GOB Chemistry Experiments

A Comprehensive Set of Imperfect Notes

The Open Education Initiative at CUNY

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Periodic Table of Chemical Elements

Z mass
Symbol
Name

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

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

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

Contents

To The Reader

Dear Reader

It is with great pleasure that we offer the alpha edition of this*Chemistry Laboratory Manual*.

Chemistry embraces everything around us such as the clothes we wear, the air we breathe, or the batteries that power our cell phones, and today's chemistry is built on centuries of exploration and discovery. Most of the lab textbooks in the market contain useful information about the basic principles of chemistry that are needed for most science and engineering students, even after graduation. Whether it's a glossary of chemical terms or basic information about the chemistry principles, it is extremely handy to have those books around for further reference. There are also many good chemistry laboratory manuals in the market result of the work of devoted faculty. They normally contain a series of laboratory procedures or inquiry-based experiments developed and carefully tested throughout the years. Nevertheless, these are rarely reference material as they have instead a more limited use. Students extensively use these manuals, sometimes writing their measurements in worksheets attached, which makes it impossible to reuse them. But more importantly: laboratory manuals are expensive. With an average price of $50, they represent a considerable investment for a student just for single-use purchase. With all these in mind, this manual aims to alleviate these burdens by providing a set of standard chemistry experiments freely available for faculty and students.

The main part of the experiments in this manual deals with classical chemistry techniques such as reading a meniscus, making dilutions, or performing a filtration or titration. These are standard techniques and experiments such as acetic acid titration have been extensively validated in the literature. The reactants employed in the experiment are standard chemicals easily accessible in any general chemistry laboratory and the safety measures are minimal.

This manual is designed to encourage students to think and develop a solid understanding of chemistry by first building a qualitative understanding and then practicing basic chemistry techniques. Each experiment begins with a background section that reviews the basic ideas behind the experiment, providing sometimes worked examples to illustrate the concepts.

We are aware of the difficulty that students have with math, as well as with plotting and graphing. With that in mind, we have included several plotting exercises in the labs that will help students practice these techniques.

Last, but by no means least, any critical input from the reader will be well-received.

The authors.

EXPERIMENT 0

Chemistry and measurements

A. Goal

The goal of this laboratory experiment is to familiarize with estimating digits and accounting for significant figures in measurements through measuring in the chemistry laboratory.d

B. Materials

- ☐ 10ml measuring cylinder
- ☐ 50ml beaker
- ☐ any size stopper
- ☐ a spatula
- ☐ string
- ☐ set of measuring cylinders: 10, 25, 50, 100 and 250mL
- ☐ metallic cylinder
- ☐ 50mL-cylinder

C. Background

Significant Figures

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

Measured numbers

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

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

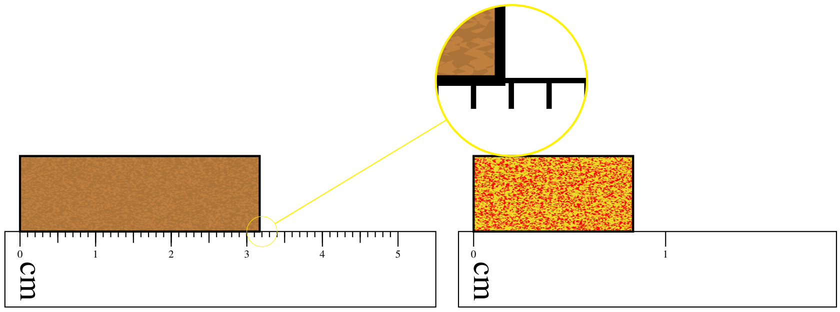

Figure 1 Some metric scales with two objects of different lengths. Measurements are (left) 3.15cm (right) and 0.8cm.

Reading menisci

Reading a liquid meniscus is similar to reading any measuring scale. There are two types of menisci (see Figure 2). A concave meniscus, which is what you normally will see, occurs when the molecules of the liquid are attracted to those of the container. This occurs with water and a glass tube. A convex meniscus occurs when the molecules have a stronger attraction to each other than to the container, as with mercury and glass. If the meniscus is concave, read at the lowest level of the curve. If the meniscus is convex, take your measurement at the highest point of the curve. Let us read the menisci from the image below. Readings are 16.0mL (left), 8.5mL (center left), 18.0mL (center right), and 18.5mL (right).

Exact numbers

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

Significant figures of numbers

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

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

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

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

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

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

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

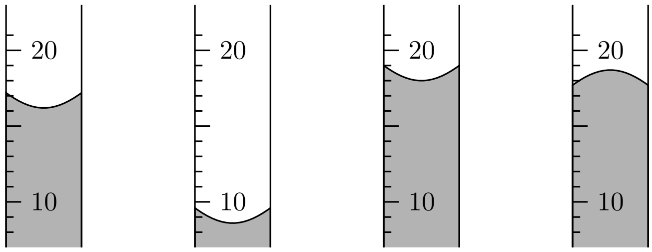

Figure 2 Some volumetric measurements in mL are presented in two types of meniscuses. The Left, center-right, and center-left are concave meniscuses, whereas the right image presents a convex meniscus. Readings are 16.0mL (left), 8.5mL (center left), 18.0mL (center right), and 18.5mL (right).

Sample Problem 1

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.

D. Procedure

1. Measuring mass The goal of this mini-experiment is to measure the mass of several objects with a scale. Overall, the goal is to learn how to use laboratory scales and how to properly report measurements. Moreover, this experiment will help you getting familiarized with a chemistry laboratory environment:

Step 1: – Locate the following objects: a 10ml measuring cylinder, a 50ml beaker, a stopper of any size and a spatula.

Step 2: – Measure the mass of each of the object using a scale. Make sure the scale is set to zero before you measure.

Step 3: – Write down the values listing the name of the object. Do not forget to indicate the unit of the measurement.

Step 4: – Indicate the measured figure (e.g. for a measure number 345.8g the estimated would be written as 0.8g) and the number of significant figures of the measurement.

Step 5: – Return each object to its original location in the lab.

2. Measuring length In this mini-experiment you will familiarize with a meterstick (or a inch ruler). I am aware you have seen one before but perhaps you have not noticed some of the nuances of this very useful measuring tool.Think about the meaning of the large and small lines on the meterstick (or a inch ruler). Discuss with your coworkers this information. What do the large lines represent? And the small lines?

Step 1: – Write down the length of the following items. Mind you need to measure these values in the lab. In the case of your height, you will find a meter stick at the lab, near the entrance.

Step 2: – You can use a string to measure the length of your wrist .

Step 3: – Write down the measured digit of each measurement and the number of significant figures. Do not forget to indicate the unit near the measurement.

Step 4: – Measure the length of the following line and write down the value on the results section.

3. Measuring volume In this mini-experiment learn how to properly read a volume measurement using chemistry equipment. In the chemistry lab, volume can be measured with a measuring cylinder. However, liquid forms a meniscus on this object due to a property of water called surface tension. Therefore, to be able to read volume you will have to properly read the meniscus and use volumetric scales.

Step 1: – Locate the lab setup with measuring cylinders of 10, 25, 50, 100 and 250 mL.

Step 2: – Fill each cylinder with a random quantity if water within the volume of the cylinders. The cylinders might be already filled.

Step 3: – For each cylinder, read the meniscus and write down the volume. Make sure you read between the lines of the cylinder to indicate the estimated digit of the measurement. Do not forget to write down the unit close to the measurement.

Step 4: – Indicate the estimated digit of the measurement and the number of significant figures.

5. Measuring volume by displacement The volume of liquids is easy to measure. However, the volume of solids is a harder property to measure experimentally. The goal of this mini-experiment is to measure the volume of an object (a cylindric piece of metal) by displacing the liquid on a measuring cylinder.

Step 1: – Find a metallic cylinder you want to measure volume.

Step 2: – Use a 50mL-cylinder large enough to easily fit the object. Make sure the object will not be stuck inside the cylinder.

Step 3: – Add water to the cylinder and write down the volume. Make sure you read the estimated digits (e.g. 50.50mL).

Step 4: – Place the object on the cylinder. You will see the level of water rise. Make sure the level is beyond the size of the object so that the object is fully submerged on water. If not, you will have to repeat the experiment adding more water initially on the cylinder.

Step 5: – Calculate the volume of the object by subtracting the the volume of water after the object is submerged–the final volume–and the volume before–the initial volume.

STUDENT INFO

Name: ______________________ Date: ______

Pre-lab Questions

Chemistry and measurements

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1. Fill the gaps with the full unit name, the abbreviation and the property measured. Indicate also whether the unit contains a prefix:

Full unit Name	**Abbreviation**	**Property measured**	**Prefix? (yes/no)**
Kilogram	________________	________________	________________
________________	mL	________________	________________
Degree Celsius	________________	________________	________________
________________	in	________________	________________

2. Indicate whether the following numbers indicate a measured value or an exact number. Measured numbers result from measurements, and exact numbers result from counting.

20Kg	________________	10.5 cm	________________
3 apples	________________	10°C	________________
30.5L	________________	90mL	________________
1cm=10^{-2}m	________________	4g	________________

3. Explain what are significant figures.

4. You measure the mass of a beaker using a scale and the results is 28.27g. Indicate the estimated digit of the measurement.

5. You measure the length of a measuring cylinder using a meter stick with a scale that indicates centimeter as well as millimeters and the results is 25.15cm. Indicate the estimated digit of the measurement.

STUDENT INFO	
Name: ____________________	Date: ______

Results
EXPERIMENT

Chemistry and measurements

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1. Measuring mass

Object Name	Mass	Estimated digit	# significant Figures

2. Measuring length

Object Name	Length	Estimated digit	# SFs
Length of your right foot			
Length of one of your fingernails			
Length of one of your wrist			
Line below			

3. Measuring volume

Object Name	Volume	Estimated digit	# SFs

5. Measuring volume by displacement

		Volume	Estimated digit	# SFs
Initial volume	1			
Final volume	2			
Object volume	2 - 1			

STUDENT INFO

Name: ______________________ Date: ______

Post-lab Questions

Chemistry and measurements

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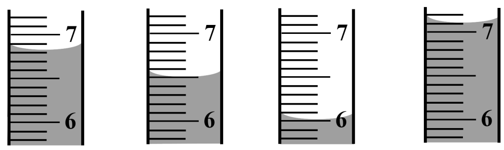
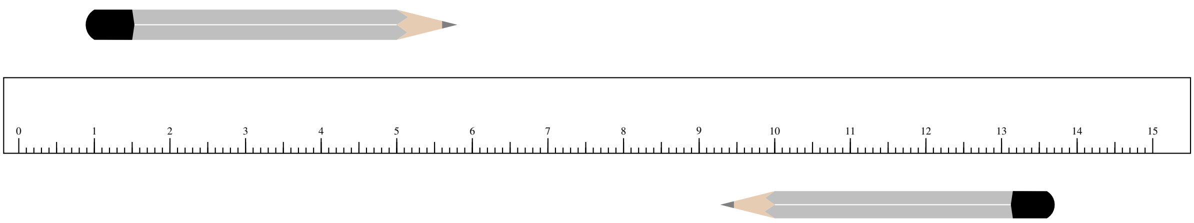
1. Indicate the measurement of the following meniscus (in mL):



2. Using the rule below in cm:



   (a) Measure the length of both pencils.

   (b) Indicate the estimated digits

3. Indicate the number of significant figures of the following measurements:

Measured number	**SFs**	**Measured number**	**SFs**
20.1Kg	____________	120.5 cm	____________
0.01 m	____________	100g	____________
0.010 s	____________	230.1dm	____________
5×10^{-5} dm	____________	6.500×10^{-1} dmg	____________

EXPERIMENT 0

Conversion Factors and Problem Solving

A. Goal

The goal of this laboratory experiment is to practice unit conversions and carry out calculations with the correct number of significant figures.

B. Materials

- ☐ rectangular wood piece
- ☐ string
- ☐ ruler
- ☐ 1L graduated cylinder
- ☐ 1qt jar

C. Background

Significant figures in calculations

Two different rules allow you to express the result of calculations with the correct number of figures.
☡ **Rule 1** ($+\ -$) *For additions or subtractions, the results has the same number of decimal places as the number with the least decimal places in the calculation.* For example:

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

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

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

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

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

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

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

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

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

Rounding

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

Sample Problem 2

Do the following calculation with the correct number of figures.

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

SOLUTION

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

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

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

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

STUDY CHECK

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

Answer: 0.35

Table 1 Different prefixes

Prefix	Symbol	Meaning	Value
exa	E	1000000000000000000	1×10^{18}
peta	P	1000000000000000	1×10^{15}
tera	T	1000000000000	1×10^{12}
giga	G	1000000000	1×10^{9}
mega	M	1000000	1×10^{6}
kilo	k	1000	1×10^{3}
hecto	h	100	1×10^{2}
deca	da	10	1×10^{1}
–	–	1	1×10^{0}
deci	d	0.1	1×10^{-1}
centi	c	0.01	1×10^{-2}
milli	m	0.001	1×10^{-3}
micro	μ	0.000001	1×10^{-6}
nano	n	0.000000001	1×10^{-9}
pico	p	0.000000000001	1×10^{-12}
femto	f	0.000000000000001	1×10^{-15}
atto	a	0.000000000000000001	1×10^{-18}

Using Conversion Factors

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

Removing or adding prefixes

Imagine that you need to remove a prefix from a unit, and convert 3 km (we will call this one the original unit) into meters (this is the final unit). First, you would need the conversion factor corresponding to the prefix (centi) from Table 1. Then you need to arrange the conversion factor by placing the prefix at the bottom of the fraction. This will cancel out the prefix in the original unit and the bottom part of the conversion factor, hence leaving the final unit on top of the conversion factor. The arrangement would be:

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

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

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

Sample Problem 3

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

SOLUTION

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

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

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

STUDY CHECK

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

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

Switching prefixes

To switch a prefix into another prefix, such as transforming 30 millimeters (30 mm) into centimeters (cm), you will need two different conversion factors: the first conversion factor will remove the original unit (mm) introducing an intermediate unit, meters (m), whereas the second conversion factor will remove the intermediate meter and introduce the final unit (cm). You will get the conversion factors from Table 1. You will arrange the first conversion factor so that the original unit cancels out with the bottom of the first conversion factor, giving you an intermediate unit. You will arrange the second conversion factor so that the intermediate unit cancels out with the bottom of the second conversion factor giving the final unit. For this example:

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

Sample Problem 4

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

SOLUTION

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

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

STUDY CHECK

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

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

Units of volume and area

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

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

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

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

Sample Problem 5

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

SOLUTION

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

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

STUDY CHECK

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

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

Table 2 Table containing some common unit equalities

Unit	Equality
Inches (in)-centimeters (cm)	$2.54^{\dagger}$ cm $=$ 1 in
miles (mi)-meters (m)	1 mi $=$ 1609.34m
minutes (min)-hours (h)	60 min $=$ 1 h
minutes (min)-seconds (s)	60 s $=$ 1 min
pound (lb)-grams (g)	454 g $=$ 1 lb
cubic centimeter (cm^3)-mililiters (mL)	1 mL $=$ 1cm^3
quart (qt)-milliliters (mL)	1 qt $=$ 946.353mL
Liter (L)-cubic decimeters (dm^3)	1 L $=$ 1dm^3
drops-mililiters* (mL)	1 mL $=$ 15 drops

* There are several definitions of a drop

† the number is exact

Liters and milliliters

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

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

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

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

Sample Problem 6

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

SOLUTION

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

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

STUDY CHECK

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

▶Answer: $40\cancel{L} \times \frac{1\cancel{mL}}{1 \times 10^{-3}\cancel{L}} \times \frac{1cm^3}{1\cancel{mL}} = 4 \times 10^4 cm^3$.

Using other equalities

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

Sample Problem 7

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

SOLUTION

We want to convert 20 inches into centimeters. The relationship between Inch and centimeter is given in Table 2. In order to do this, you need the conversion factor: $\frac{1in}{2.54cm}$ or $\frac{2.54cm}{1in}$. You will have to arrange the number (20 in) and the conversion factor in the following form:

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

STUDY CHECK

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

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

D. Procedure

1. Significant figures in additions and subtractions The goal of this mini-experiment is to familiarize with the use of significant figures in basic calculations. When faced with an addition or subtraction calculation, the rule says that the final number has to have the same number of decimal places as the number with the fewest decimal places. Carry the following calculations and give the result with the correct number of decimal places or significant figures.

Step 1: – Analyze each number separately and among all numbers identify the less number of decimal places.

Step 2: – Analyze each number separately and among all numbers identify the less number of significant figures (SFs).

Step 3: – Write down the final result with the correct number of decimals or SFs using the rounding rules (If the number you are rounding is followed by 5, 6, 7, 8, or 9, round the number up)

2. Significant figures in multiplications and divisions The goal of this mini-experiment is, again, to familiarize with the use of significant figures in basic calculations. When faced with multiplications and divisions, the rule says that the final number has to have the same number of SFs as the number with the fewest SFs. Carry the following calculations and give the result with the correct number of decimal places or significant figures.

Step 1: – Analyze each number separately and among all numbers identify the less number of decimal places.

Step 2: – Analyze each number separately and among all numbers identify the less number of significant figures (SFs).

Step 3: – Write down the final result with the correct number of decimals or SFs using the rounding rules (You can replace a digit by zero to eliminate significant figures: 123(3SF)$\approx$100(1SF))

3. Measuring volume In this mini-experiment learn how to properly compute volume using the right number of SF's.

Step 1: – Obtain a rectangular wood piece from the lab. Obtain one piece per team.

Step 2: – With a ruler measure the length of the sides of the rectangular piece of wood in cm.

Step 3: – Compute the volume by multiplying the length, heigh and depth using the right number of SF's and digits.

Step 4: – Compare your result with the other students in the team and write them down below. Do you get the same result?

4. Simple conversion factors This mini-experiment will help you out learn how to carry simple conversion factors. In particular, how to remove and add a prefix.

Step 1: – Fill the gap in the calculations displayed in the results section. Remember to place 1 in front of the unit with prefix (cm) and the corresponding power of ten in from of the unit (m).

5. Non-metric conversions This mini-experiment deals with non-metric units and their conversion to metric-based units. An example of this is inches which are 2.54cm. One can convert from non-metric In into centimeter–a metric-based unit. Below is a list of a few non-metric units

$$\boxed{1\text{in} = 2.54cm \quad 1\text{lb} = 454g \quad 1\text{qt} = 946mL}$$

Step 1: – Using a metric-based ruler and a string, measure the size of your wrist in cm. Write down your results in the table below.

Step 2: – Using an inch-based ruler and a string, measure the length of your wrist in In. Write down your results in the table below.

Step 3: – Set up the conversion factor below to convert cm into inches.

$$\_\_\_\_\_\_ cm \times \frac{\_\_\_\_\_\_}{\_\_\_\_\_\_} = \_\_\_\_\_\_ in$$

Step 4: – Write down the results in the table below.

Step 5: – Calculate the percent error using the formula (make sure you use absolute value). Write down your results in the table below:

$$\%\ \text{Error} = \left|\frac{2 - 1}{1}\right| \times 100$$

Step 6: – Compare your error with the other students in the team and write them down below. Do you get similar errors?

6. Non-metric conversions for volume This mini-experiment deals with non-metric volume units and their conversion to metric-based units. One can convert from non-metric qt (quart) into L–a metric-based unit. Below is a list of a few non-metric units

$$\boxed{1\text{L} = 1.057qt}$$

Step 1: – Measure 1qt of water and transfer it to a 1L graduated cylinder.

Step 2: – Read the volume meaurement from the 1L graduated cylinder.

Step 3: – Set up the conversion factor below to convert qt into L.

$$\_\_\_\_\_\_ qt \times \frac{\_\_\_\_\_\_}{\_\_\_\_\_\_} = \_\_\_\_\_\_ L$$

Step 4: – Write down the results in the table below.

Step 5: – Calculate the percent error using the formula (make sure you use absolute value). Write down your results in the table below:

$$\%\ \text{Error} = \left|\frac{2 - 1}{1}\right| \times 100$$

Step 6: – Compare your error with the other students in the team and write them down below. Do you get similar errors?

STUDENT INFO

Name: ____________________ Date: ______

Pre-lab Questions

Conversion Factors and Problem Solving

---

1. Fill in the gaps for the following conversion equalities:

1 Kg	=	______g
______ Km	=	10^3m
1 cm	=	______m
______ dm	=	10^{-1}m
1 Tb	=	______b
______ μL	=	10^{-6}L

2. Fill in the gaps for the following conversion factors:

$$\frac{1fs}{\_\_\_\_\_\_ s} \qquad \frac{\_\_\_\_\_\_ cm}{10^{-2}m} \qquad \frac{1nm}{\_\_\_\_\_\_ m} \qquad \frac{1Kcal}{\_\_\_\_\_\_ cal}$$

3. Round the following numbers to the indicated number of decimal places or significant figures. Mind, the rules for rounding say that if the first digit to be dropped is more or equal to five (0.262) the value of the retained digit should be increased by one ($\approx$ 0.3 one decimal place).

157.68	$\approx$	__________	(one decimal place)
47.807	$\approx$	__________	(two decimal places)
1200	$\approx$	__________	(one significant figure)

4. Do the following calculations with the correct number of significant figures. Mind when adding or substring numbers the results has to have the same number of digits as the number in the calculation with the fewest decimal places.

123.1+34.58	=	__________
45.567+2.24	=	__________

5. Do the following calculations with the correct number of significant figures. Mind when multiplying or dividing numbers the results has to have the same number of significant figures as the number in the calculation with the number of significant figures.

