals are called σ orbitals, as the lobes of the orbitals mixing go though the axes of the molecule being formed–the molecule is located in the X axis. Differently, if we combine two $2p_y$ orbitals we will obtain two π orbitals, as the lobes of the molecular orbital is perpendicular the axes of the molecule being formed

Figure 10.12: Bonding and antibonding σ orbitals resulting of combining two $2p_x$ atomic orbitals of Oxygen.



The first orbital is bonding as both lobes overlap, whereas the second orbital is antibonding as both lobes cancel out.

`The case of H2` Let us analyze the case of the formation of the H_2 molecule from two Hydrogen atoms. Each H atom has one $1s$ orbital. Therefore, according to the Molecular orbital theory, both orbitals will combine to produce two MO's. When you combine s atomic orbitals, the resulting MOs are always sigma. Sigma refers to the symmetry of the orbital. Therefore, the resulting MOs will be: σ_{1s} and σ^*_{1s}. Each AO contains one electron, hence the set of MO's will also contain two electrons that will occupy the most stable σ_{1s}. The resulting MO diagram is below. In this diagram, the atomic orbitals of H are on the left and right, whereas the MO's re in the center. We can

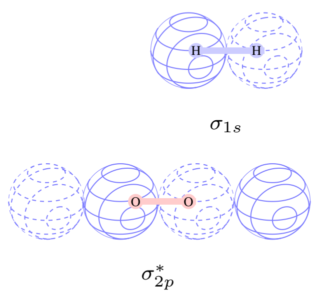


Figure 10.11: Examples of antibonding orbitals:

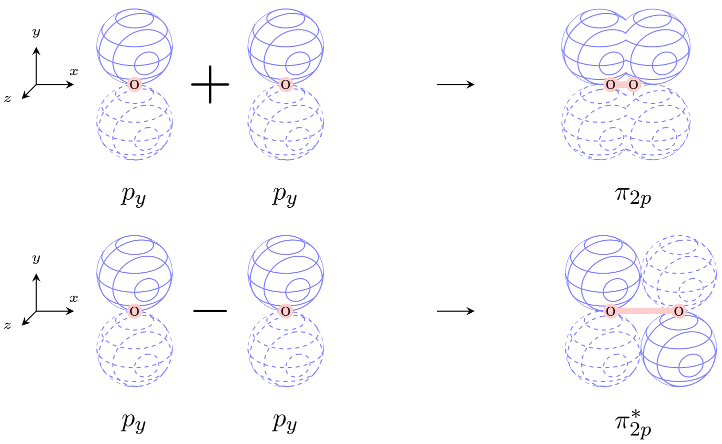
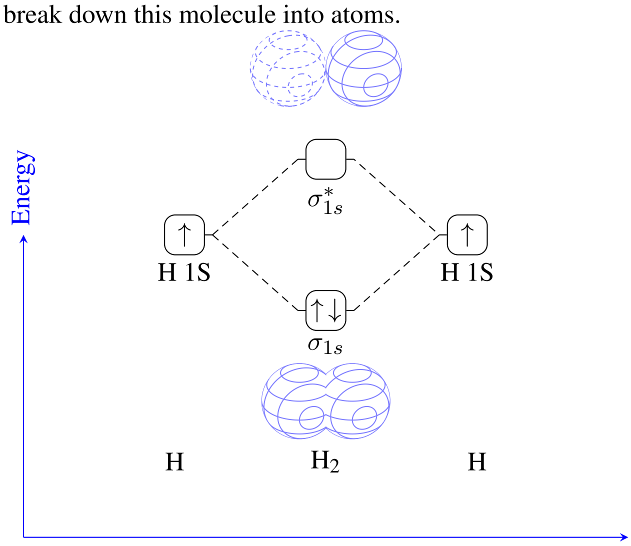
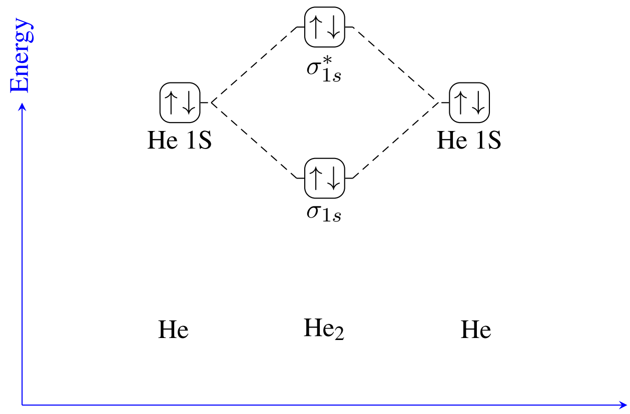


Figure 10.13: Bonding and antibonding π orbitals resulting of combining two $2p$ atomic orbitals of Oxygen.