100×12	=	____________
$0.34/3.56$	=	____________

STUDENT INFO	
Name: ____________________	Date: ______

Results
EXPERIMENT

Conversion Factors and Problem Solving

---

1. Significant figures in additions and subtractions

Calculation	Fewest # of SFs	Fewest # of decimals	Result
$45.3 + 12.63$	____________	____________	____________
$45.3 + 12.23$	____________	____________	____________
$45.33 + 12.456$	____________	____________	____________
$45 + 12.12 - 23.2$	____________	____________	____________

2. Significant figures in multiplications and divisions

Calculation	Fewest # of SFs	Fewest # of decimals	Result
$1700/123$	____________	____________	____________
$0.1245 \times 2.00 \times 0.0367$	____________	____________	____________
$54.87 \times 4.56/0.4$	____________	____________	____________

3. Measuring volume

Length	Height	Depth	Volume, cm^3
____________	____________	____________	____________

4. Simple conversion factors

$$20m \times \frac{1cm}{\underline{\hspace{2cm}}\, m} = 2000cm \qquad 76g \times \frac{1Kg}{\underline{\hspace{2cm}}\, g} = 0.076Kg$$

$$40L \times \frac{1mL}{\underline{\hspace{2cm}}\, L} = 4 \times 10^4 mL \qquad 200\mu L \times \frac{10^{-6}L}{\underline{\hspace{2cm}}\, \mu L} = 2 \times 10^{-4} L$$

$$5m \times \frac{\underline{\hspace{2cm}}\, cm}{\underline{\hspace{2cm}}\, m} = 500cm \qquad 1000g \times \frac{\underline{\hspace{2cm}}\, Kg}{\underline{\hspace{2cm}}\, g} = 1Kg$$

$$0.4cm \times \frac{\underline{\hspace{2cm}}\, m}{\underline{\hspace{2cm}}\, cm} = 4 \times 10^{-3} m \qquad 100g \times \frac{\underline{\hspace{2cm}}\, Kg}{\underline{\hspace{2cm}}\, g} = 0.1Kg$$

$$300Gb \times \frac{\underline{\qquad} b}{\underline{\qquad} Gb} = \underline{\qquad} b \qquad 200mm \times \frac{\underline{\qquad} m}{\underline{\qquad} mm} = \underline{\qquad} m$$

$$50m \times \frac{\underline{\qquad} dm}{\underline{\qquad} m} = \underline{\qquad} dm \qquad 5g \times \frac{\underline{\qquad} Kg}{\underline{\qquad} g} = \underline{\qquad} Kg$$

$$300nm \times \frac{\underline{\qquad}}{\underline{\qquad}} = \underline{\qquad} m \qquad 500Kg \times \frac{\underline{\qquad}}{\underline{\qquad}} = \underline{\qquad} g$$

$$70Tb \times \frac{\underline{\qquad}}{\underline{\qquad}} = \underline{\qquad} b \qquad 500mL \times \frac{\underline{\qquad}}{\underline{\qquad}} = \underline{\qquad} L$$

5. Non-metric conversions

Measured Length (cm)	Measured Length (in)	Converted Length (in)	% Error
__________	__________	__________	__________
	1	2	

6. Non-metric conversions for volume

Measured volume (qt)	Measured volume (L)	Converted volume (L)	% Error
__________	__________	__________	__________
	1	2	

STUDENT INFO

Name: ______________________ Date: ______

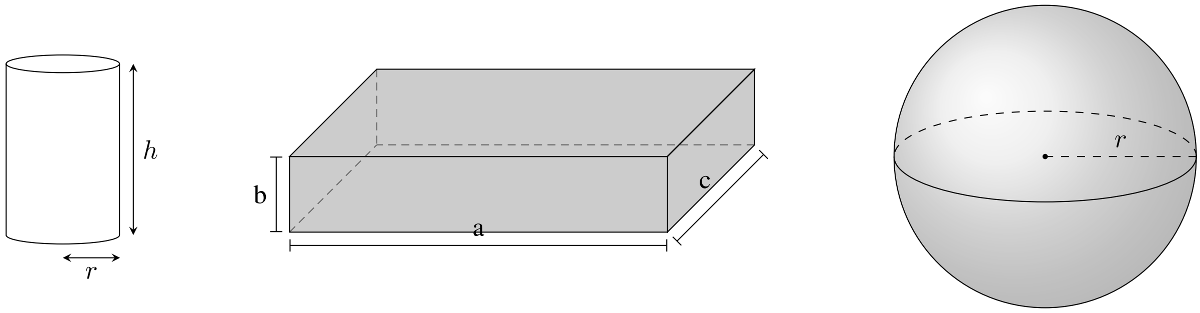
Post-lab Questions

Conversion Factors and Problem Solving

---

1. Convert $100\mu L$ into L.

2. Using a ruler in cm, calculate the volume of the following object with the correct number of digits or SFs:



$$v_{cylinder} = \pi r^2 \times h \qquad v_{cube} = a \times b \times c \qquad v_{sphere} = \frac{3}{4} \times \pi r^3$$

EXPERIMENT 0

Density and Specific Gravity

A. Goal

The goal of this laboratory experiment is to experimentally measure density as well as specific gravity for a liquid.

B. Materials

- ☐ 50ml, 100mL cylinder
- ☐ a 100mL (or 25mL) beaker
- ☐ metallic object
- ☐ unknown solution
- ☐ a set of pennies

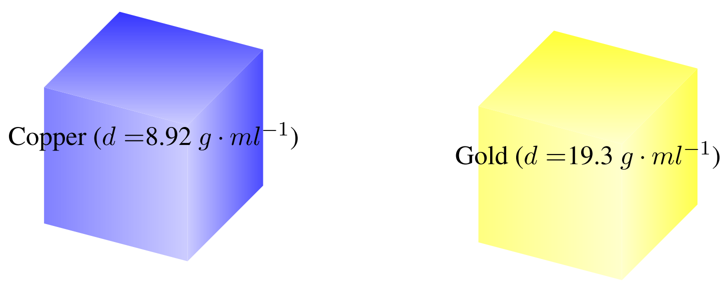
C. Background

Density

Density refers to the mass of a substance with respect to its volume. This is an unique property for each substance. Table 3 reports the density of numerous substances. Indeed, density is often used as an identification tag. The formula for density is

$$\boxed{\text{Density} = \frac{\text{Mass of substance}}{\text{Volume of substance}}} \tag{1}$$

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



Density and mixing

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

Table 3 Density of some common substances at 273.15 K and 100 kPa		
Substance	Density (g/mL)	Physical State
Helium	0.2	gas
Hydrogen	0.1	gas
Water	1.0	Liquid
Cooking oil	0.9	Liquid
Mercury	13.5	Liquid
Tetrachloroethene	1.6	Liquid
Gold	19.3	solid
Plastics	1.2	solid
Ice	0.916*	solid

*Ice is given at T $<$273.15 K

The immersion method

The density of liquids results from measuring the mass of a given volume of the liquid. Differently, density is harder to obtain for solids. For metals, we can calculate density by the immersion method: when a metal is immersed in water, the water rises. This increase in volume corresponds to the volume of the solid. This way, density results from the direct measurement of mass and the measurement of volume by displacement.

Specific gravity

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

$$\boxed{\rho = \frac{\text{density of substance}}{\text{density of water}}} \tag{2}$$

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

Sample Problem 8

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

SOLUTION

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

STUDY CHECK

A lead weight used in the belt of a scuba diver has a mass of 226 g. When the weight is placed in a graduated cylinder containing 200.0 mL of water, the water level rises to 220.0 mL. What is the density of the lead weight (g/mL)?

▶Answer:11.3 g/mL.

D. Procedure

1. Density of water The goal of this mini-experiment is to calculate the density of water. In order to do this you will measure the mass of a specific volume of water and use the formula for density:

$$d = \frac{m}{V}$$

Step 1: – Place approximately 25mL of tab water into a 100mL cylinder. Indicate the exact volume you employed in the table below.

Step 2: – Place a 100mL (or 25mL) beaker in the scale and press zero. After that add the liquid from the cylinder and write down the mass in the table below.

Step 3: – Compute the value of density. Research the expected value from the internet and make sure your value is reasonable.

2. Density of a solution In this section you will calculate the density of an unknown solution by repeating the procedure from the previous mini-experiment. You will also compute the specific gravity by diving the density of the solution by the density of water.

$$\text{specific gravity} = \frac{d}{d_{water}}$$

Step 1: – Place approximately 25mL of the solution into a 100mL cylinder. Indicate the exact volume you employed in the table below.

Step 2: – Place a 100mL (or 25mL) beaker in the scale and press zero. After that add the liquid from the cylinder and write down the mass in the table below.

Step 3: – Use a thermometer and measure the temperature of water; obtain the density of water at the measured temperature using the link below:

`https://antoine.frostburg.edu/chem/senese/javascript/water-density.html`

Step 4: – Compute the value of density.

3. Density of a solid In this mini-experiment you will calculate the density of a metal by a method called volume displacement. You will measure the volume of a liquid before and after adding the solid. The difference in volume is the volume of the solid. By means of this measurement and the mass of the solid you will be able to estimate density.

Step 1: – Obtain a metallic object and weight it. Record its mass in the table below.

Step 2: – Attach a string to the object and submerge it in a 50mL cylinder big enough to fit the object.

Step 3: – Add water until the object is covered. Record this volume in the table below. This is $V_{(After)}$.

Step 4: – Now using the string remove the object from the cylinder. Write down the liquid volume after the object is out. This is $V_{(Before)}$.

Step 5: – Calculate the volume of the object by subtracting the liquid volume before and after removing the object from the cylinder.

Step 6: – Use the formula of density to calculate the density of the solid.

5. Density by graphing The goal of this mini-experiment is to calculate the density of a metal by means of a graph. This method is useful for small pieces of metal which volume can not be computed by means of the volume displacement method. You will continuously add pieces of metal to a liquid so that the volume will progressively increase. By graphing mass vs. volume you will be able to compute density.

Step 1: – Place 25mL of water in a 100mL cylinder. Carefully record the liquid volume in the table below.

Step 2: – By means of a scale calculate the mass of the liquid and record the value reporting all decimals given by the scale.

Step 3: – Add metal pieces (or perhaps pennies) so that the liquid volume changes significantly. Write down the new volume and the new mass. Repeat this procedure until you fill up the Results table.

Step 4: – Keep on adding metal pieces until you fill in the table below.

Step 5: – Plot mass (vertical axis) vs. volume (horizontal axis) in the graph below.

Step 6: – You will calculate the density of the metal by selecting two arbitrary points from the plot (point 1 and point 2, where point 2 has a larger mass) and using the formula:

$$\text{density} = \frac{mass(2) - mass(1)}{volume(2) - volume(1)} = \underline{\hspace{5cm}} = \qquad g/mL$$

6. Reading a hygrometer The goal of this mini-experiment is to calculate the specific gravity of two liquids by reading a hygrometer.

Step 1: – You will find two hygrometers displayed in the lab, one with water and another with an unknown liquid.

Step 2: – Read the measurements of both and write down the result in the table below. Make sure your results have the right number of meadured digits (e.g. 1.0005).

STUDENT INFO

Name: ________________________ Date: ______

Pre-lab Questions

Density and Specific Gravity

1. Do oil float on water? Explain why.

2. Research the meaning of specific gravity.

3. A 3g glucose solution occupies a volume of 0.1L. Calculate the density of the solution in g/mL.

4. Oil has a density of 0.9g/mL. Calculate the mass in grams of a 100mL oil sample.

5. An electrolyte solution has a density of 1.3g/mL. Calculate the volume in L of a 2mg sample.

STUDENT INFO

Name: ____________________ Date: ______

Results

EXPERIMENT

Density and Specific Gravity

---

1. Density of water

Volume (mL)	mass (g)	Density (g/mL)

2. Density of a solution

Volume (mL)	mass (g)	Density of liquid (g/mL)	Density of water (g/mL)	Specific gravity

3. Density of a solid

Mass (g)	Volume before adding the object, $V_{(Before)}$ (mL)	Volume after adding he object, $V_{(After)}$ (mL)	$V_{(After)}$-$V_{(Before)}$ (mL)	Density (g/mL)

4. Density by graphing

Volume (mL)	mass (g)

Name of metal=____________________________

5. Reading a hygrometer

Specific gravity of water	Specific gravity of unknown

STUDENT INFO

Name: ______________________ Date: ______

Post-lab Questions

Density and Specific Gravity

---

1. A nugget of metal with a mass of 400 g is added to 25.0 mL of water. The water level rises to a volume of 40 mL. What is the density of the metal?

2. Determine the density (g/mL) of a 0.3 L sample of a salt solution that has a mass of 40 g.

3. A graduated cylinder contains 25 mL of water. What is the new water level after 35 g of silver metal is submerged in the water if the density of silver is 11g/mL?

EXPERIMENT 0

Energy and Matter

A. Goal

The goal of this laboratory experiment is to experimentally obtain heating curves of substances for boiling and freezing and to measure the heat of fusion of ice.

B. Materials

- ☐ 100mL, 250mL beaker (or a 400mL)
- ☐ hot plate
- ☐ thermometer
- ☐ clam
- ☐ stand
- ☐ 150mL cylinder
- ☐ boling chips
- ☐ jumbo testube with sasol
- ☐ double styrofoam cup
- ☐ crushed ice

C. Background

Temperature

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

$$\boxed{T_F = 1.8T_C + 32} \quad \boxed{T_F = 1.8T_K - 459.4} \quad \boxed{T_K = T_C + 273} \tag{3}$$

Sample Problem 9

Convert 25 °C to °F.

SOLUTION

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

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

2 **Step two:** use the formula $T_F = 1.8T_C + 32$ to convert from °C to °F.

3 **Step three:** solve for $T_F = 1.8 \times 25 + 32 = 77°F$.

STUDY CHECK

Convert 200°C to K.

▶Answer: 473K.

From energy to temperature

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

Heat capacity

The heat capacity c of a material is defined as:

$$c = \frac{\text{heat adsorbed}}{\text{temperature increase}} \tag{4}$$

This is a characteristic property of each material that indicates the energy required to rise its temperature and can be expressed in cal/°C or J/°C units. As this property depends on the amount of matter, oftentimes the heat capacity is expressed per mass as the specific heat capacity also known as *specific heat* (c_e) or mole unit as the *molar heat capacity* c_m. For example, the specific heat of water is 1cal/g°C that is the same as 4.184J/g°C. That means that we need to give 1 calorie to warm up one gram of water $1°C$. Similarly, the specific heat of aluminum, a metal, is 0.2cal/g°C or 0.89J/g°C; that means the energy needed to raise the temperature of an aluminum gram is 0.2 calories of 0.89 J. Mind the difference between these two values: we need to give 1 cal to increase the temperature of a gram of water in $1°C$, whereas we need to give 0.2 cal to increase the temperature of a gram of aluminum in $1°C$. Why are these two numbers so different? The answer is that water and aluminum are different materials. Normally metals warp up very easily, that is, they need less heat to increase their temperature, whereas liquids need more heat to increase their temperature. That is why pans and cooking pots tend to be metallic. Table 5 lists specific heats of common substances. Mind the specific heat if water is a well know value that you need to be familiar with:

$$\boxed{c_e^{\mathrm{H_2O}} = 4.184\mathrm{J/g°C} \qquad \mathit{or} \qquad c_e^{\mathrm{H_2O}} = 1\mathrm{cal/g°C}} \tag{5}$$

Heat

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

$$\boxed{Q = m \cdot c_e \cdot (T_f - T_i)} \tag{6}$$

where:

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

m is the mass of material in grams

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

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

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

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

Sample Problem 10

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

SOLUTION

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

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

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

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

STUDY CHECK

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

Answer: $Q = 231cal$.

Table 5 Values of specific heat for different materials

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

D. Procedure

1. Heating curve for water While heating a liquid its temperature raises up until the moment the liquid boils.

Step 1: – Place a 250mL beaker (or a 400mL) on top of a hot plate. Place a thermometer in the beaker so that it does not touch the walls of the beaker and secure it with a clam.

Step 2: – Use a cylinder and place 150mL cool of water into the beaker.

Step 3: – Using the thermometer record and write down the initial temperature of water.

Step 4: – Start heating the liquid at medium heat.

Step 5: – Record the temperature in the table below every minute (you might need to add extra space in the table to accommodate all numbers). Use a stopwatch to measure time.

Step 6: – When large bubbles continuously appear (not small bubbles), the liquid will be boiling. After that point record the temperature for only 10 minutes. In some cases, water may seem like that does not boil after 20 minutes. In those cases, consult with your instructor.

Step 7: – Turn off the hop plate when the experiment is done.

Step 8: – Using a pencil, plot the heating curve of water by graphing temperature (Vertical axis) vs. time (Horizontal axis). Make sure the time occupies the whole space in the plot. Show this plot to your instructor.

2. Cooling curve of salol Phenyl salicylate, or salol, is a chemical once used in sunscreens, phenyl salicylate and now used in the manufacture of some polymers, lacquers, adhesives, waxes, and polishes. This chemical is solid at room temperature. The goal of this mini experiment is to draw the cooling curve of melted salol.

Step 1: – Half-fill a 400mL beaker with water. Add boiling chips and start boiling the liquid with a hot plate. This is a water bath meant to melt salol.

Step 2: – Place the salol container in the water bath. Add a thermometer inside the salol tube to control its temperature. Melt the solid completely. Never warm up salol beyond 80°C.

Step 3: – When salol is all melted stop the hot plate and start recording temperature every minute. Write down the results in the table below.

Step 4: – After the solid forms, continue measuring temperature for five more minutes.

Step 5: – Stop recording when salol is fully solidified.

Step 6: – Write down the measurement in the table below.

Step 7: – Plot the heating curve of water by graphing temperature (Vertical axis) vs. time (Horizontal axis).

3. Heat of fusion of ice The goal of this mini experiment is to calculate an estimate of the heat of fusion of ice. You will do this by using a calorimeter (a double styrofoam cup) and a thermometer.

Step 1: – Weight a empty double styrofoam cup and record its mass.

Step 2: – Add 100mL of water to the cup and weight again. Record the new mass.

Step 3: – Record the initial temperature of water with a thermometer.

Step 4: – Add crushed ice to the cup with water. The amount of ice should fill half of a 100mL beaker.

Step 5: – Close the calorimeter until all the ice is melted. Record the final temperature.

Step 6: – Weight the cup with water and the melted ice and record the final mass.

Calculate the fusion heat of ice by using the following formula, in which $C_{e,water}$ is the specific heat of water (1cal/g/°C):

$$-m_{ice} \times Q_{fusion} + m_{water} \times C_{e,water} \times \Delta T = 0$$

STUDENT INFO

Name: ______________________ Date: ______

Pre-lab Questions

Energy and Matter

---

1. When ice melts is heat lost or gained? Explain.

2. Calculate the mass of 100mL of water. Density is 1g/mL.

3. What happens to the temperature of water while its boiling?

4. How many calories are needed to boil 100g of water? ($heat_{vaporization}$=540 cal/g)

5. How many calories are needed to melt 100g of ice? ($heat_{fusion}$=80 cal/g)

6. How many calories are needed to warm up 100g of water from 10 to 50°C? (C_e=1 cal/g/°C)

7. The following formula is used to calculate the heat of fusion of ice using a calorimeter, where $C_{e,water}$ is the specific heat of water (1cal/g/°C), the mass of ice is 5g, the mass of water in the calorimeter is 100g and the temperature decrease is -4°C

$$-m_{ice} \times Q_{fusion} + m_{water} \times C_{e,water} \times \Delta T = 0$$

Calculate the heat of fusion of ice.

STUDENT INFO

Name: ____________________ Date: ________

Results

EXPERIMENT

Energy and Matter

1. Heating curve for water

Time (min)	Temperature (°C)

2. Cooling curve of salol

Time (min)	Temperature (°C)

3. Heat of fusion of ice

1	Mass of the calorimeter (g)	________
2	Mass of the calorimeter+ water (g)	________
2 – 1	Mass of the water,m_{water} (g)	________
3	Initial temperature of water (°C)	________
4	Final temperature of water (when ice is melted) (°C)	________
4 – 3	Temperature change, ΔT (°C)	________
5	Mass of the calorimeter+ water + melted ice (g)	________
5 – 2	Ice mass, m_{ice} (g)	________

$Q_{fusion} =$________

STUDENT INFO

Name: ________________________ Date: ______

Post-lab Questions

Energy and Matter

---

1. Label the different areas of the heating and cooling curves you plotted with the labels ((s), (l), (g), (s+l), or (l+g)) representing solid, liquid or gas.

2. According to your plot, what is the boiling or freezing temperature of the liquid.

3. Explain the meaning of heat of fusion.

4. Explain why during a phase transition temperature is constant.

EXPERIMENT 0

Atoms and elements

A. Goal

The goal of this laboratory experiment is to practice unit conversions and carry out calculations with the correct number of significant figures.

B. Materials

☐ Display of different elements (Al, C, Cu, Fe, Mg, Ni, N, O, P, Si, S, Sn, Zn)

C. Background

The periodic table

The periodic table (see Figure 3) is a chart containing all known elements arranged in increasing number of electrons per atom in a way that elements with similar chemical and physical properties are located together. The periodic table contains all existing elements–some of them are synthetic others are natural–that form the matter arranged in columns and rows. Every element has a different name accompanied by a symbol that represents its name. The tabular arrangement of elements in the form of rows and columns allows further classification of the elements according to their properties. This section will cover the different features of the periodic table.

Elements and Symbols

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

Periods and groups

The periodic table (see Figure 3) contains all elements arranged in rows and columns. The horizontal rows are called *periods* and the vertical columns are called *groups or families*. For example, the first period contains hydrogen (H) and helium (He), whereas the second group contains Beryllium (Be), Magnesium (Mg), Calcium (Ca), Strontium (Sr), Barium (Ba) and Radium (Ra). There are seven periods (periods 1-7) and 18 groups. Some of the groups are labeled with an A (e.g. group 8A) whereas others are labeled with a B (e.g. group 8B). Group numbers can be found written with roman numbers and a letter (A or B) or with a more modern group numbering of 1-18 going across the periodic table. For example, group 2 (Mg-Ra) can also be called IIA, and group 13 (B-Ti) is also known as IIIA.

Properties in the periodic table

The physical and chemical properties of some elements of the table (see Figure 3) are similar, and these similarities led to the organization of the periodic table. Elements in the same group share properties and for example, oxygen and sulfur have similar properties: both are reactive elements. Differently, the properties across periods change going from metals to nonmetals. For example, the properties of Li and Ne are very different, and lithium is a reactive metal whereas neon is a nonreactive gas.

Metals, Nonmetals, and Metalloids

Overall, the elements of the periodic table (see Figure 3) can be classified as metals, nonmetals, and metalloids. Metals are those elements on the left of the table and nonmetals are the elements on the right of the table. The elements between metals and nonmetals are called metalloids and include only B, Si, Ge, As, Sb, Te, Po, and At. Metals are shiny solids and usually melt at higher temperatures. Some examples of metals are Gold (Au) or Iron (Fe). Nonmetals are often poor conductors of heat and electricity with low melting points. They also tend to be matt (non-shinny), malleable, or ductile. Some examples of nonmetals are Carbon (C) or Nitrogen (N). Metalloids are elements that share some properties with metals and others with nonmetals. For example, they are better conductors of heat and electricity than nonmetals, but not as good conductors as metals. Metalloids are semiconductors because they can act as both conductors and insulators under certain conditions. An example of metalloids is Silicon (Si) which should not be confused with silicone, a chemical employed in prosthetics.

Classification of elements in terms of groups

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

How to classify Hydrogen

At first sight, hydrogen (H) may seem to be put in the wrong spot on the periodic table (see Figure 3). Although it is located at the top of Group 1A (1), it is not an alkali metal, as it has very different properties. Thus hydrogen does not belong to the alkali metals, being nonmetal.

Sample Problem 11

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

SOLUTION

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

STUDY CHECK

Answer the following questions: (a) Give the group and period of the following elements, and give the name: Cl. (b) Classify as alkali metal, alkali earth metal, transition metal, halogen or noble gas, and give the name: Ne. (c) Classify as metal, nonmetal or metalloid, and give the name: W.

▶Answer: (a) Chlorine: G 17 (VIIA) P3; (b) Neon Noble gas ; (c) Tungsten metal.

The atom

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

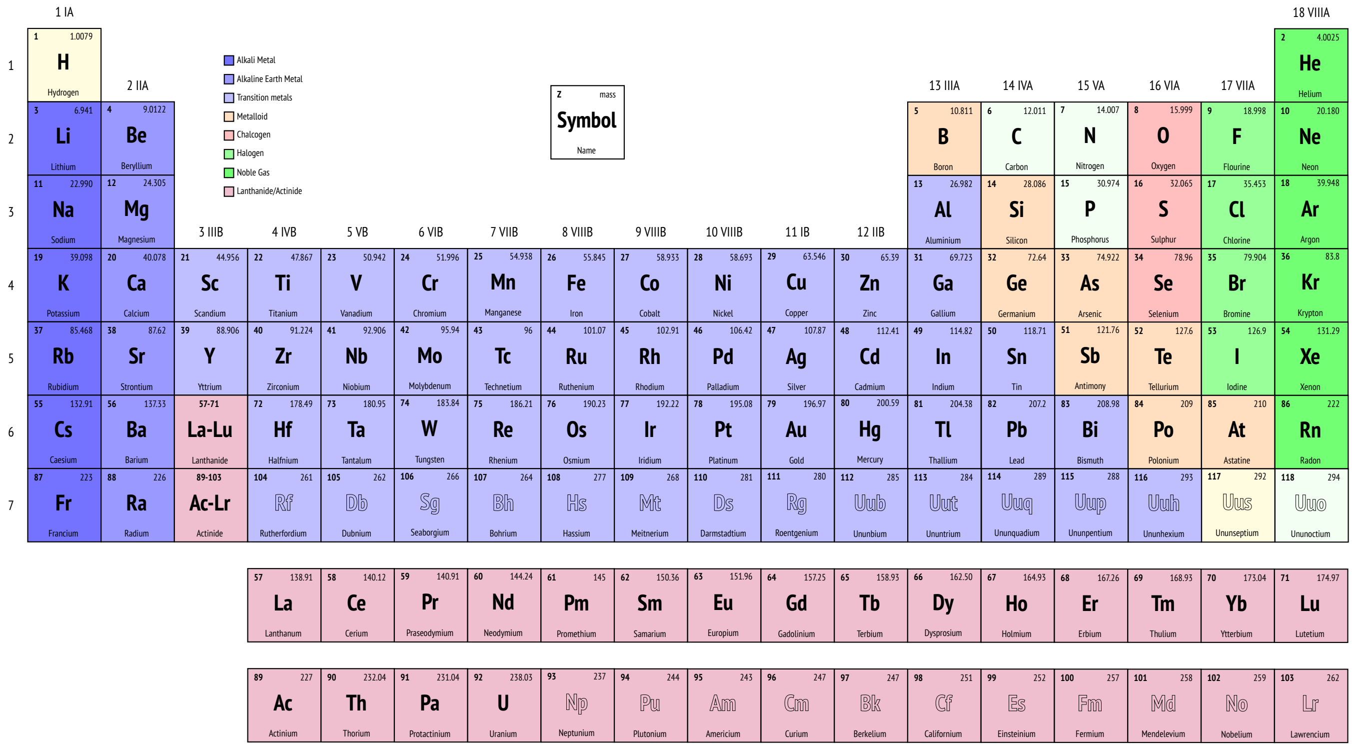


Figure 3 The periodic table of the elements

Atomic Structure

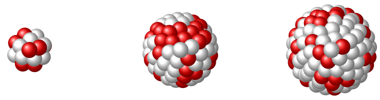
An atom is an electrically neutral, spherical entity made of a nucleus surrounded by negatively charged electrons. Atoms contain three atomic particles: the proton, neutron, and electron. Protons have a positive charge (+), whereas electrons carry a negative charge (–). Both electrons and protons have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

Atomic and mass number

Elements are made of atoms, and each atom of an element is characterized by an atomic number (Z) and a mass number (A). The atomic number (Z) of an element indicates the number of protons in an atom. This number can be easily located in the periodic table (see Figure 3). All atoms of an element have the same atomic number, whereas the atomic number of different elements differ. For example, Carbon has an atomic number of Z=6, whereas Oxygen has an atomic number of Z=8. The mass number (A) of an element indicates the combined number of protons and neutrons. Mass numbers can not be found in the periodic table. More importantly, different atoms of the same element can have different mass numbers. For example, a Carbon atom made of 6 neutrons and 6 protons has a mass number of A=12. Both A and Z for an atom X are indicated in the following form called isotope notation:

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

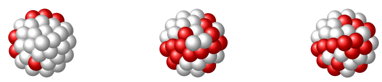
As an example, the notation $^{24}_{12}$Mg means that the atomic number of Mg is Z=12 and the mass number is A=24. Using the isotope notation, one can quickly identify the number of protons, neutrons, and electrons in an atom. As the atomic number is always indicated on the bottom part (e.g. Mg has 12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



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

Isotopes

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



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

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

Average atomic mass

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

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

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

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

Atomic masses can be simply found in any periodic table (see Figure 3) for each element. For example, the atomic mass of oxygen (O) is 15.999 amu and the atomic mass of nitrogen (N) is 14.007 amu. The atomic mass found in the periodic table is an average that results from including the mass of the different isotopes and their abundance. Table **??** lists the relative abundance of a series of common isotopes.

Sample Problem 12

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

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

SOLUTION

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

STUDY CHECK

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

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

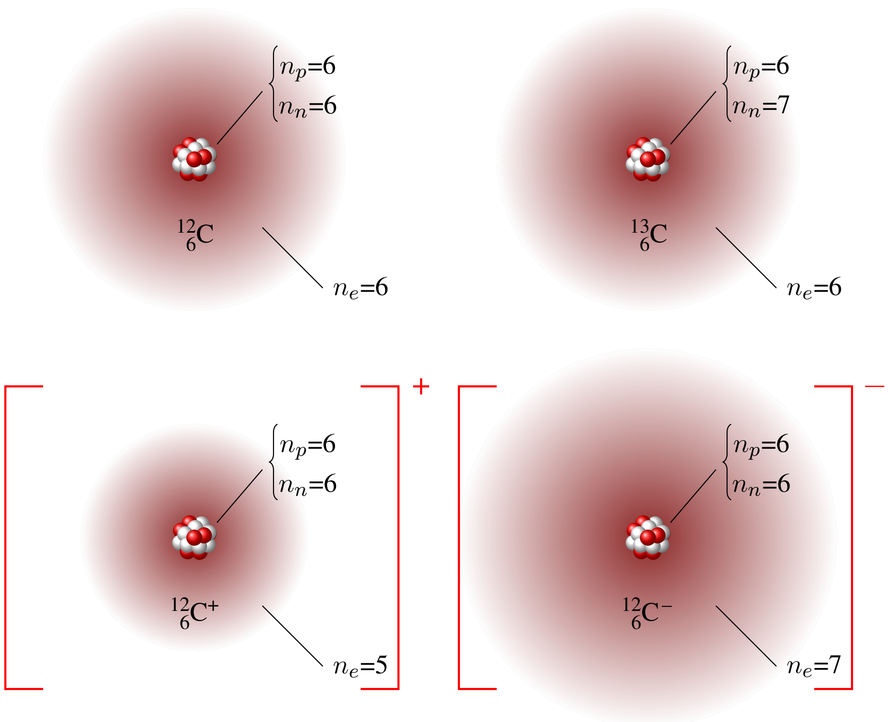


Figure 4 Representations of four different atoms, two neutral atoms on top and two ions on the bottom.

D. Procedure

Appearance of some chemical elements

Step 1: – Write the chemical symbol and describe the color of the elements listed below.

Step 2: – Describe the luster of the elements listed below (shiny/dull).

Step 3: – Based on your observations, describe the elements as metals, nonmetals or metalloids.

> **Good Lab Practice**
>
> ✍ Be gentle when handling the display of chemical elemens.

The atom and its composition

Step 1: – Fill the table below indicating the number of electrons, protons and neutrons of the following neutral isotopes.

Neutral isotopes

Step 1: – Fill the table below indicating the number of electrons, protons and neutrons of the following neutral isotopes.

Charged isotopes

Step 1: – Fill the table below indicating the number of electrons, protons and neutrons of the following charged isotopes.

Average atomic masses

Step 1: – For the element below calculate the average atomic mass by multiplying the mass of the different isotopes by its abundance and adding the contributions.

Atomic spectrum

Step 1: – Your instructor will show you the light spectra for a set of elements and compounds.

Step 2: – Describe the light color for each.

STUDENT INFO

Name: ________________________ Date: ______

Pre-lab Questions

Atoms and elements

---

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

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

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

4. 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

STUDENT INFO

Name: ______________________ Date: ______

Results
EXPERIMENT

Atoms and elements

Appearance of some chemical elements

Element	Symbol	Atomic number	Luster Shinny/dull	Metallic Character Metal/Nonmetal/ Metalloid
Aluminium	____________	____________	____________	____________
Carbon	____________	____________	____________	____________
Copper	____________	____________	____________	____________
Iron	____________	____________	____________	____________
Magnesium	____________	____________	____________	____________
Nickel	____________	____________	____________	____________
Nitrogen	____________	____________	____________	____________
Oxygen *	____________	____________	____________	____________
Phosphorus	____________	____________	____________	____________
Silicon	____________	____________	____________	____________
Silver*	____________	____________	____________	____________
Gold *	____________	____________	____________	____________
Sulfur	____________	____________	____________	____________
Tin	____________	____________	____________	____________
Zinc	____________	____________	____________	____________
Calcium	____________	____________	____________	____________

* Not given

The atom and its composition

Name	Symbol	Atomic number, Z	Mass number, A	Protons	Neutrons	Electrons
	Fe				30	
			134			55
					32	28
Fluorine			18			
	C		12			

Neutral isotopes

Isotope	Protons	Neutrons	Electrons
$^{27}_{12}Mg$			
$^{64}_{29}Cu$			
$^{79}_{34}Se$			
$^{103}_{46}Pd$			

Charged isotopes

Isotope	Protons	Neutrons	Electrons
$^{27}_{12}Mg^{2+}$			
$^{64}_{29}Cu^{+}$			
$^{18}_{8}O^{2-}$			
$^{15}_{7}N^{3-}$			

Average atomic masses

Isotope	Isotopic mass (m)	Abundance (%)	Fractional Abundance (f)	$m \times f$
$^{32}_{16}S$	31.97207	95.0		
$^{33}_{16}S$	32.97146	0.76		
$^{34}_{16}S$	33.96786	4.22		
Average mass (amu)				

Atomic spectrum

Nitrogen

Oxygen

Helium

Neon

Argon

STUDENT INFO

Name: ______________________ Date: ______

Post-lab Questions

Atoms and elements

---

1. 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.

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

EXPERIMENT 0

Specific heat and food

A. Goal

The goal of this laboratory experiment is to measure the specific heat of a metal and to review the concept of energy value.

B. Materials

- ☐ metallic object
- ☐ 250mL (or 400mL) beaker
- ☐ hot plate
- ☐ string
- ☐ double styrofoam cup (a calorimeter)
- ☐ thermometer
- ☐ food product labels

C. Background

From energy to temperature

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

Heat capacity

The heat capacity c of a material is defined as:

$$c = \frac{\text{heat adsorbed}}{\text{temperature increase}} \tag{7}$$

This is a characteristic property of each material that indicates the energy required to rise its temperature and can be expressed in $\text{cal}/°\text{C}$ or $\text{J}/°\text{C}$ units. As this property depends on the amount of matter, oftentimes the heat capacity is expressed per mass as the specific heat capacity also known as *specific heat* (c_e) or mole unit as the *molar heat capacity* c_m. For example, the specific heat of water is $1\text{cal}/\text{g}°\text{C}$ that is the same as $4.184\text{J}/\text{g}°\text{C}$. That means that we need to give 1 calorie to warm up one gram of water $1°C$. Similarly, the specific heat of aluminum, a metal, is $0.2\text{cal}/\text{g}°\text{C}$ or $0.89\text{J}/\text{g}°\text{C}$; that means the energy needed to raise the temperature of an aluminum gram is 0.2 calories of 0.89 J. Mind the difference between these two values: we need to give 1 cal to increase the temperature of a gram of water in $1°C$, whereas we need to give 0.2 cal to increase the temperature of a gram of aluminum in $1°C$. Why are these two numbers so different? The answer is that water and aluminum are different materials. Normally metals warp up very easily, that is, they need less heat to increase their temperature, whereas liquids need more heat to increase their temperature. That is why pans and cooking pots tend to be metallic. Table 5 lists specific heats of common substances. Mind the specific heat if water is a well know value that you need to be familiar with:

$$\boxed{c_e^{\text{H}_2\text{O}} = 4.184\text{J}/\text{g}°\text{C} \qquad or \qquad c_e^{\text{H}_2\text{O}} = 1\text{cal}/\text{g}°\text{C}} \tag{8}$$

Heat

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

$$\boxed{Q = m \cdot c_e \cdot (T_f - T_i)} \tag{9}$$

where:

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

m is the mass of material in grams

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

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

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

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

Sample Problem 13

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

SOLUTION

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

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

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

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

STUDY CHECK

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

Answer: $Q = 231cal$.

Table 5 Values of specific heat for different materials

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

Calories in food

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

$$\boxed{1Cal = 1000cal} \quad \text{or} \quad \boxed{\frac{1Cal}{1000cal}} \quad \text{or} \quad \boxed{\frac{1000cal}{1Cal}} \tag{10}$$

Energy values

Do you ever eat pasta? Think about how does your body feel after you eat pasta? Normally, whenever you eat pasta in a few hours you need to eat again more food. Differently, whenever you eat meat, that is enough to keep you going for a longer time. Similarly, eating a salad for lunch brings you less energy than a pizza slice. This is because each type of food–each ingredient–contains different energy. We refer to this as the energy value of food ϵ. Table 6 lists energy values for common ingredients. To compute the energy (E) provided by a certain mass of food (m) we need to multiply the mass times the energy value (ϵ):

$$E = m \cdot \epsilon \tag{11}$$

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

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

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

Sample Problem 14

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

SOLUTION

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

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

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

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

$$E_{fat} = m_{fat} \cdot \epsilon_{fat} = 28\cancel{g} \times 9\frac{kcal}{\cancel{g}} = 252kcal \qquad \text{energy from fat}$$
$$E_{carb} = m_{carb} \cdot \epsilon_{carb} = 46\cancel{g} \times 4\frac{kcal}{\cancel{g}} = 184kcal \qquad \text{energy from carbs}$$
$$E_{prot} = m_{prot} \cdot \epsilon_{prot} = 25\cancel{g} \times 4\frac{kcal}{\cancel{g}} = 100kcal \qquad \text{energy from protein}$$

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

STUDY CHECK

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

Answer: $299kcal$.

Table 6 Energy value of food

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

D. Procedure

1. Specific heat of a metal The goal of this mini experiment is to calculate the specific heat of an unknown metal. You will do this by warming up the metal in a hot water bath and by using a calorimeter to cool down the metal.

Step 1: – Obtain a metallic object. Record its mass.

Step 2: – Place a 250mL beaker (or a 400mL) on top of a hot plate. Place a thermometer in the beaker so that it does not touch the walls of the beaker and secure it with a clam. Start warming up the water at high heat so that the water boils. Add some boiling chips.

Step 3: – Tie a string to the object and submerge it in the how bath. Let it there for 10 min.

Step 4: – Obtain a double styrofoam cup and weight it. Record its mass.

Step 5: – Add 50mL of water to the cup and weight again. Record the new mass. Make sure the water is enough to fully cover the metal. If not add some more.

Step 6: – Measure the temperature of the hot bath after the metals it's been there for 10 min. Record the value.

Step 7: – Using the string and being careful not to drop the object, transfer the metal object from the hot bath to the cup with water. Cover the cup quickly and stir.

Step 8: – Using the thermometer in the calorimeter, measure the highest temperature reached by the water in the cup after you drop the object.

Step 9: – You might have to replicate the experiment.

2. Food value in food The goal of this mini experiment is to calculate the number of Calories, or Kcal, in different food products using the energy values of different food ingredients–carbs, fats and protein.

Step 1: – Obtain a food product labels.

Step 2: – Write down the name of the product in the table below.

Step 3: – Write down the mass of a serving in the table below.

Step 4: – List the grams of carbohydrates, fats and protein in your product.

Step 5: – calculate the number of Calories (kcal) for each food type in a serving. The accepted energy values of carbohydrates, fats and proteins are 4, 9 and 4 Cal/g.

Step 6: – Calculate the total number of calories in a serving and compare the value with the one in the food label.

STUDENT INFO

Name: ________________________ Date: ______

Pre-lab Questions

Specific heat and food

---

1. What is the formula to calculate the energy needed to warm up a metal? Explain the meaning of each variable.

2. The specific heat of Al is 0.215 cal/g/°C whereas the one for brass is 0.09 cal/g/°C. Explain the implications.

3. How many calories are absorbed by a 45.2g piece of aluminum (C_e =0.215 cal/g/°C) if its temperature rises from 10°C to 40°C.

4. What is the difference between cal and Cal?

5. A pepperoni pizza slice contains 10g of fat (9kcal/g), 36g of carbs (4kcal/g) and 14 g of protein (4kcal/g), where the caloric values are indicated in parenthesis. What is the total caloric intake of the slice?

STUDENT INFO	
Name: ______________________	Date: ______

Results

EXPERIMENT

Specific heat and food

1. Specific heat of a metal

1	Mass of the metal, m_{metal} (g)	__________
2	Mass of the calorimeter (g)	__________
3	Mass of the calorimeter + water (g)	__________
3 – 2	Mass of the water, m_{water} (g)	__________
4	Temperature of boiling water (°C)	__________
5	Initial temperature of water in calorimeter (°C)	__________
6	Final temperature of water in calorimeter (°C)	__________
6 – 5	Temperature change , ΔT (°C)	______________

Calculate the specify heat of the metal by means of the following formula in which $C_{e,water}$ is the specific heat of water (1cal/g/°C):

$$m_{metal} C_{e,metal} \times \Delta T + m_{water} \times C_{e,water} \times \Delta T = 0$$

$C_{e,metal} =$______________

2. Food value in food

Name of food product	________
Mass of a serving (g)	________
Mass of carbohydrate(g)	________
Mass of fats(g)	________
Mass of protein(g)	________
Calories from carbohydrates (Cal, or Kcal)	________
Calories from fat (Cal, or Kcal)	________
Calories from protein (Cal, or Kcal)	________

STUDENT INFO

Name: ______________________ Date: ______

Post-lab Questions

Specific heat and food

---

1. What percentage of the total calories in your food product is coming from fats?

2. What percentage of the total calories in your food product is coming from carbs?

EXPERIMENT 0

Nuclear Chemistry

A. Goal

The goal of this laboratory experiment is to measure the radioactivity of a set of chips and to test the effect of the use of a shield and distance.

B. Materials

- ☐ ST360 radiation counter
- ☐ set of chips
- ☐ set of shields

C. Background

Radiation, particles & radioisotopes

Light elements have normally stable nuclei. Differently, heavier elements with atomic numbers larger than 20 tend to often have several isotopes–remember these are atoms of the element with a different number of neutrons–that have unstable nuclei. For these unstable isotopes, the forces that keep the nucleus together are not strong enough to stabilize the nuclei. An unstable nucleus is radioactive, which means that it will spontaneously emit radiation in the form of small particles. Not all radioactivity is the same and there exist different types of radiation, which we will address in the following. Table 7 reports common nuclear symbols.

alpha radiation

Alpha radiation–referred to as α–is a type of radiation that contains alpha particles. These particles are indeed helium nuclei, with 2 protons, 2 neutrons, and a (2+) positive charge. Alpha particles are often represented as α or $^4_2\mathrm{He}$.

beta radiation

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

gamma radiation

Gamma radiation–referred to as γ–is a type of radiation that contains high-energy photons. These particles are indeed photons with no mass or charge. Gamma particles are often represented as γ or $^0_0\gamma$.

protons

Protons in this chapter are often referred to as p or $^1_1\mathrm{H}^+$. These are positive charges.

positrons

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

neutrons

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

Radioisotope notation

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

Sample Problem 15

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

SOLUTION

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

STUDY CHECK

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

Table 7 Nuclear symbols

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

Nuclear reactions

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

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

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

alpha decay

Some isotopes produce alpha radiation, that is, they produce α particles on its decay. A nuclear reaction that produces an α particle (${}^{4}_{2}\mathrm{He}$) is called alpha decay. In alpha decay, the emitter decreases its mass number A four units and its atomic number Z two units.

$$\text{Emitter} \longrightarrow {}^{4}_{2}\mathrm{He} + \text{New isotope}$$

beta decay

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

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

positron emission

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

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

gamma decay

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

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

Sample Problem 16

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

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

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

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

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

SOLUTION

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

STUDY CHECK

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

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

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

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

Radiation protection

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

alpha particles

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

beta particles

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

gamma particles

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

D. Procedure

0. Instructions for the use of ST360 radiation counter

Step 1: – Turn on the radiation counter (red button on the back).

Step 2: – Use display select to select *time*. Set the time to 60 seconds by pressing the display Up or Down.

Step 3: – Use display select until the light cursor is next to *high voltage*. Set the voltage to 900V by pressing the display Up or Down.

Step 4: – Use display select until the light cursor is next to counts.

Step 5: – Place a radioactive chip on the chip support and measure. In case of using a shield, place it between the counter and the chip.

Step 6: – Press Count and wait until the stop button sign lights up, as the machine stops automatically when the 60s is up.

Step 7: – Right down the unit of the reading as count per minute (cpm).