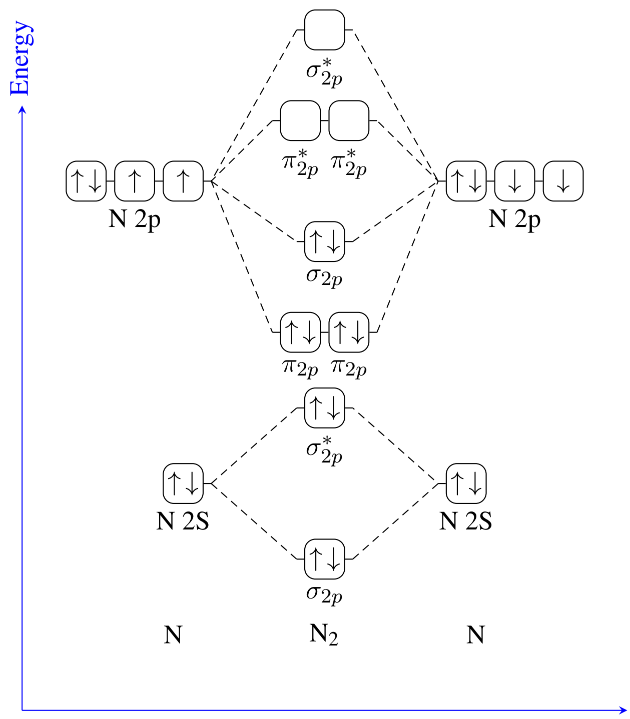
also give the MO configuration as: $\mathrm{H}_2 = \sigma^2_{1s}$. The hydrogen molecule is more stable than the separate hydrogen atoms. Why is that? the molecular orbitals of the molecule are lower in energy than the atomic orbitals of the hydrogen atoms. This means, the have more energy–as energy is negative that also means they are more stables. That is the reason why the hydrogen molecule is a stable existing molecule and takes energy to break down this molecule into atoms.



The case of He$_2$ molecule Let us analyze the case of the formation of the hypothetycal Ne_2 molecule from two He atoms. Each He atom has one $1s$ orbital with two electrons. Therefore, according to the Molecular orbital theory, both orbitals will combine to produce two MO's with a total of four electrons. The resulting MOs will be as well: σ_{1s} and σ^*_{1s}. This time, MO configuration is: $\mathrm{He}_2 = \sigma^2_{1s}\sigma^{*2}_{1s}$. In general antibonding orbitals are not stable. In the He molecule we stabilize the molecule by forming two σ^2_{1s} orbitals, but we also destabilize the molecule by forming σ^{*2}_{1s}. Hence the He_2 molecule will not be stable in compared to the atoms:



From MO diagram to MO configuration Obtaining a MO diagram is not obvious, and these diagrams can only be obtained after very complicated quantum mechanics simulations. However, after the MO diagram is given, one can obtain the MO configuration. From this configuration we can calculate two main properties: the bond order–related to the length of the molecule–and the magnetic character of the molecule. Let us use the case of N_2:



In this diagram, the lower MO's are the most stables and should be filled first. The higher MO are less stable and they are listed in the right side of the MO configuration. For example, the MO configuration of N_2 would be:

$$N_2 = \sigma^2_{2s}\sigma^{*2}_{2s}\pi^4_{2p}\sigma^2_{2p}$$

Bond order of a MO configuration Lets go back to the MO configuration for N_2. In this configuration we have some of the electrons occupying bonding MO and other occupying antibondoing MO's. The bond order is just the number of bonding electrons–the number of electrons occupying bonding MO's–minus the number of antibonding electrons–the number of electrons occupying antibonding MO's–divided by two. The formula is:

Add this formula to your flashcard.

$$\boxed{BO = \frac{(n - n^*)}{2}}$$ Bond Order

where:

n is the number of electrons occupying bonding MO's

n^* is the number of electrons occupying antibonding MO's

The bond order is related to the stability of the molecule and to the length of its chemical bond. The larger the bond order the most stable is the molecule as more electrons occupy bonding orbitals. The larger the bond order the smaller the chemical bond, and the atoms are more loose.