1. Background radiation The air has a certain radioactivity called background radioactivity. This is a very small activity but still affect the radioactive measurements and hence it should be taken into account. In this experiment you will measure the background radiation by means of a Geiger counter. You will have to repeat the measurement several times and average the radiation measured in order to obtain a reliable number.

Step 1: – Do not use any of the chips and make sure they are in the secured protecting box.

Step 2: – Start the Geiger counter. Set up the measurement time to 60 seconds and the measuring voltage according to your professor's instructions. Mind to select a voltage of 900V for all measurements (Press Display/High Voltage/Up/Down until you reach 900V). Press measure (press Display until the light cursor is next to count; then press Count until the stop button lights up.) and write down the background radioactivity in counts per minute in the table below.

Step 3: – Repeat the measurement two more times and calculate the average by adding the three measurements and dividing by three. Make sure the measurements are consistent with each other.

2. Radioactive chips In this section you will calculate the radioactivity of a set of different radioactive chops. You will still use the Geiger counter and after measuring the number of counts per minutes you will have to subtract the background radiation to your measurement.

Step 1: – Select three of the radioactive chips.

Step 2: – Place one of the chips 5cm away from the counter by means of the plastic stand.

Step 3: – Start the Geiger counter. Set up the measurement time to 60 seconds and the measuring voltage according to your professor's instructions. Press measure and write down the background radioactivity in counts per minute in the table below.

Step 4: – Repeat the measurement for the other two chips.

Step 5: – Repeat the measurement for other materials such as tea, instant coffee, potassium chloride or dry seaweed.

Step 6: – Now subtract the background radiation measured in the previous section to teach of the measurements.

3. Radioactivity protection In this section you will only use one of the chips from the previous experiment. For this one chip you will use different barriers to shield radiation and estimate the shielding impact.

Step 1: – Select radioactive chip from the previous experiment that gave you the highest counts per minute.

Step 2: – Place the chips at a 10cm distance from the counter by means of the plastic stand.

Step 3: – Select three of the shielding and place one of these in between the counter and the chip.

Step 4: – Measure the number of counts per minute and subtract the background radiation.

Step 5: – Write down the measurement in the table below. Compute the activity taking into account the background radiation. These results can potentially be a negative value.

Step 6: – Repeat the procedure for the other two shielding.

4. Effect of distance on radioactivity In this section you will only use one of the chips from the previous experiment, again the stronger one. You will place this chip at different distances from the counter and measure the impact of distance on radioactivity.

Step 1: – Place the chips at five different distances from the counter by means of the plastic stand. The distances are indicated in the table below.

Step 2: – Measure the number of counts per minute for each distance and subtract the background radiation.

Step 3: – Write down the measurement in the table below.

Step 4: – Plot activity without the background (right column, vertical axis) vs distance (horizontal axis) connecting the points with a line in the graph below.

STUDENT INFO	
Name: ______________________	Date: ______

Pre-lab Questions

Nuclear Chemistry

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1. Classify the following decays as α, β, γ or positron emission:

   (a) $${}^{235}_{92}\mathrm{U} \longrightarrow {}^{231}_{90}\mathrm{Th} + {}^{4}_{2}\mathrm{He}$$

   (b) $${}^{235}_{92}\mathrm{U} \longrightarrow {}^{235}_{92}\mathrm{Np} + {}^{0}_{0}\gamma$$

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

   (d) $${}^{222}_{88}\mathrm{Ra} \longrightarrow {}^{4}_{2}\mathrm{He} + {}^{226}_{88}\mathrm{Ra}$$

2. Iodine-131 is used to treat certain types of thyroid cancer and some rarer types of cancer. Given that its half-life is 8days, calculate the amount of iodine of a 4g sample that remains after 10 days.

3. Strontium-89 is used to treat some types of secondary bone cancer. Given that its half-life is 50days, calculate the time needed to reduce a 4g sample into 2g.

4. Iron-59 is used in studies of iron metabolism in the spleen. A Iron-59 sample has an activity of 20cpm and 10cpm 46 days after the first measurement. Calculate the half life of the isotope.

STUDENT INFO	
Name: ______________________	Date: ______

Results
EXPERIMENT

Nuclear Chemistry

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1. Background radiation

Measurement 1 (cpm)	Measurement 2 (cpm)	Measurement 3 (cpm)	Average Radiation (cpm)

2. Radioactive chips

Isotope name	Activity (cpm)	Activity - Background (cpm)

3. Radioactivity protection

Isotope name=____________________________

Shielding name	Activity (cpm)	Activity - Background (cpm)

5. Effect of distance on radioactivity

Isotope name=____________________________

Distance (cm)	Activity (cpm)	Activity - Background (cpm)
1cm		
2cm		
3cm		
4cm		
5cm		

Isotope Name=____________________________

STUDENT INFO	
Name: ____________________	Date: ______

Post-lab Questions

Nuclear Chemistry

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1. From the radioactive chips you studied indicate the nature of the radiation (written on the chip) produced by the strongest chip.

2. From the common materials, you tested for radiation (tea, seaweed, coffee, and KCl), which one gave you the highest radioactive measurement?

3. From the different shielding you studied indicate the nature of the one that protected the most from radiation.

4. From the graph you made estimate the number of counts per minute at 3.5cm from the counter.

EXPERIMENT 0

Compounds and their structure

A. Goal

This experiment will go over the ideas of molecular geometry and bond-hybridization. On one hand, you will learn how to predict the geometry of a molecule and how to differentiate, for example, a linear molecule from a bent molecule. Also you will learn to predict the hybridization of atomic orbitals involved in a chemical bond.

B. Materials

☐ This is a theory-based experiment.

C. Background

Electron-dot structures of atoms & molecules

Protons, neutrons, and electrons make the atoms. Electrons–in particular valence electrons–are responsible for the main chemical properties of an atom. These electrons are loosely bound and can be exchanged easily with other atoms, in contrast to the strongly-tied core electrons. The electron-dot structure of an atom or a molecule–also called Lewis structures–is a visual representation of the electronic arrangement of the valence electrons. Atoms in a molecule will tend to be surrounded by eight electrons so that their electron configuration resembles a noble gas. This arrangement is known as the octet rule. This rule is responsible for the common negative charge of F, and the positive charge of Na: F ($[He]2s^22p^5$) can easily receive an extra electron producing ionic F^- ($[He]2s^22p^6$=[Ne]), and atomic Na ($[Ne]3s^1$) can lose an electron producing ionic Na^+ ($[He]2s^22p^6 = [Ne]$). There are a few exceptions. A remarkable one is the case of the hydrogen atom that follows the duet rule.

Valence electrons of atoms, and molecules

The electrons of an atom are divided into core electrons and valence electrons. The valence electrons of an atom are involved in chemical bonds as they are less bonded to the nucleus. *The number of valence electrons of an atom* corresponds to the group number. For example, hydrogen H belongs to the group IA, and hence it has one valence electron. Similarly, oxygen O belongs to the group VIA, having six valence electrons. Similarly, we can count the *number of valence electrons of a molecule* by adding the valence electrons of the atoms that make the molecule. For example, water (H_2O) has eight valence electrons as each oxygen has one valence electron and oxygen has six. The number of *pairs of electrons* is just the overall number of valence electrons divided by two. For example, water has eight valence electrons that correspond to 4 pairs of electrons.

Sample Problem 17

Indicate the number of valence electrons for the following atoms: N, O, C and S, and the number of pairs of electrons of the following molecules: NH_3, and CO_2.

SOLUTION

Nitrogen is in group VA and hence it has five valence electrons ($5e^-$). Oxygen belongs to the group VIA and C belong to IVA, hence they have wiz and four valence electrons, respectively. For the molecules, we have that ammonia has 8 electrons (nitrogen has five valence electrons and each hydrogen has one electron) that correspond to four pairs, whereas carbon dioxide has 16 electrons (carbon has four electrons and each oxygen has six) and eight pairs.

STUDY CHECK

Indicate the number of valence electrons for the following atoms: Cl and B.

▶Answer: Cl ($7e^-$), B ($3e^-$).

The octet rule

Atoms exchange electrons when they combine to form molecules. This electron exchange is the driving force that drives the formation of molecules from single atoms. The octet rule states that each atom in a stable molecule should be surrounded by eight (octet) electrons achieving noble gas electron configurations. There are two important exceptions to this rule as H is surrounded only by two electrons (this is called the duet rule), and B by six. This rule comes from the experimental observation of numerous molecules.

Electron-dot structure of an atom

The electron-dot structure of an atom is a visual representation of the arrangement of the valence electrons of the atom. To write the electron-dot structure of an atom, you just need to write down the symbol of the atom surrounded by the valence electrons located in the four directions of the space: top, bottom, right, and left. To place the electrons, you start in any of the directions and fill one electron at a time. For example, for the case of three electrons, we would have: $\cdot\dot{\mathrm{B}}\cdot$. After all four directions have been filled, you need to start pairing the electrons. For example, for the case of five electrons, we would have: $\cdot\underset{\cdot}{\dot{\mathrm{P}}}:$. Another example, oxygen has six valence electrons and hence, the electron-dot structure would be $\cdot\underset{\cdot}{\ddot{\mathrm{O}}}:$ Similarly, the electron-dot structure of fluorine would be $:\underset{\cdot}{\ddot{\mathrm{F}}}:$. For ions, you need to add (if its an anion) or subtract (if its a cation) valence electrons, and for example the electron-dot structure of the oxide anion O^{2-} is $:\underset{\cdot\cdot}{\ddot{\mathrm{O}}}:^{2-}$. The electron-dot structure of atoms is useful to predict–or make sense–of the atomic valence. Mind that the number of valence electrons of an atom is not the same as the valence of the atom. The valence of an atom is a number used to combine with other atoms forming compounds. For example, the electron-dot structure of nitrogen is $\cdot\underset{\cdot}{\dot{\mathrm{N}}}:$ and this atom needs to gain three electrons to reach the noble gas configuration with eight electrons $:\underset{\cdot\cdot}{\ddot{\mathrm{N}}}:^{3-}$ hence its valence is -3.

Sample Problem 18

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: $\cdot\underset{\cdot}{\ddot{\mathrm{N}}}\cdot$ and $\cdot\underset{\cdot}{\dot{\mathrm{C}}}\cdot$. Cl^- has eight valence electrons, that is seven plus one, and hence its electron-dot structure will be $:\underset{\cdot\cdot}{\ddot{\mathrm{Cl}}}:^-$.

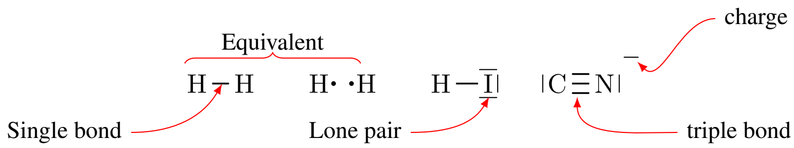
STUDY CHECK

Write down the electron-dot structure for N^{3-}

▶Answer: $:\underset{\cdot\cdot}{\ddot{\mathrm{N}}}:^{3-}$.

An introduction to electron-dot structures

Below, you will find some examples of electron-dot structures. Mind that the lines represent pairs of electrons hence below there are two equivalent representations for the hydrogen molecule. In these structures, you will find two different types of lines. Some pairs of electrons connect atoms. We call these types of pairs bonds. Other pairs lay on atoms. We call these lone pairs. Each atom can have a different number of lone pairs. For example, in the Lewis structures below carbon has one lone pair whereas iodide has three pairs. Bonds can be simple or multiple, double or triple. Finally, some molecules are charged and the charge is normally indicated on the top right side of the representation.



Electron-dot structure of diatomic molecules

Electron-dot structures–or Lewis structures–of diatomic molecules are the most simple electron-dot structures of molecules that you will see. To obtain these structures, you need to follow the next steps. The first step is (1) to set up the atoms in the molecule in the form of a line. After that, (2) you need to count the total number of valence electrons in the molecule by adding the valence electrons of each atom (remember the number of valence electrons corresponds to the group number in the A notation). Then (3) compute the pairs of electrons represented by lines–the total number of valence electrons divided by two. Finally, (4) you need to start distributing the electron pairs in the molecule in a very specific way, first connecting the atoms among themselves, and then placing the remaining pairs surrounding the atoms. Following the octet rule, each atom except for H and B should be surrounded by four pairs, counting as pairs the bonds and lone pairs.

Sample Problem 19

Construct the electron-dot structure of HCl.

SOLUTION

We first arrange the atoms in the molecule as indicated below and then we count the number of valence electrons: H(1) and Cl(7) that gives a total of eight electrons. We have four pairs of electrons.

H Cl

Now we distribute the pair on each atoms knowing that each atom has to have four pairs with the exception of hydrogen that can only be surrounded by one pair. We can use lines instead of pairs

H:C̈l: or H–Cl|

STUDY CHECK

Construct the electron-dot structure of HF.

Answer: H–F|.

Number of bonds and atomic nature

The number of covalent bonds that a nonmetal form is related to the number of electrons needed to complete the octet. For example, Hydrogen ($1s^2$) tends to form one bond to make compounds, whereas Nitrogen ($[He]2s^22p^3$) forms three bonds. For example, in the NH_3 molecule, each H forms one bond whereas N forms three bonds. Table **??** gives the relation between the number of bonds formed by several elements.

Electron-dot structure of general molecules

Now we will address how to build up electron-dot structures of more complex molecules given that one of the atoms is the central atom and the others are connected to this central atom. The first step is (1) to arrange the atoms in the molecule, in the form of a central atom and the remaining atoms around it; the central atom is the one with a lower index in the molecule (e.g. in H_2O is O or in NH_3 is N). After that, (2) you need to count the total number of valence electrons in the molecule, dividing this number by two to obtain the number of pairs of electrons represented by lines. In the following (3) you need to connect the surrounding atoms to the central atom with electron pairs, and then (4) place electron pairs on top of the surrounding atoms, always placing a maximum of four atoms. Finally (5) place the remaining pairs in the central atom. Overall, each atom should be surrounded by four pairs (this is the octet rule) except H and B which should be surrounded by one and three pairs respectively. When drawing Lewis structures it is not important the atom arrangement (if the molecule looks like a line, a triangle or so) as long as the connectivity (which atom goes in the center and the surroundings) is correct.

Sample Problem 20

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. $\mathrm{H{:}\ddot{\underset{\cdot\cdot}{O}}{:}H{:}}$ 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: $\begin{array}{c}\mathrm{H-\overline{N}-H}\\ |\\ \mathrm{H}\end{array}$; three bonds and one lone pair.

Multiple bonds

Often you are going to encounter electron-dot structures like the ones below

$$\mathrm{{:}N\equiv N{:}} \text{ and } \mathrm{\ddot{O}=\ddot{O}}$$

in which the atoms are connected through multiple bonds, double or triple bonds. Multiple bonds are formed while constructing electron-dot structures to impose the octet rule. Look for example the lewis structure for the HCN molecule below

$$\mathrm{H-C-\ddot{\underset{\cdot\cdot}{N}}{:}}$$

In this structure, carbon does not follow the octet rule. We can enforce the octet rule by moving lone pairs from the atoms into the bond forming the structure below

$$\mathrm{H-C\equiv N{:}}$$

In this structure both carbon and nitrogen follow the octet rule. Hence, we need to add one more step to the Lewis structure construction scheme: convert lone pairs of electrons into bonds to enforce the octet rule.

Sample Problem 21

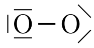
Construct the electron-dot structure of O_2.

SOLUTION

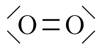
1. **Step one:** We first arrange the atoms in the molecule as

$$\mathrm{O\ O}$$

Now we count the total number of valence electrons, including all atoms: 2xO(6) that gives a total of twelve electrons. Let us count the pairs of electrons; we have twelve electrons and that is six pairs. Then we distribute the pairs, fist connecting the atoms O − O (we have five extra pairs to distribute at this point), and we place the remaining pairs on top of the oxygen atoms

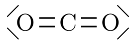


The right oxygen do not follow the octet rule. In order to enforce the octet rule we move lone pairs into the bond



STUDY CHECK

Construct the electron-dot structure of CO_2.

▶Answer: 

Atomic charges in a molecule

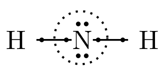
The electron-dot structure of a molecule results from counting the overall number of valence electrons of the molecule given that each atom brings a different number of valence electrons (n_e^{free}). For example, two H atoms bring one electron each, whereas O brings two electrons, giving a total of six valence electrons. When arranging the electron pairs in the molecule, the number of electrons surrounding an atom is called the valence of this atom in the molecule (n_e^{bonded}). We calculate the number of valence electrons of an atom in a molecule by accounting for the number of lone pairs on this atom and half the number of bonds:

$$n_e^{bonded} = \text{number of lone pairs} + 1/2\text{number of bonds}$$

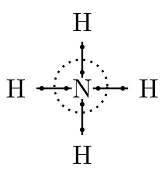
It is important to note that the valence of a free atom and the valence of this atom in a molecule is not necessarily the same. Indeed the difference between the valence electron of a free atom and the same atom in a molecule is the effective charge of that atom in the molecule, Q:

$$Q_{eff} = n_e^{free} - n_e^{bonded}$$

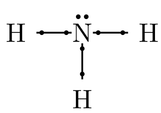
When the valence of an atom in a molecule is larger than the valence of the free atoms we have negative effective charges. In the example below



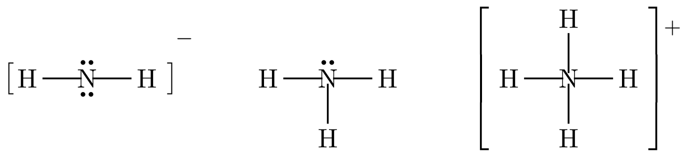
the number of electrons surrounding nitrogen is six electrons, more than the number of electrons originally brought to the molecule (five). We conclude that the atoms have a negative charge, and the effective atomic charge of nitrogen is Q=-1. In the next example,



the central atom, nitrogen, still has five valence electrons. After counting the electrons surrounding nitrogen, this time we find that this atom is surrounded by four electrons, less than the number of electrons originally brought to the molecule. We can conclude that nitrogen has a positive charge. In particular, the effective atomic charge of nitrogen in this molecule is the number of valence electrons minus the number of surrounding electrons. In this case, the atomic charge is Q=+1. When the valence of an atom in a molecule is the same as the valence of the free atoms we have zero effective charges. In this last example



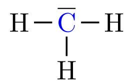
the central atom, nitrogen, has five valence electrons. After counting the electrons surrounding nitrogen–remember in a bond each atom shared an electron and hence each line around an atom counts as one electron–we find that this atom is surrounded by five electrons. As the number of valence electrons brought to the molecule is the same as the number of electrons surrounding the atom, we say the atomic charge of this atom is zero (Q=0). Hence, the atom is neutral. In all the molecules above, hydrogen remains neutral and hence the atomic charge of nitrogen corresponds to the molecular charge of the molecule. We can hence summarize the three scenarios indicated, as we have a neutral molecule in the center, a positive molecule on the right, and a negative molecule on the left.



The atomic charge of an isolated atom can be well-defined. However, the atomic charge of an atom in a molecule is arbitrarily defined, and more than one definition exists. Formal charges are one of the possible definitions of atomic charges in a molecule, whereas redox numbers are an alternative definition of atomic charges in molecules. None of these definitions is exactly correct. For example, redox numbers tend to overestimate the atomic charges, as they assume that all shared electrons in a bond belong to the most electronegative atom. Normally, negative formal charges tend to reside on electronegative atoms and not on electropositive atoms. At the same time, the sum of all effective charges needs to give the overall charge of the species. Furthermore, atoms in molecules tend to achieve formal charges as close to zero as possible. One can use formal charges to assess the validity of a Lewis structure. When comparing a series of equivalent Lewis structures for a molecule, the structures that best describe the bonding in the molecule tend to be those with small effective charges located on electronegative atoms.

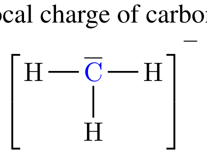
Sample Problem 22

Indicate the atomic charges of the blue highlighted atom



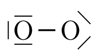
SOLUTION

The carbon atoms brings four electrons and in the molecule it is surrounded by eight electrons, five of which belongs to it. Hence the charge of C is -1; this means that carbon has one extra electron. Each hydrogen brings one electron and in the molecule each hydrogen has one electron (they share two electrons with C, one for C and one for H). The final lewis structure with the local charge of carbon can be indicated as:



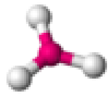
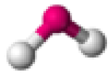
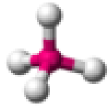
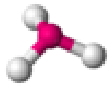
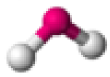
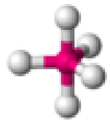
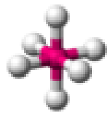
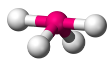
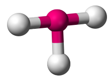
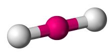
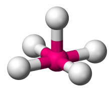
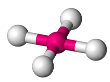
STUDY CHECK

Indicate the atomic charges of all atoms in the Lewis structure below



▶Answer: left oxygen -1; right oxygen $+1$

Table 5 Molecular geometries

Electron groups (AEs)	Electron-group geometry	Bonded atoms (Bs)	Lone pairs (Es)	ABE Code	Molecular shape	Bond Angle	3D model
2	Linear	2	0	AB_2	Linear	180°	
3	Trigonal Planar	3	0	AB_3	Trigonal Planar	120°	
3	Trigonal Planar	2	1	AB_2E	Bent	120°	
4	Tetrahedral	4	0	AB_4	Tetrahedral	109°	
4	Tetrahedral	3	1	AB_3E	Trigonal pyramidal	109°	
4	Tetrahedral	2	2	AB_2E_2	Bent	109°	
5	trigonal bipyramidal	5	0	AB_5	trigonal bipyramidal	90°, 120°,180°	
6	octahedral	6	0	AB_6	octahedral	90°, 180°,180°	
5	trigonal bipyramidal	4	1	AB_4E	see-saw	180°,120°, 90°	
5	trigonal bipyramidal	3	2	AB_3E_2	T-shaped	90°, 180°	
5	trigonal bipyramidal	2	3	AB_2E_3	Linear	180°	
6	octahedral	5	1	AB_5E	square pyramidal	90°	
6	octahedral	4	2	AB_4E_2	square planar	90°, 180°	

Steps to obtain Lewis structures

The following steps can be used to obtain the Lewis structure of a general molecule:

1. **Step one:** Arrange the atoms in the molecule, in the form of a central atom and the surrounding atoms
2. **Step two:** Obtain the number of pairs of valence electrons by dividing the total number of valence electrons of the molecule by two
3. **Step three:** Connect the surrounding atoms to the central atom with electron pairs
4. **Step four:** Place electron pairs on top of the surrounding atoms, always placing a maximum of four atoms
5. **Step five:** Place the remaining pairs in the central atom.
6. **Step six:** Convert lone pairs of electrons into bonds to enforce the octet rule
7. **Step seven:** For extended octets place the extra electrons on the central atom
8. **Step seight:** When numerous equivalent Lewis structures exist, the best structures would have low formal charges, with negative charges located on electronegative atoms

Atomic charges in a molecule

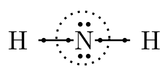
The electron-dot structure of a molecule results from counting the overall number of valence electrons of the molecule given that each atom brings a different number of valence electrons (n_e^{free}). For example, two H atoms bring one electron each, whereas O brings two electrons, giving a total of six valence electrons. When arranging the electron pairs in the molecule, the number of electrons surrounding an atom is called the valence of this atom in the molecule (n_e^{bonded}). We calculate the number of valence electrons of an atom in a molecule by accounting for the number of lone pairs on this atom and half the number of bonds:

$$n_e^{bonded} = \text{number of lone pairs} + 1/2\text{number of bonds}$$

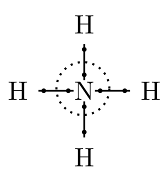
It is important to note that the valence of a free atom and the valence of this atom in a molecule is not necessarily the same. Indeed the difference between the valence electron of a free atom and the same atom in a molecule is the effective charge of that atom in the molecule, Q:

$$Q_{eff} = n_e^{free} - n_e^{bonded}$$

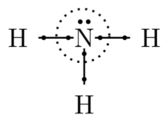
When the valence of an atom in a molecule is larger than the valence of the free atoms we have negative effective charges. In the example below



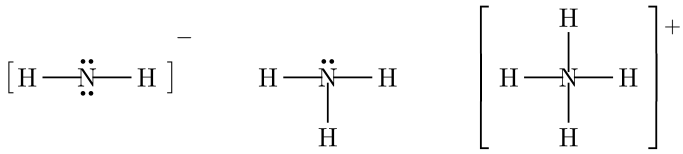
the number of electrons surrounding nitrogen is six electrons, more than the number of electrons originally brought to the molecule (five). We conclude that the atoms have a negative charge, and the effective atomic charge of nitrogen is Q=-1. In the next example,



the central atom, nitrogen, still has five valence electrons. After counting the electrons surrounding nitrogen, this time we find that this atom is surrounded by four electrons, less than the number of electrons originally brought to the molecule. We can conclude that nitrogen has a positive charge. In particular, the effective atomic charge of nitrogen in this molecule is the number of valence electrons minus the number of surrounding electrons. In this case, the atomic charge is Q=+1. When the valence of an atom in a molecule is the same as the valence of the free atoms we have zero effective charges. In this last example



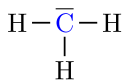
the central atom, nitrogen, has five valence electrons. After counting the electrons surrounding nitrogen–remember in a bond each atom shared an electron and hence each line around an atom counts as one electron–we find that this atom is surrounded by five electrons. As the number of valence electrons brought to the molecule is the same as the number of electrons surrounding the atom, we say the atomic charge of this atom is zero (Q=0). Hence, the atom is neutral. In all the molecules above, hydrogen remains neutral and hence the atomic charge of nitrogen corresponds to the molecular charge of the molecule. We can hence summarize the three scenarios indicated, as we have a neutral molecule in the center, a positive molecule on the right, and a negative molecule on the left.



The atomic charge of an isolated atom can be well-defined. However, the atomic charge of an atom in a molecule is arbitrarily defined, and more than one definition exists. Formal charges are one of the possible definitions of atomic charges in a molecule, whereas redox numbers are an alternative definition of atomic charges in molecules. None of these definitions is exactly correct. For example, redox numbers tend to overestimate the atomic charges, as they assume that all shared electrons in a bond belong to the most electronegative atom. Normally, negative formal charges tend to reside on electronegative atoms and not on electropositive atoms. At the same time, the sum of all effective charges needs to give the overall charge of the species. Furthermore, atoms in molecules tend to achieve formal charges as close to zero as possible. One can use formal charges to assess the validity of a Lewis structure. When comparing a series of equivalent Lewis structures for a molecule, the structures that best describe the bonding in the molecule tend to be those with small effective charges located on electronegative atoms.

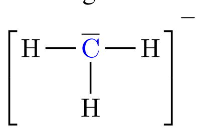
Sample Problem 23

Indicate the atomic charges of the blue highlighted atom



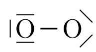
SOLUTION

The carbon atoms brings four electrons and in the molecule it is surrounded by eight electrons, five of which belongs to it. Hence the charge of C is -1; this means that carbon has one extra electron. Each hydrogen brings one electron and in the molecule each hydrogen has one electron (they share two electrons with C, one for C and one for H). The final lewis structure with the local charge of carbon can be indicated as:



STUDY CHECK

Indicate the atomic charges of all atoms in the Lewis structure below



▶Answer: left oxygen -1; right oxygen $+1$

Molecular shape

Molecules consist of arrangements of atoms presented in different forms. Let us use as an example the H_2O molecule, which contains two hydrogen atoms and one oxygen. Knowing that both hydrogens are connected to oxygen through a covalent bond, one can envision several molecular geometries such as a linear geometry or maybe a v-shaped geometry with oxygen at the point. The geometry of a molecule determines its properties, and small geometrical changes can have severe consequences on the functioning of molecules. For example, at high temperatures, when proteins in the body denaturalize losing their unique structure they also lose their functionality. The goal of this section is to identify the approximate shape of a given molecule.

The VSEP model

The VSEPR model, also known as the valence shell electron-pair repulsion model, is a model that predicts the geometries of molecules made of nonmetals. This model predicts the atomic arrangement of the molecules with an emphasis on the shape of the arrangement. However, it is not a very accurate model to predict geometries and there are better methods to obtain molecular geometries. The model predicts, for example, that water molecules have a v-shaped geometry and not a linear geometry while giving an estimate of the angle between the two O-H bonds. Still, VSEPR is not accurate enough to predict the O-H bond length and more advanced methods should be used for this purpose. The VSEPS model is based on the premise that the structure around a given atom results from minimizing the electron-pair repulsion. This way, the bonding and nonbonding pairs of electrons around a central atom are differently accounted for. Let us analyze a few cases in which the central atom only has bonding pairs of electrons. For example, the BeH_2 molecule has two bonding pairs around Be and the arrangement that maximizes the distance between both pairs hence minimizing repulsion is a linear arrangement. Hence, the BeH_2 molecule is linear with a 180° angle between both Be-H bonds. Another example would be the BH_3 molecules, a molecule with three bonding pairs. The geometry that maximizes the distance between the three pairs hence minimizing repulsion is a trigonal planar structure in which the three bonding pairs are in the same plane with an angle of 120° between the three bonds. A final example would be the methane molecule (CH_4), a molecule with four bonding pairs. A tetrahedral arrangement with 109.5° between the C-H bonds is the most stable arrangement for this case.

VSEP model for 5 and 6 bonds

For the case of five bonds, the geometrical arrangement that minimizes the electron-pair interaction is a trigonal bipyramidal arrangement consisting of two pyramidal arrangements sharing a common base. The PH_5 molecule presents this arrangement. In this arrangement, there are two different bond angles: 90° between the vertical and in-plane and 120° for the in-plane bonds. Finally, the octahedral structure minimizes the pair repulsion in the case of six bonds and for example, the SH_6 molecule has an octahedral arrangement. In this arrangement, all bonds have a 90° angle. All atomic arrangements discussed above are presented in the diagram below.

ABE Molecular code

We will use the ABE code to identify the molecular geometry of more complex molecules, with bonds and lone pairs. This code is based on the Lewis structure of the molecule, with B refers to the number of atoms connected to the central atom in the molecule (number of bonded atoms), and E is the number of lone pairs on the central atom. The overall number of bonded atoms and lone pairs is called the number of electron groups. Corresponding geometry for different ABE codes is tabulated. For example, an AB_2 molecule will be linear, whereas an AB_2E_2 is bent. The electron-dot structure of water and ammonia are:

$$\mathrm{H}-\overline{\underline{\mathrm{O}}}-\mathrm{H} \quad \text{and} \quad \begin{array}{c}\mathrm{H}-\overline{\mathrm{N}}-\mathrm{H}\\ | \\ \mathrm{H}\end{array}$$

Water has two bonds with the central atom and hence two Bs and two lone pairs on top of the central atom and hence two Es. The ABE code of water is AB_2E_2 and its geometry is bent. The ABE code of ammonia is AB_3E, as the molecule has three atoms connected to the central nitrogen and N has a single lone pair. Its geometry would be trigonal pyramidal. Angles between the different bonds for the different atomic arrangements are also tabulated. For example, the angle between the two H-O bonds of water would be 104.5°, whereas the angle between two of the N-H bonds of ammonia would be 107°. The overall number of bonding and lone pairs is referred to as the number of electron regions and the molecular geometry of the molecule is not necessarily the geometry of the electron regions. For example, the molecule methane has four bonds and a tetrahedral geometry. Ammonia has two bonds and two lone pairs. The geometry of the electron regions is also tetrahedral with three bonds pointing toward the lower part of the tetrahedron and the lone pair pointing toward the upper part. At the same time, the molecular geometry of ammonia is trigonal pyramidal. For the case of water, we have that again the geometry of the four electron regions is tetrahedral whereas the molecular geometry is bent. We can also conclude that lone pairs require more room than bonding pairs and this has an impact on the molecular angles. For example, the angle between two bonds in a tetrahedron is 109.5° being this value is the same as the molecule angles of methane. Differently, the molecular angles of ammonia–a molecule with one lone pair–are 107°, and the molecular angle of water–a molecule with two lone pairs–is 104.5°. These results indicate that as the number of lone pairs increases the bonding pairs are more squeezed together.

Steps to use the VSEPR model

The following steps can be used to obtain the molecular geometry using the VSEPR model:

1. **Step one:** Identify the central and the peripheral atoms.
2. **Step two:** Obtain the Lewis structure of the molecule
3. **Step three:**Obtain the ABE code with B being the number of peripheral atoms and E being the number of lone pairs. A represents the central atom. Multiple bonds (double, triple) count as a single B.
4. **Step four:** Use Table **??** to obtain the molecular geometry

STUDENT INFO

Name: ______________________ Date: ______

Pre-lab Questions

Compounds and their structure

---

1. Calculate the number of valence electrons for the following atoms and molecules:

$BeCl_2$	______________________	CFH_3	______________________
NF_3	______________________	SH_2	______________________

2. Draw the lewis structure of the following compounds: NF_3, CFH_3, $BeCl_2$.

STUDENT INFO

Name: ____________________ Date: ______

Results
EXPERIMENT

Compounds and their structure

---

1. Lewis structure of molecules

1. Calculate the number of valence electrons for the following atoms and molecules:

N	____________________	CO_3^{2-}	____________________
O	____________________	SO_4^{2-}	____________________
P	____________________	ICl_4^-	____________________
Br	____________________	CH_2Cl_2	____________________

2. Draw the lewis structure of the following compounds: H_2O, HF, HCl.

3. Draw the lewis structure of the following compounds and indicate their polarities: CH_4, CH_2Cl_2, $CHCl_3$.

Formula	Lewis Structure	# valence e^-	# e^- pairs	Geometry	Angles	Polar?
NH_3						
H_2O						
CH_4						
CH_2Cl_2						

Formula	Lewis Structure	# valence e^-	# e^- pairs	Geometry	Angles	Polar?
PCl_3						
ClO_2^-						
BrO_2^-						
CCl_3^-						

EXPERIMENT 0

Chemical reactions and equations

A. Goal

The goal of this laboratory experiment is to practice ballancing equations while observing chemical reactions happen. The goal is also to understand the existing different types of chemical reactions.

B. Materials

- ☐ Magnesium ribbon
- ☐ Bunsen burner
- ☐ metallic Zn and Cu
- ☐ a 1M $CuSO_{4(aq)}$ solution
- ☐ an 1M $HCl_{(aq)}$ solution
- ☐ a series of 0.1M $CaCl_{2(aq)}$,$Na_3PO_{4(aq)}$,$FeCl_{3(aq)}$, and $KSCN_{(aq)}$ solutions
- ☐ $Na_2CO_{3(s)}$
- ☐ a 3% $H_2O_{2(aq)}$ and 0.1M $KCl_{(aq)}$ solution

C. Background

Chemical reactions

When we eat we burn food with molecular oxygen (O_2) to produce carbon dioxide and water. Similarly, when we start the engine of the car to go to work, gasoline burns to produce the same chemicals: CO_2 and H_2O. These are two examples of chemical reactions, but there are many other examples. Nitrogen from the air reacts with hydrogen to produce ammonia, a common chemical used in the production of fertilizers. This section covers the basics of chemical reactions. You will learn how to balance reactions and how to classify reactions.

Simple chemical reactions

Magnesium is a metal that reacts with oxygen to produce magnesium oxide. Magnesium is solid $Mg_{(s)}$ whereas oxygen is gas and contains two oxygen atoms per molecule $O_{2(g)}$. Magnesium oxide, the result of the reaction, is solid $MgO_{(s)}$. The reaction between magnesium and oxygen to produce magnesium oxide

$$2\,Mg_{(s)} + 1\,O_{2(g)} \longrightarrow 2\,MgO_{(s)}$$

Mg and O_2 combine–that is why we use a plus sign– to produce MgO–we use an arrow to indicate that a chemical is being produced. Also, the symbols *(s)* or *(g)* indicates solid or gas state. The reactants are located before the arrow and the products are after the arrow. The numbers in front of the reactants and products (2,1 and 2) are called stoichiometric coefficients, and we will talk more about them in the following sections.

Reading a chemical reaction

Chemical reactions can be read in words. To read a chemical reaction you need to connect the reactants with the word "react" and then use the words "to produce" and after that, you need to read the products. The numbers in front of the reactants and products represent the number of moles, and you need to include those numbers in the reading. For example, the following reaction

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

should be read as: "two moles of Mg react with one mole of O_2 to produce two moles of MgO".

Balanced chemical reactions

Chemical reactions contain molecules, which are made of atoms. Some chemical reactions are balanced, and others need to be balanced. To identify a balanced reaction, you should use the stoichiometric coefficients and the indexes in the molecular formulas to break down the reactants and products into atoms. In a balanced chemical reaction, the atoms of reactants should be the same as the atoms of the products. Consider the following reaction,

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

The table below shows all reactants and products in the form of atoms.

$2\,Mg\,(s) + O_2(g)$	$\longrightarrow$	$2\,MgO\,(s)$
Reactants	$\longrightarrow$	Products
2Mg	✓	2Mg
O_2=2O	✓	2O

The number of Mg atoms in the reactants and products is the same and equals two. On the other hand, the number of O atoms in the reactants and products is the same, equal to two. For this reason, we say this reaction is *balanced*. Now consider the following reaction:

$$C\,(s) + O_2(g) \longrightarrow CO\,(g),$$

The number of C atoms in the reactants and products is the same and equals one. In contrast, the number of O atoms in the reactants and products differ, and for this reason, we say this reaction is *not balanced*.

$C\,(s) + O_2(g)$	$\longrightarrow$	$CO\,(g)$
Reactants	$\longrightarrow$	Products
1C	✓	1C
O_2=2O	✗	O

Balancing chemical reactions

To balance a reaction, we need to introduce the stoichiometric coefficients that make the number of atoms of reactants and products the same. To balance the number of oxygens, we will multiply CO by two, and that will give us two oxygens and two carbons as well. If we do this, now the carbon atoms of reactants and products will not be the same. We can solve this by multiplying C(s) by two. The following table summarizes the changes we made:

$2\,C\,(s) + O_2(g)$	$\longrightarrow$	$2\,CO\,(g)$
Reactants	$\longrightarrow$	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 24

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

Five types of reactions

Most chemical reactions can be classified according to five types: combination, decomposition, single replacement, double replacement, and combustion.

In a *combinations reaction* two reactants combine to generate a product. An example of a combination is the reaction between Mg and oxygen to produce MgO:

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

In a *decomposition reaction* a single reactant breaks down into several products. An example of a decomposition reaction is the thermal reaction of $CaCO_3$ to produce calcium oxide (CaO) and carbon dioxide

$$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g) \text{ (decomposition)}$$

In a single replacement reaction, an element replaces another element in a chemical. An example would be the reaction of Zn with HCl, in which Zn replaces hydrogen:

$$Zn(s) + 2\,HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g) \text{ (Single replacement)}$$

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

$$AgNO_3(aq) + NaCl(aq) \longrightarrow NaNO_3(aq) + AgCl(s) \text{ (Double replacement)}$$

Finally, in a combustion reaction, a carbon-based chemical reacts with oxygen to produce carbon dioxide and water. An example would be the combustion of methane (CH_4):

$$CH_4(g) + 2\,O_2(g) \longrightarrow 2\,H_2O(g) + CO_2(g) \text{ (Combustion)}$$

D. Procedure

Magnesium and oxygen

Step 1: – Obtain a 1-in strip of magnesium.

Step 2: – Start the Bunsen burner and place the strip in the blue cone of the flame using crucible tongs to hold the strip.

Step 3: – As soon as the metal starts to burn, remove it from the flame without direcly looking into the buring flame of magnesium which can damage your sight.

Step 4: – In the Experiment section, describe the appearance of the reactants and products of this reaction.

Step 5: – In the Experiment section, balance the reaction involved.

Step 6: – In the Experiment section, identify the type of reaction (combination, decomposition, single replacement, double replacement, combustion).

Step 7: – Use the disposals to get rid of any reactant or product.

Zinc and copper(II) sulfate

Step 1: – Pour 3mL (aproximately 60 drops) of a 1M $CuSO_4$ solution into a test tube. You can use a measuring cylinder to measure volume. Place the test tube into a test tube rack.

Step 2: – Obtain a small piece of zinc and place it into the test tube containing $CuSO_4$.

Step 3: – Wait 30 min for the reaction to complete. In the meantime you can proceed with the rest of the experiment.

Step 4: – In the Experiment section, describe the appearance of the reactants and products of this reaction.

Step 5: – In the Experiment section, balance the reaction involved.

Step 6: – In the Experiment section, identify the type of reaction (combination, decomposition, single replacement, double replacement, combustion).

Step 7: – Use the disposals to get rid of any reactant or product.

Metals and hydrochloric acid

Step 1: – Pour 3mL (aproximately 60 drops) of a 1M HCl solution into a test tube. You can use a measuring cylinder to measure volume. Place the test tube into a test tube rack.

Step 2: – Obtain small pieces of Zn, and place it into the test tube containing the acid.

Step 3: – Observe the reaction and record any evidence of reaction (heat exchange, bubbles, color change, solid appearing).

Step 4: – In the Experiment section, describe the appearance of the reactants and products of this reaction.

Step 5: – In the Experiment section, balance the reaction involved.

Step 6: – In the Experiment section, identify the type of reaction (combination, decomposition, single replacement, double replacement, combustion).

Step 7: – Now, repeate the procedure with Cu.

Step 8: – Now, repeate the procedure with Mg.

Step 9: – Use the disposals to get rid of any reactant or product.

$CaCl_2$ and Na_3PO_4

Step 1: – Pour 3mL (aproximately 60 drops) of a 0.1M $CaCl_2$ solution into a test tube. You can use a measuring cylinder to measure volume. Place the test tube into a test tube rack.

Step 2: – Pour 3mL (aproximately 60 drops) of a 0.1M Na_3PO_4 solution into a test tube. You can use a measuring cylinder to measure volume. Place the test tube into a test tube rack.

Step 3: – Pour the content of one test tube into the other.

Step 4: – Observe the reaction and record any evidence of reaction (heat exchange, bubbles, color change, solid appearing).

Step 5: – In the Experiment section, describe the appearance of the reactants and products of this reaction.

Step 6: – In the Experiment section, balance the reaction involved.

Step 7: – In the Experiment section, identify the type of reaction (combination, decomposition, single replacement, double replacement, combustion).

Step 8: – Use the disposals to get rid of any reactant or product.

$FeCl_3$ and KSCN

Step 1: – Pour 3mL (aproximately 60 drops) of a 0.1M $FeCl_3$ solution into a test tube. You can use a measuring cylinder to measure volume. Place the test tube into a test tube rack.

Step 2: – Pour 3mL (aproximately 60 drops) of a 0.1M KSCN (names potassium thiocyanide) solution into a test tube. You can use a measuring cylinder to measure volume. Place the test tube into a test tube rack.

Step 3: – Pour the content of one test tube into the other.

Step 4: – Observe the reaction and record any evidence of reaction (heat exchange, bubbles, color change, solid appearing).

Step 5: – In the Experiment section, describe the appearance of the reactants and products of this reaction.

Step 6: – In the Experiment section, balance the reaction involved.

Step 7: – In the Experiment section, identify the type of reaction (combination, decomposition, single replacement, double replacement, combustion).

Step 8: – Use the disposals to get rid of any reactant or product.

HCl and Na_2CO_3

Step 1: – Pour 3mL (aproximately 60 drops) of a 1M HCl solution into a test tube. You can use a measuring cylinder to measure volume. Place the test tube into a test tube rack.

Step 2: – Pour a peasize of Na_2CO_3, a solid reagent, into the test tube containing the acid.

Step 3: – Observe the reaction and record any evidence of reaction (heat exchange, bubbles, color change, solid appearing). For this experiment, place a lighted match insude the neck of the test tube and record what you see.

Step 4: – In the Experiment section, describe the appearance of the reactants and products of this reaction.

Step 5: – In the Experiment section, balance the reaction involved.

Step 6: – In the Experiment section, identify the type of reaction (combination, decomposition, single replacement, double replacement, combustion).

Step 7: – Use the disposals to get rid of any reactant or product.

H_2O_2 and KI

Step 1: – Pour 3mL (aproximately 60 drops) of a 3% H_2O_2 (names hydrogen peroxide) solution into a test tube. You can use a measuring cylinder to measure volume. Place the test tube into a test tube rack.

Step 2: – Pour 3mL (aproximately 60 drops) of a 0.1M KI solution into a test tube (this chemical acts as a catalyst). You can use a measuring cylinder to measure volume. Place the test tube into a test tube rack.

Step 3: – Pour the content of one test tube into the other.

Step 4: – Observe the reaction and record any evidence of reaction (heat exchange, bubbles, color change, solid appearing).

Step 5: – In the Experiment section, describe the appearance of the reactants and products of this reaction.

Step 6: – In the Experiment section, balance the reaction involved.

Step 7: – In the Experiment section, identify the type of reaction (combination, decomposition, single replacement, double replacement, combustion).