Sample Problem 97

Given the following MO configurations: (a) $\sigma_{2s}^2\sigma_{2s}^{*2}\pi_{2p}^4\sigma_{2p}^1$ and (b) $\sigma_{2s}^2\sigma_{2s}^{*2}\pi_{2p}^4\sigma_{2p}^2\pi_{2p}^{*3}$. Calculate the bond order and compare the length of the chemical bond of both molecules.

SOLUTION

The bond order is the number of bonding electrons minus the number of antibonding electrons divided by two. For the first example, we have seven bonding electrons and tow antibonding. Hence the bond order will be 2.5. For the second example, we have eight bonding electrons and five antibonding. Hence the bond order will be 1.5. The larger the BO the smaller the bond, hence the second molecule has a smaller bond.

STUDY CHECK

Calculate the bond order for $\sigma_{2s}^2\sigma_{2s}^{*2}\pi_{2p}^4\sigma_{2p}^2$.

Answer: 3.

Paramagnetism and diamagnetism We can also predict the magnetic character of a molecule by means of the MO configuration. Paramagnetic molecules are attracted by magnetic fields, whereas diamagnetic molecules are repelled by magnetic fields. The paramagnetic character is due to the presence of unpaired electrons. For example: $\sigma_{2s}^2\sigma_{2s}^{*1}$ is a paramagnetic molecule as we have one unpaired electron in the σ_{2s}^* orbital. Differently, $\sigma_{2s}^2\sigma_{2s}^{*2}$ is a diamagnetic molecule, as it has no unpaired electrons.

Sample Problem 98

Given the following MO configurations, predict the magnetic character: (a) $\sigma_{2s}^2\sigma_{2s}^{*2}\pi_{2p}^4\sigma_{2p}^1$ and (b) $\sigma_{2s}^2\sigma_{2s}^{*2}\pi_{2p}^4\sigma_{2p}^2\pi_{2p}^{*3}$.

SOLUTION

The first example has an unpaired σ electron and hence it is paramagnetic. The second base also has a single unpaired electron, this time in the π_{2p}^* orbital. Mind π orbitals have capacity of four and hence can place two separate pairs of electrons.

STUDY CHECK

Given the following MO configurations, predict the magnetic character: (a) $\sigma_{2s}^2\sigma_{2s}^{*2}\pi_{2p}^4\sigma_{2p}^2$ and (b) $\sigma_{2s}^2\sigma_{2s}^{*2}\pi_{2p}^4\sigma_{2p}^2\pi_{2p}^{*2}$.

Answer: both diamagnetic.

CHAPTER 10

Electron-Dot structures of molecules

1. The electron-dot structure of HI is:

(a) $|\underline{\mathrm{H}}-\overline{\underline{\mathrm{I}}}|$
(b) $|\overline{\underline{\mathrm{H}}}-\overline{\mathrm{I}}$
(c) $|\overline{\underline{\mathrm{H}}}-\overline{\underline{\mathrm{I}}}$
(d) $\overline{\underline{\mathrm{H}}}-\overline{\underline{\mathrm{I}}}$
(e) $\mathrm{H}-\overline{\underline{\mathrm{I}}}|$

Ans: (e)

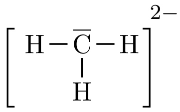
2. Draw the electron-dot structure of BH_3:

3. Draw the electron-dot structure of CH_4:

4. Draw the electron-dot structure of NH_4^+:

5. Draw the electron-dot structure of H_3O^+:

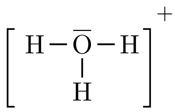
6. Indicate the charge of the C atoms in the following electron-dot structure:



(a) +1
(b) +2
(c) -1
(d) -2
(e) 5

Ans: (d)

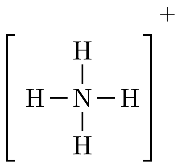
7. Indicate the charge of the O atoms in the following electron-dot structure:



(a) +1
(b) +2
(c) -1
(d) -2
(e) 4

Ans: (a)

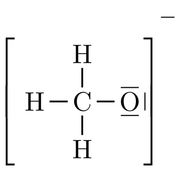
8. Indicate the charge of the O atoms in the following electron-dot structure:



(a) +1
(b) +2
(c) -1
(d) -2
(e) 4

Ans: (a)

9. Indicate the charge of the O atoms in the following electron-dot structure:



(a) +1
(b) +2
(c) -1
(d) -2
(e) 4

Ans: (c)

Molecular shape

10. Identify the molecular shape of the molecule: H_2

(a) Linera
(b) Trigonal planar
(c) Bent
(d) Tetrahedral
(e) Trigonal pyramidal

Ans: (a)

11. Identify the molecular shape of the molecule: $BeCl_2$

(a) Linera (b) Trigonal planar (c) Bent (d) Tetrahedral (e) Trigonal pyramidal

Ans: (a)

12. Identify the molecular shape of the molecule: BF_3

(a) Linera (b) Trigonal planar (c) Bent (d) Tetrahedral (e) Trigonal pyramidal

Ans: (b)

13. Identify the molecular shape of the molecule: NH_3

(a) Linera (b) Trigonal planar (c) Bent (d) Tetrahedral (e) Trigonal pyramidal

Ans: (e)

14. Identify the molecular shape of the molecule: CH_4

(a) Linera (b) Trigonal planar (c) Bent (d) Tetrahedral (e) Trigonal pyramidal

Ans: (d)

POLARITY

15. The H_2O molecule is

(a) polar (b) nonpolar

Ans: (a)

16. The HCl molecule is

(a) polar (b) nonpolar

Ans: (a)

17. The H_2 molecule is

(a) polar (b) nonpolar

Ans: (b)

18. The NH_3 molecule is

(a) polar (b) nonpolar

Ans: (a)

HYBRID ORBITALS

19. Indicate the hybridization of: NH_3

(a) sp (b) sp^2 (c) sp^3 (d) sp^3d (e) sp^3d^2 (f) sp^3d^3 (g) sp^3d^4 (h) sp^3d^5

Ans: (c)

20. Indicate the hybridization of: CH_4

(a) sp (b) sp^2 (c) sp^3 (d) sp^3d (e) sp^3d^2 (f) sp^3d^3 (g) sp^3d^4 (h) sp^3d^5

Ans: (c)

21. Indicate the hybridization of: H_2O

(a) sp (b) sp^2 (c) sp^3 (d) sp^3d (e) sp^3d^2 (f) sp^3d^3 (g) sp^3d^4 (h) sp^3d^5

Ans: (c)

22. Indicate the hybridization of: $\mathrm{H_2C=O}$

(a) sp (b) sp^2 (c) sp^3 (d) sp^3d (e) sp^3d^2 (f) sp^3d^3 (g) sp^3d^4 (h) sp^3d^5

Ans: (b)

23. Indicate the hybridization of: $\mathrm{H-C\equiv C-H}$

(a) sp (b) sp^2 (c) sp^3 (d) sp^3d (e) sp^3d^2 (f) sp^3d^3 (g) sp^3d^4 (h) sp^3d^5

Ans: (a)

24. Indicate the hybridization of: XeF_2

(a) sp (b) sp^2 (c) sp^3 (d) sp^3d (e) sp^3d^2 (f) sp^3d^3 (g) sp^3d^4 (h) sp^3d^5

Ans: (d)

MOLECULAR ORBITAL THEORY

25. Using the MO order provided below, obtain the MO configuration for B_2:
$\sigma_{2s}\sigma^*_{2s}\pi_{2p}\sigma_{2p}\pi^*_{2p}\sigma^*_{2p}$

Ans: $\sigma^2_{2s}\sigma^{*2}_{2s}\pi^2_{2p}$

26. Using the MO order provided below, obtain the MO configuration for C_2:
$\sigma_{2s}\sigma^*_{2s}\pi_{2p}\sigma_{2p}\pi^*_{2p}\sigma^*_{2p}$