Step 8: – Use the disposals to get rid of any reactant or product.

STUDENT INFO

Name: ________________________ Date: ______

Pre-lab Questions

Chemical reactions and equations

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1. Balance the following reaction:

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

2. Balance the following reaction:

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

3. Balance the following reaction:

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

4. Classify next reaction as combination, decomposition, single replacement, double replacement, or combustion:

$$2\,RbNO_{3(aq)} + BeF_{2(aq)} \longrightarrow Be(NO_3)_{2(aq)} + 2\,RbF_{(aq)}$$

STUDENT INFO	
Name: ____________________	Date: ______

Results
EXPERIMENT

Chemical reactions and equations

Magnesium and oxygen

Reactants appareance	
Products appareance	
Sign of reaction (heat exchange, bubbles, color change, solid appearing)	
Type of reaction (combination, decomposition, single replacement, double replacement, combustion)	
$\_\_\_Mg(s) + \_\_\_O_2(g) \longrightarrow \_\_\_MgO(s)$	

Zinc and copper(II) sulfate

Reactants appareance	
Products appareance	
Sign of reaction (heat exchange, bubbles, color change, solid appearing)	
Type of reaction (combination, decomposition, single replacement, double replacement, combustion)	
$\_\_Zn(s) + \_\_CuSO_4(aq) \longrightarrow \_\_Cu(s) + \_\_ZnSO_4(aq)$	

Metals and hydrochloric acid

Reactants appareance	
Products appareance	
Sign of reaction (heat exchange, bubbles, color change, solid appearing)	
Type of reaction (combination, decomposition, single replacement, double replacement, combustion)	
$\_\_Zn(s) + \_\_HCl(aq) \longrightarrow \_\_ZnCl_2(aq) + \_\_H_2(g)$	

Reactants appareance	
Products appareance	
Sign of reaction (heat exchange, bubbles, color change, solid appearing)	
Type of reaction (combination, decomposition, single replacement, double replacement, combustion)	

$$\_\_\_ Cu(s) + \_\_\_ HCl(aq) \longrightarrow \_\_\_ CuCl_2(aq) + \_\_\_ H_2(g)$$

Reactants appareance	
Products appareance	
Sign of reaction (heat exchange, bubbles, color change, solid appearing)	
Type of reaction (combination, decomposition, single replacement, double replacement, combustion)	

$$\_\_\_ Mg(s) + \_\_\_ HCl(aq) \longrightarrow \_\_\_ MgCl_2(aq) + \_\_\_ H_2(g)$$

$CaCl_2$ and Na_3PO_4

Reactants appareance	
Products appareance	
Sign of reaction (heat exchange, bubbles, color change, solid appearing)	
Type of reaction (combination, decomposition, single replacement, double replacement, combustion)	

$$\_\_\_ CaCl_2(aq) + \_\_\_ Na_3PO_4(aq) \longrightarrow \_\_\_ Ca_3(PO_4)_2(aq) + \_\_\_ NaCl(aq)$$

$FeCl_3$ and KSCN

Reactants appareance	
Products appareance	
Sign of reaction (heat exchange, bubbles, color change, solid appearing)	
Type of reaction (combination, decomposition, single replacement, double replacement, combustion)	

$$\_\_\_ FeCl_3(aq) + \_\_\_ KSCN(aq) \longrightarrow \_\_\_ Fe(SCN)_3(aq) + \_\_\_ KCl(aq)$$

HCl and Na_2CO_3

Reactants appareance	
Products appareance	
Sign of reaction (heat exchange, bubbles, color change, solid appearing)	
Type of reaction (combination, decomposition, single replacement, double replacement, combustion)	
$\_\_Na_2CO_3(s) + \_\_HCl(aq) \longrightarrow \_\_CO_2(g) + \_\_H_2O(l) + \_\_NaCl(aq)$	

H_2O_2 and KI

Reactants appareance	
Products appareance	
Sign of reaction (heat exchange, bubbles, color change, solid appearing)	
Type of reaction (combination, decomposition, single replacement, double replacement, combustion)	
$\_\_H_2O_2(aq) \xrightarrow{KI} \_\_H_2O(l) + \_\_O_2(g)$	

STUDENT INFO

Name: ____________________ Date: ______

Post-lab Questions

Chemical reactions and equations

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1. Balance the following reaction:

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

2. Classify next reaction as combination, decomposition, single replacement, double replacement, or combustion:

$$Pb_{(s)} + FeSO_{4(s)} \longrightarrow PbSO_{4(s)} + Fe_{(s)}$$

3. Classify next reaction as combination, decomposition, single replacement, double replacement, or combustion:

$$C_6H_{12(g)} + 9\,O_{2(g)} \longrightarrow 6\,CO_{2(g)} + 6\,H_2O_{(g)}$$

4. Why the reaction between sodium carbonate and hydrochloric acid can suffocate a lighted match?

5. Balance the following reactions:

A Liquid hexane C_6H_{14} reacts with molecular oxygen O_2 to form carbon dioxide and water.

B Solid potassium reacts with molecular oxygen O_2 to form potassium oxyde.

6. Complete and balance the following equations:

A $CaCO_3(s) \longrightarrow$ (Decomposition)

B $H_2(g) + F_2(g) \longrightarrow$ (Combination)

C $NaCl(aq) + AgNO_3(aq) \longrightarrow$ (Double displacement)

D $Zn(s) + CuF_2(aq) \longrightarrow$ (Single displacement)

EXPERIMENT 0

Moles and Chemical Formulas

A. Goal

The goal of this laboratory experiment is to compute the chemical formula of an oxide and to calculate the water percent in a hydrate.

B. Materials

- ☐ rectangular wood piece
- ☐ string
- ☐ ruler
- ☐ 1L graduated cylinder
- ☐ 1qt jar

C. Background

Converting moles into grams and into atoms

A standard way to measure chemicals in the lab is by weight. We can weigh different quantities and the larger the quantity the larger the weight. For a chemical, the weight of a mole is called the molar (or molecular) weight. For example, if we weigh a mole of water (H_2O) we will be weighing 18 grams of water, or if you weigh a mole of table salt (NaCl) the scale will show 58 grams. In this section, you will learn how to calculate the molar mass of a chemical and how to use this property to convert from weight to moles (and moles to weight).

Molar mass of a chemical

Chemicals are made of atoms, and each atom has a specific atomic weight (AW) listed in the periodic table. For example, the atomic weight of Na is 23 grams whereas the atomic weight of Cl is 35 g. The weight of all the atoms of a molecule is called the molecular weight (we call this also molar weight or MW). For example, the molecular weight of NaCl is 58 g, as the weight of Na and Cl is 23 and 35g. Another example would be water, H_2O with a molecular weight of 18g–as the atomic weight of H and O is 1 and 16 g, respectively, and the molecule has two H atoms. The units for molecular weight is $\frac{g}{mol}$, also written as g/mol or $g/cdotmol^{-1}$. To compute the molar mass of a molecule you need to break down the molecule into atoms using the coefficients in its formula. For example, the formula for acetic acid, the acidic chemical in vinegar, is $C_2H_4O_2$ which means a molecule contains 2C, 4H and 2O atoms. If you add the atomic masses of 2C, 4H, and 2O you will get $60g/mol$. If the chemical formula has a parenthesis, you need to open up the parenthesis to calculate the total number of atoms. As an example, $Ca(NO_3)_2$ contains 1Ca, 2N, and 6O, and its molar mass is $164.09g/mol$.

From moles to grams

The molar mass is used to convert moles to grams or grams to mol. For example, the molar mass of water is $18g/mol$. This means:

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

that is the same as

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

From grams to atoms

In the previous sections, we covered how to convert grams to moles, moles to molecules, or molecules to atoms. You can follow the diagram below to switch from one of these properties (atoms, molecules, moles, grams) to another. For example, if you want to convert grams into moles, you will only need one step and you will only have to use a single property: the molar mass. Differently, if you need to convert grams into molecules you will have to use two different steps and use two different properties: the molar mass and Avogadro's number.

Sample Problem 25

Calculate: (a) The atomic weight of Mg; (b) the molecular mass of sulfuric acid, H_2SO_4

SOLUTION

(a) According to the periodic table the atomic weight (AW) of Mg is $24.31g/mol$. (b) The molar mass of H_2SO_4 is the result of adding the atomic masses of 2H (AW=$1g/mol$) atoms, 1 S (AW=$32g/mol$) and 4O (AW=$16g/mol$) atoms, that gives $98.08g/mol$.

STUDY CHECK

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

▶Answer: $180.06g/mol$.

D. Procedure

1. Empirical formula of an oxyde The goal of this mini experiment is to calculate the formula of an oxide. You will do this by burning up a metal .

Step 1: – Obtain a crucible with a lid, a clay triangle and an iron ring attached to a ring stand. Place the covered crucible in the clay triangle on an iron ring attached to a ring stand. Adjust the height of the ring so that the bottom of the crucible will be in the hottest part of the flame. The correct arrangement of the equipment, crucible, and burner is shown in the figure (Right panel).

Step 2: – Learn how to use the Bunsen burner. Heat the covered crucible in the hottest part of the flame for about 5 min while keeping the lid ajar, making sure that the bottom of the crucible attains a red glow.

Step 3: – Stop the burner and allow the crucible to cool down completely. Weight the covered crucible and record the mass of the covered crucible.

Step 4: – Obtain 0.2 g of magnesium ribbon. Clean the surface of the metal with metallic wool until it shines. Cut the magnesium ribbon into tiny bits using scissors, and place them inside the crucible. Cover the crucible, obtain and record the mass again. Now you know the mass of the crucible+lid+Mg.

Step 5: – Set the crucible on the clay triangle with the lid on and heat the crucible in the hottest part of the flame another 5 min. Keep the lid close. Using the crucible tongs, lift the lid carefully by a slight amount. The metal should glow brightly without flames. Continue until all Mg is burned and the product does not glow.

Step 6: – Patiently cool down the crucible with lid. The content should be white or slightly gray. At this point, add a few drops of water using a plastic pipet on the crucible content. You might notice a smell of ammonia at this point.

Step 7: – Place the lid back onto the crucible, slightly ajar, and heat the crucible in the hottest part of the flame for 15 more minutes. After that time, allow the covered crucible and its content to cool down. Obtain the mass of the covered crucible.

Calculations

1 This is the mass of the empty and clean crucible with lid.

2 This is the mass of the clean crucible with lid and the Mg.

3 This is the mass of Mg added to the crucible: 2 – 1

4 This is the moles of Mg (Atomic weight 24.305 $g \cdot mol^{-1}$): $n_{Mg} = \frac{3\,g}{24.305\,g \cdot mol^{-1}}$

5 This is the mass of the clean crucible with lid and the product.

6 This is the mass of product: 5 – 1

7 This is the mass of O in the product: 6 – 3

8 This is the moles of O (Atomic weight 15.999 $g \cdot mol^{-1}$) in the product: $n_O = \frac{7\,g}{15.999\,g \cdot mol^{-1}}$

2. Thermal decomposition of a hydrate The goal of this mini experiment is to calculate the percentage of water contained in an hydrate. You will achieve this goal by heating up the hydrate and measuring its mass before and after heating. The difference in mass will be the mass of water contained in the hydrate.

Step 1: – Place a clean, covered crucible in a clay triangle on an iron ring attached to a ring stand. Adjust the height of the ring so that the bottom of the crucible will be in the hottest part of the flame. The correct arrangement of the equipment, crucible, and burner is shown in the figure (Right panel).

Step 2: – Learn how to use the Bunsen burner. Heat the covered crucible in the hottest part of the flame for about 5 min, making sure that the bottom of the crucible attains a red glow.

Step 3: – Stop the burner and allow the crucible to cool down completely.

Step 4: – Weight the covered crucible and record the mass of the covered crucible.

Step 5: – Weight about 1.5 g of the hydrate.

Step 6: – Add the the hydrate sample onto the crucible and cover the crucible again. Weight the covered crucible with the chemical and record the exact mass in the results sheet.

Step 7: – Heat up the crucible in the hottest part of the flame for about 15 min. The bottom of the crucible should be red hot during this time.

Step 8: – When the crucible is cool, weight and record the mass of the product.

Calculations

1 Record the mass of the empty crucible with the lid. Remember to weight the crucible in the balance only when completely cool.

2 Record the mass of the empty crucible with the lid with hydrate.

3 The mass of hydrate added to the crucible should be:

$$\text{Mass of hydrate} = 2 - 1$$

4 After you heat the crucible with hydrate a product will form. Weight the crucible and lid with the final product inside.

5 You should calculate the mass of product by doing:

$$\text{Mass Product} = 4 - 1$$

6 Calculate the mass % of water in the hydrate:

$$\frac{(\text{Mass hydrate}) - (\text{Mass Product})}{(\text{Mass of hydrate})} \times 100 = \frac{3 - 5}{3} \times 100$$

STUDENT INFO

Name: ____________________ Date: ______

Pre-lab Questions

Moles and Chemical Formulas

---

1. Fill the following conversion to convert 15g of N into moles (AW(N)=14$g \cdot mol^{-1}$)

$$15 \cancel{\text{ g of N}} \times \frac{\underline{\qquad\qquad} \text{ moles of N}}{\underline{\qquad\qquad} \cancel{\text{ g of N}}} = \underline{\qquad\qquad} \text{ moles of N.}$$

2. Fill the following conversion to convert 2 moles of O into moles (AW(O)=16$g \cdot mol^{-1}$)

$$2 \cancel{\text{ moles of O}} \times \frac{\underline{\qquad\qquad} \text{ g of O}}{\underline{\qquad\qquad} \cancel{\text{ moles of O}}} = \underline{\qquad\qquad} \text{ g of O.}$$

3. The thermal decomposition of 5g of an hydrate gives a final product mass of 2.5g. Calculate the percent of water in the hydrate.

4. Name or give the formula of the following compounds:

Magnesium sulfate	____________________
$MgSO_4 \cdot H_2O$	____________________
Barium chloride	____________________
$BaCl_2 \cdot 2\,H_2O$	____________________
Chromium(III) chloride	____________________
$CoCl_2 \cdot 6\,H_2O$	____________________
Nickel(II) sulfate heptahydrate	____________________

STUDENT INFO

Name: ______________________ Date: ______

Results
EXPERIMENT

Moles and Chemical Formulas

---

1. Empirical formula of an oxyde

1	Mass of empty crucible and lid (g)	____________
2	Mass of crucible and lid with Mg (g)	____________
3	Mass of Mg (g)	____________
4	Moles of Mg (mole)	____________
5	Mass of crucible and lid with oxide (g)	____________
6	Mass of oxide(g)	____________
7	Mass of O (g)	____________
8	Moles of O (moles)	____________

	Mg	O
Moles (moles) 8		
Moles/smallest number (round to closest integer)		
Empirical Formula		

2. Thermal decomposition of a hydrate

Hydrate decomposition Data

1	Mass of empty crucible and lid (g)	____________
2	Mass of crucible and lid with hydrate (g)	____________
3	Mass of hydrate (g)	____________
4	Mass of crucible, lid and product (g)	____________
5	Mass of product (g)	____________
6	Mass % of water	____________

STUDENT INFO

Name: ________________________ Date: ______

Post-lab Questions

Moles and Chemical Formulas

---

1. Calculate the formula of the oxide resulting of mixing Mg and O according to the respective valences of the elements?

2. Calculate the formula of the nitride resulting of mixing Mg and N according to the respective valences of the elements?

3. The product of burning 5 grams of a hydrate weights 4.5g. Calculate the water % mass of the hydrate.

4. The formula for an hydrate is $FeSO_4 \cdot 7\,H_2O$. Calculate the water % mass of the hydrate.

5. Name the following chemical: $FeSO_4 \cdot 7\,H_2O$.

EXPERIMENT 0

Solutions, Electrolytes, Concentration

A. Goal

The goal of this experiment is to identify the polarity of a given solute as well as the electrolyte character of a series of solutes. The goal is also to calculate the molarity of an already prepare solution.

B. Materials

- ☐ Eight test tubes in a rack with cyclohexane,sucrose, iodine, potassium permanganate and vegetable oil
- ☐ Two electrodes connected to a lightbulb and a set of electrolytes: NaCl, sucrose, HCl, CH_3-COOH,NH_3 , CH_3-CH_2OH, and NaOH
- ☐ A 250mL beaker with 200mL of water
- ☐ evaporating dish
- ☐ unknown solution
- ☐ hot plate, stand and ring

C. Background

Solutions and composition

Solutions are homogeneous mixtures of two components. The state of the matter of both components of the mixture or their polarity affects the formation of a solution. For example, a solution will not result from mixing oil and water as they have different polar characteristics and it will form from mixing table salt and water as both are polar chemicals. At the same time, the more solute you add to a solution the more concentrated the solution will be. This section covers polarity and the composition of solutions.

What makes a solution?

Solutions are homogeneous mixtures of a solute and a solvent (see Figure **??**). Homogeneous means that if you look at the mixture you will not be able to differentiate both components and you will only see it as a whole. In a solution, the solute is the component of the mixture in less amount, whereas the solvent is the component in a larger amount. Think about mixing sugar with water. Sugar is sweet and water tasteless. When you mix both, you form a solution of sugar (solute) in water (solvent) and you will not see sugar in the solution as it is dissolved. In this particular example, sugar will be the solute in the solution, as the sugar is in less amount than water. Is important to remember that a solution is a result or mixing a solute and a solvent:

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

Types of solutions

You can prepare different types of solutions by mixing a solid and a liquid, like when you mix sugar and water, or salt and water. You can create solutions as well by mixing two liquids or two solids. Examples are vinegar–a liquid solution of acetic acid (liquid) in water (liquid)–or steel– a solid solution that contains iron and carbon, both solids.

Empirical rules of polarity

The affinity between two chemicals is related to a concept called polarity. Molecules contain electrons and depending on the electron distribution within the molecules, molecules can be polar or non-polar. Molecules with an even electron distribution are non-polar as they have no permanent dipole moment. An example of this is H_2 molecule. Differently, HF is a polar molecule, as F concentrates more on the electron density of the molecule than H. The polar nature of substances–with a permanent dipole moment–is related to miscibility and molecules with similar polar character will mingle and mix creating a single visible phase (see Table 8). As an example, water (H_2O, polar) and methanol (CH_3OH, polar) will mix together. Differently, water (polar) and oil (non-polar) are immiscible due to their different polar nature and they will not mix. Even if the rules or polarity are based on the nature and structure of the molecule, one can use very simple empirical rules to classify molecules as polar or non-polar. These rules work in general well for the case of diatomic and very large molecules:

1. **Rule one:** Diatomic molecules made of the same element (e.g. H_2) are non-polar.
2. **Rul two:** Diatomic molecules made of different elements (e.g. HI) are polar.
3. **Rul three:**Poliatomic molecules (with more than four atoms) made of C and H (e.g. CH_4) are in general non-polar.
4. **Rul four:** Poliatomic molecules (with more than four atoms) containing C, H, and a different atom (e.g. CH_3F) are in general polar.

Mixing and polarity

A solution is formed when both the solute and the solvent mix. However, they will only mix if they have the same polarity. As an example, water (H_2O) is a polar molecule, and methanol (CH_3-OH) is too. Hence they will both mix and form a solution. If the elements of a mixture have different polarity they will not mix. An example is benzene (C_6H_6, nonpolar) and water, or for example oil (nonpolar) and water (polar).

Sample Problem 26

Classify the following molecules as polar or nonpolar: H_2, HCl, CH_3CH_3, and CH_3CH_2Cl.

SOLUTION

H_2 is a non-polar molecule, being a diatomic molecule containing two atoms of the same element. Differently HCl is polar. CH_3CH_3 is a non-polar poliatomic molecule made of C and H atoms, whereas CH_3CH_2Cl is polar.

STUDY CHECK

Classify the following molecules as polar or nonpolar: HF, Cl_2, C_2H_4, and C_2H_3Cl.

Table 8 Polarity and mixing

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

Concentration of solutions

The concentration of a solution refers to the amount of solute with respect to the amount of solution. The larger concentration the larger the number of solute particles with respect to the particles of solvent. Concentration is one of the most important properties of a solution as it affects the physical properties of a solution such as the freezing and boiling point. There are many different concentration units, such as molarity, mass percent concentration, or volume percent concentration. All these different units overall express the ratio between the particles–mass or volume–of solute and solvent.

Mass percent concentration

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

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

Volume percent concentration

The volume percent concentration (% v/v) is the volume of solute per volume of solution in percent form

$$\boxed{\%v/v = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100}$$

Mass volume percent concentration

The mass/volume percent concentration (% m/v) is the mass of solute per mL of solution in percent form.

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

Molarity concentration

The molarity (M) is the moles of solute per L of solution.

$$\boxed{M = \frac{\text{moles of solute}}{\text{L of solution}}}$$

Sample Problem 27

A NaCl solution is prepared by mixing 4g of NaCl (MW=58.4g/mol) with 50 g of water until a final volume of 52mL of solution. Calculate: (a) the mass percent (m/m) concentration; (b) the molarity.

SOLUTION

(a) to calculate the mass percent (m/m) we just need the grams of solute and the grams of solution–that is four plus fifty. Both numbers are already given:

$$m/m = \frac{\text{g of solute}}{\text{g of solution}} \times 100 = \frac{\text{4 g of solute}}{\text{54 g of solution}} \times 100 = 9.2\%$$

(b) To calculate molarity we need the moles of solute and the liters of solution. We have the mL of solution, that can be converted to L: 52mL= $5.2 \times 10^{-2}L$. To calculate the moles of solute, we will use the grams of solute and the molar mass to convert this value into moles: $4g/58.4g/mol$ =0.068moles. Plugging all values into the molarity formula:

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

STUDY CHECK

(a) A solution is prepared by mixing 8g of NaCl (MW=74g/mol) with water until a 250mL volume. Calculate the

> molarity; (b) A KCl solution is prepared by mixing 45g of KCl with 200g of H_2O. Calculate the percent (m/m) of the solution.

D. Procedure

1. Polarity and miscibility The goal of this mini-experiment is to identify the polarity of a given solute by mixing it with a polar and nonpolar solvent. To chemicals with the same polarity will mix with each other due to favorable interactions. Think about oil and soap, both are nonpolar chemicals and hence they mix well. Differently, chemicals with different polarity do not mix well. Think this time about water and oil. Water is polar and oil is nonpolar. Both chemicals do not mix well. By using a polar solvent (water) and a nonpolar solvent (cyclohexane) you will be able to track the polarity of a given solute by studying the miscibility of the solute with both solvents. Ultimately, polarity is due to differences in electronegativity an the existence of a permanent dipolar moment in the molecule. Small molecules (diatomic) made of different elements will always be polar, as the differences of electronegativity will not compensate with each other leading to a permanent dipole moment. Molecules made mainly of C and H will normally be nonpolar as both elements have not appreciable electronegativity differences.

Step 1: – Use eight test tubes in a rack. Four of these will be filled with 5 drops of water–a polar solvent–whereas the remaining four tubes will be filled with cyclohexane–a nonpolar solvent.

Step 2: – Add a few drops of a few crystals of the following solutes both in water and in cyclohexane. If the solute mixes with water that means it will be polar. If the solute mixes with cyclohexane that would mean it is nonpolar.

2. Electrolytes In this mini-experiment you will study the electrolyte character of a series of solutes with different nature. By means of two electrodes connected to a lightbulb you will be able to appreciate the degree these chemicals conduct electricity. If the lightbulb glows the chemical will be an electrolyte. Depending on the brightness of the glow the chemical will be a strong or weak electrolyte.

Step 1: – You or the professor will use a setup with two electrodes connected to a lightbulb. Place 20mL of the different solutions in the table below in a beaker.

Step 2: – Lower the electrodes to the solution and observe the glow.

Step 3: – Observe the glow and classify the chemical as nonelectrolyte, strong electrolyte or weak electrolyte.

3. Molarity of a solution The goal of this mini-experiment is to calculate the molarity of an already prepare solution. In order to calculate molarity, we need the moles of solute and the volume of solution. You will take a given solution volume by using a pipet. At the same time you will learn how to use a pipet–a very common chemistry measuring tool. Then you will evaporate the solution so that only the solute will remain. By weighting this solute and given the molar mass you will convert grams into moles and compute molarity.

Step 1: – Fill a 250mL beaker with 200mL of water. Set the beaker on a hot plate and start heating at medium high heat.

Step 2: – Weight an evaporating dish. Write down the mass in the table below.

Step 3: – Place the evaporating dish on top of the beaker so that it receives indirect heat. Use a metallic ring to secure the beaker.

Step 4: – Use a small beaker to measure approximately 20mL of the unknown solution. Use a 10-mL graduated pipet to transfer exactly 10mL of the solution into the evaporating dish. Weight the evaporating dish with the solution.

Step 5: – The solution will start to dry. When the evaporating dish is completely dry, stop the heater and wait for the dish to cool down. Weight the evaporating dish with the solute.

Calculate the molarity of the solution by using the following formula:

$$M = \frac{n_{solute}}{v_{solution}}$$

STUDENT INFO

Name: ______________________ Date: _____

Pre-lab Questions

Solutions, Electrolytes, Concentration

---

1. Define electronegativity.

2. Compare the electronegativity of hydrogen and chlorine. Will HCl be polar or nonpolar?

3. Compare the electronegativity of two hydrogen atoms. Will H_2 be polar or nonpolar?

4. Classify the following chemicals as ionic, covalent, organic chemical, organic acid or organic base: NaCl, CH_3-COOH, CH_3NH_2, HF, CO.

STUDENT INFO

Name: ______________________ Date: ______

Results
EXPERIMENT

Solutions, Electrolytes, Concentration

---

1. Polarity and miscibility

Solute	Soluble in Water (H_2O)?	Soluble in Cyclohexane (C_6H_{10})?	Polar/Nonpolar
I_2(iodine)			
Sucrose			
$KMnO_4$(potassium permanganate)			
Vegetable oil			

2. Electrolytes

Chemical	Light intensity (No light, weak light, strong light)	Electrolyte type (Non electrolyte/weak electrolyte/strong electrolyte)	Particles in solution (Molecules/ions/Molecules+Ions)
NaCl			
Sucrose			
HCl			
CH_3-COOH			
NH_3			
CH_3-CH_2OH			
NaOH			

3. Molarity of a solution

1	Mass of the evaporating dish (g)	__________
2	Volume of solution, $v_{solution}$ (L)	__________
3	Mass of the evaporating dish with the solution (g)	__________
3 − 1	Mass of the solution (g)	__________
4	Mass of the evaporating dish with dry solute	__________
4 − 1	Mass of solute, m_{Solute}	__________
$(4 - 1) \times \frac{1 \text{ mol NaCl}}{58 \text{ g NaCl}}$	Moles of solute, n_{solute} (mol)	__________

$M =$ __________

STUDENT INFO

Name: ______________________ Date: ______

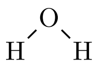
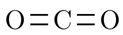
Post-lab Questions

Solutions, Electrolytes, Concentration

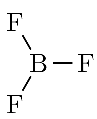
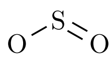
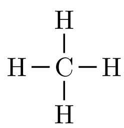
---

1. Indicate whether the following diatomic molecules are polar or nonpolar: Cl_2, N_2, O_2, HCl, HI, and HF.

2. Given the geometry of the following small polyatomic molecules are polar or nonpolar:

 and 

3. Given the geometry of the following small polyatomic molecules are polar or nonpolar:

  and 

4. Indicate whether the following chemicals are nonelectrolytes, weak electrolytes or strong electrolytes: NaF, CH_3-CH_3, $CH_3-CH_2-CH_2OH$, HF, and CH_3-CH_2-COOH.

5. Use your results to calculate the mass percent of the solution.

6. Calculate the volume of a 3M solution that contains 4 moles of solute.

7. Calculate the number of moles in 20mL of a 4M solution.

EXPERIMENT 0

Soluble and Insoluble Salts

A. Goal

The goals of this experiment are to test the solubility rules and to investigate the changes of solubility with temperature.

B. Materials

- ☐ spot plate
- ☐ Type A reagents: $NaCl(aq)$, $Na_2SO_4(aq)$ and $Na_3PO_4(aq)$
- ☐ Type B reagents: $Ca(NO_3)_2(aq)$ and $AgNO_3(aq)$
- ☐ KNO_3
- ☐ hot plate, beaker, ring and stand

C. Background

Electrolytes and insoluble compounds

On one hand, electrolytes are compounds that conduct electricity once dissolved in water. Differently, nonelectrolytes are compounds that do not conduct electricity once dissolved in water. On the other hand, insoluble compounds are not soluble in water, whereas soluble compounds can be dissolved in water. This section covers the properties of electrolytes and insoluble (and soluble) compounds. At the end of this section, you should be able to classify a chemical in terms of its electrolyte type and solubility character.

Solubility formula

Solubility (s) is the grams of a solute per 100 g of solvent:

$$\boxed{s = \frac{\text{g of solute}}{\text{100 g of solvent}}}$$

A saturated solution can be achieved when you fit the maximum amount of solute in the solvent. If you continue adding solute to a saturated solution it will precipitate and a solid will form.

Soluble and insoluble salts

Soluble compounds dissolve in water, whereas insoluble compounds do not. For example, barium chromate ($BaCrO_{4(s)}$) is an insoluble salt. How do we know that? Table 9 will help you predict the solubility of a salt. To do this, you need to start by assessing the right ion (the anion, $CrO_4{}^{2-}$) located in the left column of Table 9. After that, you need to assess the left ion (the cation, Ba^{2+}) located in the right column. If you follow this, you will see that chromate is insoluble and barium is not part of any exception. Let us predict for example the soluble/insoluble nature of $CaSO_4$, calcium sulfate. We start by looking for $SO_4{}^{2-}$ in the left column to find out is soluble. Next, we continue in the same line as $SO_4{}^{2-}$ and look for the ion in the left Ca^{2+}. In conclusion, even when $SO_4{}^{2-}$ is soluble, when combined with Ca^{2+}, we have that $CaSO_4$ is insoluble, and overall $CaSO_{4(s)}$ is insoluble.

Sample Problem 28

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

SOLUTION

(a) $K_2CO_{3(aq)}$ is soluble, as $CO_3{}^{-2}$ is insoluble but when combined with K^+ the salt becomes soluble. (b) All nitrates are soluble without exceptions. (c) $Ca(OH)_{2(aq)}$ is soluble.

STUDY CHECK

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

Table 9 Soluble and insoluble compounds

Ions that form *soluble* compounds...	... **except** when combined with
Group I ions (Na^+, Li^+, K^+, etc)	no exceptions
Ammonium ($NH_4{}^+$)	no exceptions
Nitrate ($NO_3{}^-$)	no exceptions
Acetate (CH_3COO^-)	no exceptions
Hydrogen carbonate ($HCO_3{}^-$)	no exceptions
Chlorate ($ClO_3{}^-$)	no exceptions
Halide (F^-, Cl^-, Br^-, I^-)	Pb^{2+}, Ag^+ and $Hg_2{}^{2+}$
Sulfate ($SO_4{}^{2-}$)	Ag^+, Ca^{2+}, Sr^{2+}, Ba^{2+}, $Hg_2{}^{2+}$ and Pb^{2+}
Ions that form *insoluble* compounds...	... **except** when combined with
Carbonates ($CO_3{}^{2-}$)	group I ions (Na^+, Li^+, K^+, etc) or ammonium ($NH_4{}^+$)
Chromates ($CrO_4{}^{2-}$)	group I ions (Na^+, Li^+, K^+, etc) or Ca^{2+}, Mg^{2+}or ammonium ($NH_4{}^+$)
Phosphates ($PO_4{}^{3-}$)	group I ions (Na^+, Li^+, K^+, etc) or ammonium ($NH_4{}^+$)
Sulfides (S^{2-})	group I ions (Na^+, Li^+, K^+, etc) or ammonium ($NH_4{}^+$)
Hydroxides (OH^-)	group I ions (Na^+, Li^+, K^+, etc) or Ca^{2+}, Mg^{2+}, Sr^{2+}or ammonium ($NH_4{}^+$)

Strong electrolytes

Strong electrolytes completely dissociate in water. Hence, in a solution of a strong electrolyte, you will only have ions and never molecules. Strong electrolytes are typically ionic compounds such as $MgCl_2$ or NaCl (table salt). We represent the dissociation of a strong electrolyte with a single arrow, meaning that the reaction proceeds to completion, and for the example below, in the solution, we will only have ions ($Mg^{2+}_{(aq)} + 2\,Cl^-_{(aq)}$) and not molecules ($MgCl_{2(s)}$):

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

Weak electrolytes partially dissociate in water, and this is indicated using a chemical reaction with a double arrow. Hence in a solution of a weak electrolyte, you will have ions as well as molecules at the same time. Examples of weak electrolytes are hydrofluoric acid, water, ammonia, or acetic acid. The dissociation of hydrochloric acid (HF) proceeds as:

$$HF_{(l)} \xrightleftharpoons{H_2O} H^+_{(aq)} + F^-_{(aq)}$$

Acetic acid (CH_3COOH) is an important weak electrolyte and its dissociation proceeds somehow in a peculiar way:

$$CH_3COOH_{(l)} \xrightleftharpoons{H_2O} CH_3COO^-_{(aq)} + H^+_{(aq)}$$

Nonelectrolytes

Nonelectrolytes do not dissociate in water. Hence a solution of a nonelectrolyte will only contain molecules and not ions. Examples of nonelectrolytes are carbon-based chemicals such as methanol, ethanol, urea, or sucrose. The dissociation of urea for example CH_4N_2O proceeds as:

$$CH_4N_2O_{(s)} \xrightarrow{H_2O} CH_4N_2O_{(aq)}$$

Identify the electrolyte character of a chemical

You can use Table 10 to identify the electrolyte character of a chemical. Ionic compounds are in general strong electrolytes, and most acids are as well. There are four important weak electrolytes: water, acetic acid, ammonia, and hydrofluoric acid. Covalent compounds are in general nonelectrolytes. Organic compounds, compounds based on carbon atoms (e.g. $C_{12}H_{22}O_{11}$) are in general nonelectrolytes.

Sample Problem 29

For the following chemicals indicate whether you will have in the solution (a) only ions, (b) ions and some molecules, or (c) molecules: NH_3, KOH, and $C_{12}H_{22}O_{11}$.

SOLUTION

Ammonia (NH_3) is a weak electrolyte and a solution of ammonia will contain ions and well as ammonia molecules. Differently KOH is a strong electrolyte and in solution you would find only ions (K^+ and OH^-). Sucrose ($C_{12}H_{22}O_{11}$) is a nonelectrolyte and in solution you will find molecules.

STUDY CHECK

For the following chemicals indicate whether you will have in the solution only ions, ions and some molecules, or molecules: (a) H_2SO_4, HNO_3 and CH_3OH.

Table 10 Different types of electrolytess

Electrolyte Type	Dissociation	Particles in solution	Examples
Strong	Fully	Mostly ions	Ionic Compounds and most acids and bases (hydroxides): NaCl, NaOH, HCl, $MgCl_2$, H_2SO_4, etc
Weak	Partially	Ions & molecules	NH_3, CH_3COOH (acetic acid), HF, H_2O
Nonelectrolytes	No	molecules	Most covalent compounds: CH_3OH (methanol), CH_3CH_2OH (ethanol), $C_{12}H_{22}O_{11}$ (sucrose), CH_4NO_2 (urea)

D. Procedure

1. Solubility rules The goal of this mini experiment is to test the solubility rules and predict the soluble character of a chemical.

Step 1: – Find a spot plate. Arrange in the following order the set of reactants Type A: NaCl(aq), Na_2SO_4(aq) and Na_3PO_4(aq). Arrange in the following order the set of reactants Type B: $Ca(NO_3)_2$(aq) and $AgNO_3$(aq).

Step 2: – Make mixtures of each pair of compounds listed in the results table by adding 2-3 drops of each solution in the same spot of the spot plate. When the resulting mixture is cloudy that means a precipitate has formed.

Step 3: – Write down the result on the Results table as soluble (S) or insoluble (I).

2. Change of solubility with temperature The goal of this mini experiment is to investigate the change of solubility of KNO_3 with temperature. You will do so by adding different amounts of solute and measuring the temperature at what the solute dissolved. You will work in different teams and each team will share their results with the rest of the class.

Step 1: – You will be assigned an amount of solute between 3 and 7 grams. Weight the solid and write down exactly how much solute did you use. If for example you are assigned 3g you can weight 3.1g.

Step 2: – Measure 5mL of water with a measuring cylinder. Add the liquid and the solid to a test tube that should be damped to a stand inside a water bath. Use a thermometer and start warming up the water bath.

Step 3: – Heat the solution with either a hot plate or a Bunsen burner until the solid dissolves completely. If you use a hot plate, make sure you secure the beaker with a ring. At that point, stop the heating and take the tube out of the bath keeping the thermometer inside the solution. When crystals start to appear, write down the temperature of the solution. You can calculate solubility in $g \cdot mL^{-1}$ using the formula:

$$\text{Solubility} = \frac{\text{mass of KNO}_3}{\text{5 mL of H}_2\text{O}} \times 100$$

Step 4: – When you have all results from the class, plot solubility (vertical axis) vs. temperature (horizontal axis).

STUDENT INFO

Name: ____________________ Date: ______

Pre-lab Questions

Soluble and Insoluble Salts

---

1. Name and predict the solubility (*soluble* or *insoluble*) of the following substances.

Compound	Name	Solubility
$CaCl_2$		
$NaNO_3$		
NH_4Br		
$AgCl$		
$Ni(OH)_2$		
Ag_3PO_4		
$Ca_3(PO_4)_2$		

2. Are salts more or less soluble in liquid at high or low temperature?

3. Think about a soda can. Are gases more or less soluble in liquid at high or low temperature?

STUDENT INFO	
Name: ______________________	Date: ______

Results

EXPERIMENT

Soluble and Insoluble Salts

---

1. Solubility rules

	NaCl	Na_3PO_4(aq)	Na_2SO_4
$Ca(NO_3)_2$(aq)			
$AgNO_3$(aq)			

(write *S* for soluble product and *I* for insoluble product.)

2. Change of solubility with temperature

Mass of KNO_3 (g)	Temperature when crystals appear (°C)	Solubility (g/ml)
–	100	13
3		
3.5		
4		
4.5		
5		
5.5		
6		
6.5		
7		

STUDENT INFO

Name: ____________________ Date: ______

Post-lab Questions

Soluble and Insoluble Salts

---

1. According to your graph, estimate solubility at 37 °C.

2. Solubility for a given chemical is 0.1 $g \cdot mL^{-1}$ at 30 °C. How many grams of solute will dissolve in 25mL of water at that temperature.

3. Solubility of table salt is 0.4 $g \cdot mL^{-1}$ at 25 °C. Will 50 grams of table salt dissolve in 100mL of water at that temperature.

EXPERIMENT 0

Reaction Rates & Chemical Equilibrium

A. Goal

The goal of this laboratory experiment is to identify the factors that impact the rate of a chemical reaction such as temperature, the presence of a catalyst, while testing Le Chatelier's principle.

B. Materials

- ☐ 0.1, 2M, and 3M HCl
- ☐ 400mL beaker
- ☐ ice
- ☐ thermometer
- ☐ $NaHCO_3$
- ☐ Mg
- ☐ 3% H_2O_2
- ☐ MnO_2, Zn, fresh and boiled potato
- ☐ 0.01M $Fe(NO_3)_3$ and 0.01M KSCN
- ☐ 1M $Fe(NO_3)_3$

C. Background

Le Chatelier principle

At this point, we have covered the idea of equilibrium and we have seen that the forward and reverse reactions have the same rate at equilibrium. Now, what happens if you alter this equilibrium? Le Châtelier principle claims a reaction will go back to its original equilibrium state by shifting left or right.

Le Chatelier principle

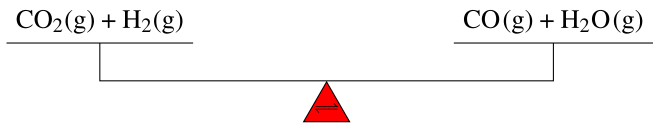
When a reaction is in equilibrium the forward and reverse reactions proceed at the same speed. Also in an equilibrium state, the concentrations of reactants and products have very specific values. Imagine that you create stress conditions by adding reactants or products or even changing the temperature. This stress will have an impact on the equilibrium and the reaction eventually will reach a new state of equilibrium by somehow counteracting this stress. Le Châtelier principle says that when stress is placed in a reaction (adding or removing reactant or products, increasing or decreasing temperature) the equilibrium will be shifted in the direction that relieves that stress. Table **??** displays different aspects regarding Le Châtelier's principle in terms of parameter change and consequence.

Change in concentration

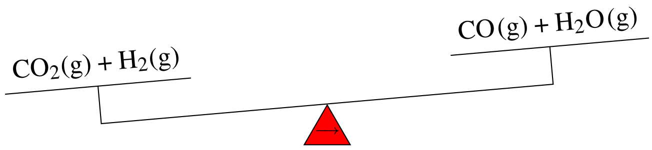
Let us consider the following equilibrium:

$$\mathrm{CO_2(g) + H_2(g) \rightleftharpoons CO(g) + H_2O(g)}$$

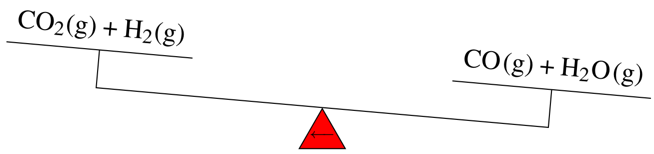
K_c for this reaction equals one at 1200K. This means that the concentration of reactants and products are the same. We can represent this using this balance or seesaw



If we add some CO_2 the equilibrium will be affected. To counteract this stress, the reaction will restore the equilibrium by decreasing the amount of CO_2. This can only be achieved by displacing the equilibrium to the right so that CO_2 is removed. Mind that CO_2 is consumed if the reaction moves from reactants $\longrightarrow$ to products and it is produced when going from products $\longleftarrow$ to reactants. We can represent this with the following seesaw.



As we added CO_2 the reactants now weigh more and hence the reaction has to proceed to the right $\longrightarrow$. Now imagine we remove some $CO_2(g)$. Again, the equilibrium will be affected and the reaction will restore its equilibrium state by doing the opposite, that is producing $CO_2(g)$ as the reaction proceeds from reactants $\longleftarrow$ to products. Again using the seesaw:



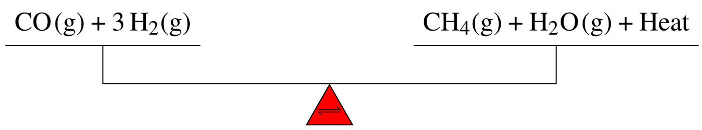
We can also add a different chemical that is not involved in the equilibrium. In this case, the equilibrium will not be affected by this change.

Temperature change

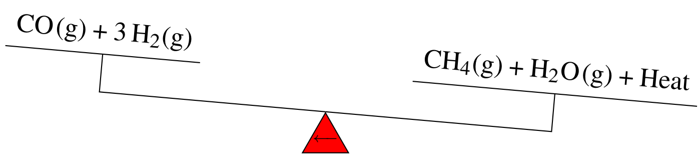
Let us consider the following equilibrium that produces heat–remember we describe these types of reactions as exothermic:

$$\mathrm{CO\,(g) + 3\,H_2(g) \rightleftharpoons CH_4(g) + H_2O\,(g) + Heat}$$

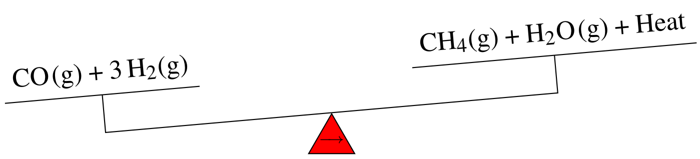
Again, this reaction is in equilibrium so we can use the same seesaw analogy.



If we increase the temperature of the system, the equilibrium will be affected. To go back to an equilibrium state the reaction will decrease the temperature of the container. As the reaction produces heat, a way to decrease the system temperature is to generate reactants ($\longleftarrow$). Again, using the scale that means:



Differently now if we decrease the temperature, the reaction will increase the temperature by going back to its equilibrium state going from reactants $\longrightarrow$ to products. This is because heat is produced as a product

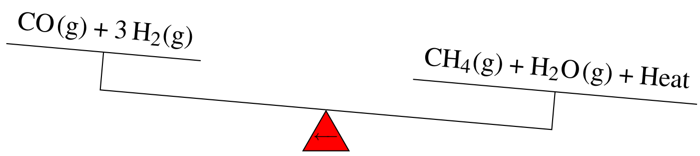


Volume change

We can also think about increasing or decreasing the volume in which the reaction takes place. This change will have an impact on the reaction equilibrium as the concentrations of reactants and products will be altered by this change. Changes in volume will shift the reaction towards the left or right depending on the overall stoichiometric change of the reaction, that is on whether the reaction produces or consumes molecules. For reactions that generate matter, that is, in the case that $\Delta n > 0$, increasing the volume will follow the increase of the number of moles. In other words, by increasing the volume, the equilibrium will shift towards the products, that is towards the right. For reactions that consume matter ($\Delta n < 0$), increasing the volume will shift the equilibrium towards the reactants, that is towards the left. For example, the reaction below consumes molecules:

$$\mathrm{CO\,(g) + 3\,H_2(g) \rightleftharpoons CH_4(g) + H_2O\,(g)}$$

if we increase the volume of the container in which the reaction takes place, the equilibrium will shift toward the left:



The opposite shift will follow a volume decrease as the reaction shift towards the right.