Ans: $\sigma^2_{2s}\sigma^{*2}_{2s}\pi^4_{2p}$

27. Using the MO order provided below, obtain the MO configuration for O_2:
$\sigma_{2s}\sigma^*_{2s}\sigma_{2p}\pi_{2p}\pi^*_{2p}\sigma^*_{2p}$

Ans: $\sigma^2_{2s}\sigma^{2*}_{2s}\sigma^2_{2p}\pi^4_{2p}\pi^{2*}_{2p}$

28. Using the MO order provided below, obtain the MO configuration for F_2^+:
$\sigma_{2s}\sigma^*_{2s}\sigma_{2p}\pi_{2p}\pi^*_{2p}\sigma^*_{2p}$

Ans: $\sigma^2_{2s}\sigma^{2*}_{2s}\sigma^2_{2p}\pi^4_{2p}\pi^{3*}_{2p}$

REVIEW PART E

1. How many valence electrons are in the electron-dot structure of H_2O?
   (a) 2 (b) 4 (c) 6 (d) 8 (e) 10
2. Which of the following substances contains a polar covalent bond?
   (a) H_2 (b) CH_4 (c) NaCl (d) MgF_2 (e) N_2
3. Which one of the following compounds will NOT be soluble in water?
   (a) NaOH (b) PbS (c) K_2SO_4 (d) $LiNO_3$ (e) $MgCl_2$
4. What is the hybridization of the S atom in SF_4?
   (a) sp^2 (b) sp^3 (c) sp^3d (d) sp^3d^2 (e) spd
5. What is the hybridization of the xenon atom in XeF_4?
   (a) sp^2 (b) sp^3 (c) sp^3d (d) sp^3d^2 (e) spd
6. The molecular orbital diagram for the carbide ion ($C_2{}^{2-}$) would show which of the following molecular orbitals?
   (a) $(\sigma_{2s})^2(\sigma_{2s}*)^2(\pi_2p)^4(\sigma_{2p})^2(\pi_{2p}*)^4(\sigma_{2p}*)^2$
   (b) $(\sigma_{2s})^2(\sigma_{2s}*)^2(\pi_{2p})^2$
   (c) $(\sigma_{2s})^2(\sigma_{2s}*)^2(\pi_{2p})^4(\sigma_{2p})^2(\pi_{2p}*)^4$
   (d) $(\sigma_{2s})^2(\sigma_{2s}*)^2(\pi_{2p})^4(\sigma_{2p})^2$
   (e) $(\sigma_{2s})^2(\sigma_{2s}*)^2(\pi_{2p})^4$
7. When a double bond is formed between two atoms, one of the bonds is a sigma bond and the other is a pi bond. The π bond is created by the overlap of...
   (a) sp^2 hybrid orbitals
   (b) sp^3 hybrid orbitals
   (c) p orbitals
   (d) s orbitals
   (e) d orbitals
8. For the N2 molecule, there are .... electrons in the σ_{2p} bonding orbital?
   (a) 0. (b) 1. (c) 2. (d) 3. (e) 4.
9. Molecular orbital theory predicts the $O_2{}^-$ ion (a minus one charge) has:
   (a) no unpaired electrons.
   (b) one unpaired electron.
   (c) two unpaired electrons.
   (d) three unpaired electrons.
   (e) six unpaired electrons.
10. The N_2 molecule is:
    (a) paramagnetic. (b) diamagnetic. (c) submagnetic. (d) supermagnetic. (e) Superbowl magnetic.
11. Molecular orbital theory predicts the $F_2{}^{2+}$ ion has a bond order of:
    (a) 0.0 (b) 0.5 (c) 1.0 (d) 1.5 (e) 2.0
12. Draw the molecular orbital diagram for the molecular ion, $N_2{}^+$. The number of electrons in the σ_{2p} molecular orbital is:
    (a) 0 (b) 1 (c) 2 (d) 3 (e) 4
13. What is the bond order in $O_2{}^+$?
    (a) 3.5 (b) 2.0 (c) 1.5 (d) 2.5 (e) 0
14. The number of unpaired electrons in the B_2 molecule is...
    (a) zero (b) 1 (c) 2 (d) 3 (e) 4
15. Antibonding molecular orbitals are produced by
    (a) constructive interaction of atomic orbitals.
    (b) destructive interaction of atomic orbitals.
    (c) the overlap of the atomic orbitals of two negative ions
    (d) all of these
    (e) none of these
16. Green light has a wavelength of 5.50×10^2 nm. The energy of a photon of green light is
    (a) 3.64×10^{-34}J (b) 2.17×10^5J (c) 3.61×10^{-19}J (d) 1.09×10^{-27}J (e) 5.45×10^{12}J
17. The number of electron levels in a magnesium atom is
    (a) 1. (b) 2. (c) 3. (d) 4. (e) 5.
18. The maximum number of electrons that may occupy the third energy level is
    (a) 2. (b) 8. (c) 10. (d) 18. (e) 32.
19. What is the element with the electron configuration $1s^22s^22p^63s^23p^5$?
    (a) Be (b) Cl (c) F (d) S (e) Ar
20. Valence electrons are electrons located
    (a) in the outermost energy level of an atom.
    (b) in the nucleus of an atom.
    (c) in the innermost energy level of an atom.
    (d) throughout the atom.
    (e) in the first three shells of an atom.
21. How many f orbitals have n=3?
    (a) 0 (b) 3 (c) 5 (d) 7 (e) 1
22. Which of the following electron configurations is impossible?
    (a) $1s^22s^22p^63s^23p^1$ (b) $1s^22s^42p^63s^23p^3$ (c) $1s^22s^22p^63s^23p^5$ (d) $1s^22s^22p^63s^23p^6$ (e) $1s^22s^22p^63s^23p^3$
23. The ionization energy of atoms

oss a period.
n within a group.
n within a group.
ng down within a group.

elements has the lowest electroneg-

(d) O
(e) F

F, Li, and Na., the element with the

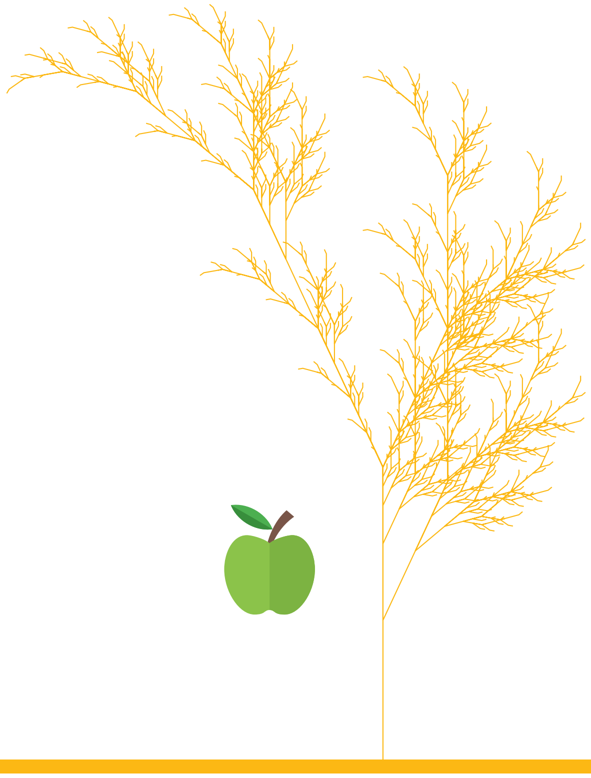
(d) Li.
(e) Na.

as in order of increasing ionization

(d) $Na < Li < F < O < C$
(e) $Na < Li < C < F < O$

Answers:

1. (d)	**2.** (b)	**3.** (b)	**4.** (c)
5. (d)	**6.** (d)	**7.** (c)	**8.** (c)
9. (b)	**10.** (b)	**11.** (e)	**12.** (b)
13. (b)	**14.** (c)	**15.** (b)	**16.** (c)
17. (c)	**18.** (d)	**19.** (b)	**20.** (a)
21. (a)	**22.** (b)	**23.** (b)	**24.** (a)
25. (e)	**26.** (b)		



College Chemistry

A Comprehensive Set of Imperfect Notes

This set of lectures intend to cover the first semester of a College Chemistry class. The intention here is to present the 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 as well as a periodic table. The organization of the notes is based on 10 chapters and five parts, each made of two chapters. This organization is intended to help the reader digest the large content typically covered in a General Chemistry class. Every part ends with a review quiz that assesses content. Finally, this set of notes are made to complement and not replace any existing textbook.