Sample Problem 30

For the next endothermic reaction indicate whether the reaction will shift right ($\longrightarrow$) or left ($\longleftarrow$) after the following changes:

$$\mathrm{N_2(g) + O_2(g) + Heat \rightleftharpoons 2\,NO\,(g)}$$

(a) adding reactants (b) adding products (c) decreasing the temperature.

SOLUTION

(a) Adding reactants always displaces the equilirbium so that reactants are consumed, hence the reaction will proceed $\longrightarrow$. (b) After adding products the reaction will tend to reduce the amount of products, and hence it will go $\longleftarrow$. (c) The reaction is endothermic that means that it consumes heat. If we decrease the temperature it will tend to increase the temperature and hence heat needs to be formed. This will only happen if the reaction proceeds ($\longleftarrow$).

STUDY CHECK

For next exothermic reaction indicate whether the reaction will shift right ($\longrightarrow$) or left ($\longleftarrow$) after the following changes:

$$\mathrm{C\,(g) + O_2(g) \rightleftharpoons CO_2(g) + Heat}$$

(a) removing reactants (b) removing products (c) decreasing the temperature.

D. Procedure

1. Factors that affect the chemical rate. Effect of temperature The goal of next three mini-experiments is to identify the factors that impact the rate of a chemical reaction. Reactions proceed at a certain rate, some are fast others are slow. By playing with a few factor you can increase the speed of a reaction generating more products in less time, or even slow down a reaction, avoiding the formation of products. You will study three different reaction and address the impact of three factors (1) the concentration of reactants, (2) temperature and (3) adding a catalyst on the chemical rate. This mini-experiment addresses the impact of temperature on reaction rate of the decomposition of sodium hydrogen carbonate:

$$\mathrm{NaHCO_3(aq) + H_3O^+(aq) \longrightarrow CO_2(g) + 2\,H_2O(l)}$$

Step 1: – Place 10mL of 0.1 M HCl in each of two different test tubes.

Step 2: – Place one of the test tubes in a cold bath with ice–this is a 400mL beaker half-filled with ice and water. Cool the test tube to a temperature of $10°$C.

Step 3: – Place the second test tube in a hot bath–this is a 400mL beaker half-filled with hot water from the tab. Warm up the test tube to a temperature close to $40°$C.

Step 4: – Remove both test tubes and place them in a test tube rack. Immediately, add one scoop of $NaHCO_3$ (sodium bicarbonate or sodium hydrogencarbonate) to each tube. You will observe the appearance of bubbles. Write down which test tube produces bubbles first.

2. Factors that affect the chemical rate. Effect of reactant concentration This mini-experiment addresses the impact of temperature on reaction rate of the reaction of magnesium with hydrochloric acid:

$$\mathrm{Mg(s) + 2\,HCl(aq) \longrightarrow MgCl(aq) + H_2(g)}$$

Step 1: – Place a 1-in piece of Mg in each of three different test tubes. Label each test tube 1 to 3.

Step 2: – Measure 10mL of 1M HCl in a graduated cylinder and add it to test tube 1. Immediately, start recording the time and stop when all Mg has disappeared.

Step 3: – Repeat the previous step but now with 2M HCl and then 3M HCl. Write down the three different times in the table below.

3. Factors that affect the chemical rate. Presence of a catalyst. This mini-experiment addresses the impact of catalysts on the reaction rate of the decomposition of hydrogen peroxide:

$$\mathrm{2\,H_2O_2(aq) \longrightarrow O_2(g) + 2\,H_2O(l)}$$

You will add different possible catalysts into the reaction mixture and study where more oxygen bubbles are being produced.

Step 1: – Place 2mL of 3% H_2O_2 into each of five different test tubes. Label the test tubes from 1 to 5. Test tube 1 will be the reference test tube.

Step 2: – Add a small spatula tip of MnO_2 to test tube 2 and record your observations in comparison to test tube 1. If you see more bubbles than on test tube 1 that would mean the substance you used is a catalyst.

Step 3: – Repeat the previous step now using a set of possible catalysts in the table below. Record your observations. If you see more bubbles than on test tube 1 that would mean the substance you used is a catalyst.

4. Le Chatelier principle Reaction proceed from reactants to products but when products are former, reactions can also proceed from products to reactants. This establishes an equilibrium. When a reaction reached equilibrium, the forward and reverse reactions proceed at the same speed, so what is formed is also being consumed. You can alter a reaction in equilibrium pushing chemistry to the right of to the left so that mostly reactants or mostly products are being formed. You can do this by adding or removing reactants or by increasing or decreasing temperature. Le chatelier principle rationalizes the behavior of chemical reactions in equilibrium predicting the shift of the equilibrium. When reactants are added the reaction shifts to the right, when products are added the reaction differently shifts to the left. When reactants are removed, the reaction shifts to the left, and when products are removed it shifts to the right. In this mini-experiment you will address the impact of an equilibrium shift for the following reaction:

$$\underbrace{\mathrm{Fe^{3+}(aq)\ +\ SCN^{-}(aq)}}_{\text{yellow}} \rightleftharpoons \underbrace{\mathrm{FeSCN^{2+}(aq)}}_{\text{red}}$$

Step 1: – Measure 10mL of 0.01M $Fe(NO_3)_3$ and 10mL of 0.01M KSCN in a graduated cylinder. Pour both into a small beaker. Set up four test tubes in a rack add 3mL of previous mixture into each test tube. Label the test tubes from 1 to 4.

Step 2: – Test tube 1 will be the reference. Add 10 drops of water to this test tube.

Step 3: – Add 10 drops of 1M $Fe(NO_3)_3$–this is a product–to test tube 2. Record the color in comparison to test tube 1.

Step 4: – Add 10 drops of 1M KSCN–this is a product–to test tube 3. Record the color in comparison to test tube 1.

Step 5: – Add 10 drops of 1M HCl to test tube 4. This will remove Fe by forming $FeCl_4^-$. Record the color in comparison to test tube 1.

STUDENT INFO

Name: ________________________ Date: ______

Pre-lab Questions

Reaction Rates & Chemical Equilibrium

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1. Define speed of reaction.

2. Define forward reaction and reverse reaction.

3. For the following reaction, write down the forward reaction and the reverse reaction:

$$CO_2(g) + H_2O(l) \rightleftharpoons HCO_3{}^{-}{}_{(Aq)} + H^{+}(Aq)$$

STUDENT INFO	
Name: ____________________	Date: ______

Results
EXPERIMENT

Reaction Rates & Chemical Equilibrium

---

1. Factors that affect the chemical rate. Effect of temperature

Test tube (Hot/Cold)	Observation (intense/weak bubble formation)

2. Factors that affect the chemical rate. Effect of reactant concentration

Molarity of HCl	Total Time

3. Factors that affect the chemical rate. Presence of a catalyst.

Test tube	Observation (bubbles/no bubbles)	Catalyst (yes/no)?
Reference		
MnO_2		
Zn		
Fresh potato		
Boiled potato		

4. Le Chatelier principle

Test Tube (1,2,3,4)	Color	Color vs reference (Deeper or lighter)	Equilibrium shift ($\longrightarrow$ or $\longleftarrow$)
		N/A	

STUDENT INFO

Name: ______________________ Date: ______

Post-lab Questions

Reaction Rates & Chemical Equilibrium

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1. What is the impact of temperature on the reaction rate?

2. What is the impact of adding a catalyst on the reaction rate?

3. For the reaction below write down the expression of K_c

$$H_2CO_{3(aq)} \rightleftharpoons HCO_3{}^-{}_{(Aq)} + H^+_{(Aq)}$$

4. The chemical equilibrium that controls the PH of blood is

$$CO_{2(g)} + H_2O_{(l)} \rightleftharpoons HCO_3{}^-{}_{(Aq)} + H^+_{(Aq)}$$

   Respiratory alkalosis is caused by a lack of carbon dioxide in the blood that results from poor lung function or depressed breathing. When a patient has respiratory alkalosis, breathing from a paper bag can help. Based on the equilibrium, explain why this simple technique works.

5. In the miniexperiment 1, how did temperature affect the bubble production and why?

6. In the miniexperiment 2, how did molarity affect the time for Mg to dissapear and why?

7. In the miniexperiment 3, how do you explain the catalytic activity difference of fresh and boiled potatoes?

EXPERIMENT 0

Acid-Base Titration

A. Goal

The goal of this experiment is to calculate the *molar concentration* of a sample of acetic acid by means of a standard chemical procedure known as *titration*. In order to do that you will react the weak acid with a basic solution of sodium hydroxide (NaOH), which has a known concentration. You will also use phenolphthalein as an *indicator*.

B. Materials

- ☐ 5 mL glass-pipet
- ☐ 50 mL buret
- ☐ buret clamp and stand
- ☐ acetic acid s and NaOH
- ☐ phenolphthalein

C. Background

Titrations

A titration is a chemical technique used to calculate the unknown molarity of an acid or base. It is based on the principle that acids neutralize bases and we can figure out the molarity of the unknown chemical (the titrate) by knowing the reacting amounts. A titration uses chemical equipment: a burette, Erlenmeyer, and an indicator (see Figure 6). The unknown chemical is called the titrate and the known chemical is called the titrant. The goal of a titration is to calculate the volume of titrant needed to neutralize the titrate. We reach the endpoint of a titration when the titrant and titrate completely neutralize. At the end point, the mixture of titrant and titrate has a specific PH. Even though the chemical procedure in the lab is similar when titrating strong or weak acids or bases, the calculations needed to calculate the PH at the endpoint differ. This section will cover the principles and calculations involved in titrations.

Figure 6 An acid base titration using phenolphthalein as indicator. From left to right, before, at and after the end point.

Neutralization Reactions

Titrations involve a neuralization reaction in which an acid neutralizes a base. Acids produce protons H^+ and bases hydroxyls OH^- that neutralize each other forming water, H_2O. More importantly, they react in very specific ratios. Let us take a look at the reaction of hydrochloric acid with sodium hydroxide to produce water and sodium chloride:

$$HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)} \qquad \text{Neutralization Reaction}$$

In this reaction, one mole of HCl reacts with one mole of NaOH. The fact that one more reacts with one more can be used as a principle for acid-base titration. We will have to use the stoichiometry of the reaction to calculate the volume of titrant needed to neutralize the titrate. Imagine you have an unknown sample of HCl and you need to determine the amount of acid in the solution. If you know that this sample reacts with a specific amount of NaOH as you know that they react in a one-2-one ratio then you would know the acidic content. This is the idea behind titration: a laboratory procedure in which an unknown sample is neutralized with a known solution. A chemical *indicator*, which changes color depending on the acidity of the medium, is used to visually reveal the moment in which the acid and the base are completely neutralized. The point at which the indicator changes color is called the *endpoint*. At the endpoint, the acid and the base are neutralized.

Endpoint formula

At the *equivalence point*, also called the *stoichiometric point*, the moles of acid and the moles of the base are the same. A simple formula is extensively used to calculate the unknown acid concentration in a titration:

$$\boxed{n_H \cdot c_a \cdot V_a = n_{OH} \cdot c_b \cdot V_b} \tag{12}$$

where:

$n_H \cdot c_a \cdot V_a$ and $n_{OH} \cdot c_b \cdot V_b$ are moles of protons and hydroxyls, respectively

c_a and V_a are acid concentration and volume respectively

c_b and V_b are base concentration and volume respectively

n_H and n_{OH} are the number of protons of the acid and hydroxyls of the base

The units on this formula need to be consistent and for example, if the units of V_a are mL then the units of V_b should also be mL. This formula can be used when we titrate a known acid with a known base and we need to calculate the volume of titrant needed to reach the endpoint.

Equation 12 can also be used to identify if we have passed the endpoint in a titration. Imagine we titrate 2mL of 3M H_2SO_4 (titrant) with 2mL of 1M NaOH (titrate) and we want to know if we are before, after, or at the endpoint. Using Equation 12 we have:

$$2 \cdot 3M \cdot 2mL = 1 \cdot 1M \cdot V_b$$

That is, we would need 12 mL of the base. Therefore, as we only used 2mL we would be before the endpoint and we would have not reached the endpoint.

Sample Problem 31

A 50mL sample of an unknown acid is neutralized with 25 mL of a NaOH 3M solution. Calculate the molarity of the unknown acid.

SOLUTION

We will use Equation 12, given: $c_b = $ 3M, $V_b = 25mL$ and $V_a = 50mL$.

$$c_a \cdot 50mL = 3M \cdot 25mL$$

and the results is 1.5M.

STUDY CHECK

A 15mL sample of an unknown acid is neutralized with 45 mL of a NaOH 1M solution. Calculate the molarity of the unknown acid.

Answer: 3M

D. Procedure

1. Acetic acid titration

Step 1: – Obtain a 5 mL glass-pipet and a 50 mL buret with a stand and buret clamp.

Step 2: – Obtain about 30 mL of acetic acid solution (vinegar) in a 50 mL beaker and about 80 mL of the NaOH solution in a clean, dry Erlenmeyer flask. Keep the NaOH solution containing Erlenmeyer closed with a rubber stopper.

Step 3: – Clean your buret and fill it with the NaOH solution using a plastic funnel.

Step 4: – Record the initial volume in the buret as 0mL. Read accordingly to the tool precision, including your significant or estimated value.

Step 5: – Pipet 5.00 mL of acetic acid into a clean 125 mL Erlenmeyer flask that has 20 mL of distilled water and 2 drops of phenolphthalein.

Step 6: – Record the molarity of the NaOH solution (c_b) indicated in the lable of the stock solution bottle. This value will be the same for all experiments.

Step 7: – Place the flask under the buret. Use a piece of white paper under the flask to distinguish better the color change.

Step 8: – Add the NaOH solution from the buret in 1 mL portions, while swirling the solution in the flask.

Step 9: – The titration is completed when an addition of 1 mL causes the color to change from colorless to any shade of pink. Record the final buret volume.

Step 10: – Repeat the steps above four times and average the resulting acetic acid concentration.

Step 11: – Make sure you dispose of the solutions and leftovers in the corresponding disposals.

Calculations

1 Record the initial volume of the buret. This value is not necessarily 0.00 mL.

2 Record the final volume of the buret, after you reached the end point.

3 The volumen of NaOH used should be: 2 – 1

4 You can calculate the molarity of the acetic acid solution by means of:

$$c_a = \frac{3 \cdot c_b}{5 \text{ mL}}$$

where c_b is the given molarity of the NaOH solution found in the bottle.

5 Is the average of the 4 concentrations calculated.

$$\frac{\sum 4}{4}$$

STUDENT INFO

Name: ______________________ Date: ______

Pre-lab Questions

Acid-Base Titration

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1. A 10.00 mL sample of aqueous HCl requires 31.00 mL of 0.0900 M NaOH to reach the endpoint. What is the molar concentration of HCl. The equation for the reaction is:

$$HCl + NaOH \longrightarrow NaCl + H_2O$$

2. The molarity of a vinegar solution is 0.90 M. Calculate the number of acetic acid moles in 10. mL of this solution. Write down your result using scientific notation.

3. Nitric acid (HNO_3) is an acid with three protons. Suppose you titrate 5.00 mL of of this acid with NaOH 0.10 M. Knowing that the end point is reached after 25.00 mL of the base is added, find the molarity of the acid solution.

STUDENT INFO

Name: ____________________ Date: ______

Results

EXPERIMENT

Acid-Base Titration

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1. Acetic acid titration

		1	2	3	4
1	Initial Buret Volume (mL)				
2	Final Buret Volume (mL)				
3	NaOH Volume used (mL)				
4	CH_3COOH Concentration (M)				
5	Mean CH_3COOH Concentration (M)				

STUDENT INFO

Name: ______________________ Date: ______

Post-lab Questions

Acid-Base Titration

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1. You need to prepare a sample containing 0.20 g of $CuSO_4$ from a solution that is 10.% $CuSO_4$ by mass. What mass of solution do you need?

2. A 10.00 mL sample of aqueous HNO_3 requires 20.00 mL of 0.201 M NaOH to reach the endpoint. Calculate the molarity of HNO_3.

3. You titrate a vinegar sample–an acetic acid solution in water–with 0.30 M NaOH. Using 10. mL of vinegar, you reach the endpoint after adding 10. mL of the base. Indicate the molarity of the acetic acid solution.

EXPERIMENT 0

Organic Compounds: Alkanes

A. Goal

The goal of this laboratory experiment is to practice unit conversions and carry out calculations with the correct number of significant figures.

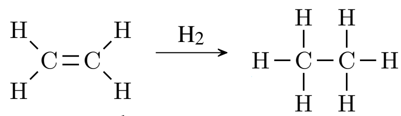
B. Materials

☐ This is theory-based experiment.

C. Background

Alkanes

This first section will introduce organic chemistry, covering the simplest organic compounds: alkanes. Alkanes–also called hydrocarbons–are simply made of carbon and hydrogen with all C-C bonds being single bonds. First, you will be introduced to a few organic chemicals, and you will learn about a series of different organic formulas that can represent the same compound. Then, you will learn the basic naming rule of alkanes, which extends to other–more complex–organic chemicals. Hydrocarbons composed of single C-C bonds are called saturated hydrocarbons, whereas hydrocarbons containing multiple C-C bonds are described as being unsaturated. All carbons in a saturated hydrocarbon are bound to four atoms. You can produce saturated hydrocarbons by reacting an unsaturated hydrocarbon with hydrogen:



unsaturated saturated

Molecular formula for alkanes

The naming of alkanes results from the combination of a prefix and a suffix. On one hand, the suffix is always *ane*. On the other hand, the prefix depends on the number of carbons in the molecule. Table 11 shows a list of the different prefixes. For example, the alkane with a single carbon is called methane (CH_4). Other examples of alkanes are ethane which contains two carbons (C_2H_6) or propane with three carbons (C_3H_8). The molecular formula for an alkane with n carbon atoms is:

$$C_nH_{2n+2} \tag{13}$$

Hence, we have that the molecular formula for methane ($n = 1$) is CH_4 and the molecular formula for octane ($n = 8$) is C_8H_{18}. *Molecular formulas* represent only the molecular compositions, showing only the elements in the molecule.

A review of the different structural formulas

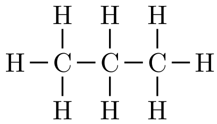
At this point, you have seen four different ways to represent organic molecules. Using propane as an example, here are all the formulas: We have that the *molecular formula* (e.g. C_3H_8 for propane) is mainly used to indicate the composition of the molecule in the form of Carbon and Hydrogen atoms. A second way to represent propane is using its *expanded* *structural formula*, that is by representing all atoms in the molecule and all atomic connections. A third molecular representation is the *condensed structural formula* that uses CH_3 and CH_2 units, only representing the C-C bonds. Finally, the *skeletal formula* is perhaps the most simplistic representation as only the C-C bonds are represented in the form of simple lines. It is important to understand that *all formulas are just different ways to represent the same molecule*.

C_3H_8

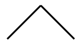
Molecular formula

$CH_3-CH_2-CH_3$

Condensed structural formula



Expanded structural formula



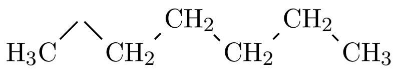
Skeletal structural formula

Sample Problem 32

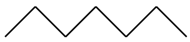
Write down the condensed and skeletal formulas for heptane.

SOLUTION

Heptane has seven carbons, hence its condensed formula will have two CH_3 units and five CH_2 units:



The skeletal formula for would be:

▶Answer: 

STUDY CHECK

Draw the skeletal formula of decane.

▶Answer: 

Table 11 Prefixed for alkane naming

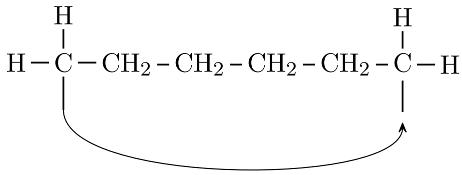
# Carbons	prefix	# Carbons	prefix
1	Meth	6	Hex
2	Eth	7	Hepta
3	Prop	8	Octa
4	But	9	Nona
5	Pent	10	Deca

Cycloalkanes

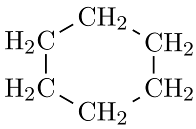
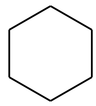
Alkanes are perfect examples of hydrocarbons, with C-C chains and all carbon atoms saturated with hydrogen. Cycloalkanes are simply alkanes with a cyclic structure. We will cover the molecular, condensed, and skeletal formulas for these chemicals.

Cyclic alkanes

Consider the expanded structure of hexane. A cycloalkane results from removing the left and right hydrogen while connecting the molecule in the form of a cycle:

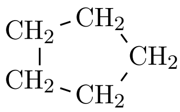
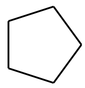


As the most stable structure for six lines is the hexagon, the resulting structure of cyclohexane would be:

 or  cyclohexane

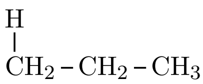
Naming cycloalkanes

The naming of alkanes and cycloalkanes is very similar. You just need to add the cyclo prefix to the name. For example, the alkane with five carbons is called pentane, whereas the corresponding cycloalkane is called cyclopentane:

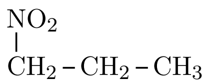
 or  Cyclopentane

Alkanes with substituents

Oftentimes alkanes have other groups of atoms called substituents attached to the hydrocarbon chain. This section covers the naming of alkanes with substituents. Here is an example of an alkane and an alkane with a substituent:



Propane



Nitropropane

In the substituted molecule, a nitro group has replaced a hydrogen atom.

Substituents

There are many different substituents–also called groups–that can be found attached to an alkane chain. Their names are indicated in Table **??**. The easiest substituents are halogens; atoms of chorine ($Cl-$), bromine ($Br-$) or iodine ($I-$) can replace hydrogen atoms in an alkane. The name of these substituents–chloro, bromo, and iodo– resembles the name of the corresponding atom. Other substituents can contain carbon, like a methyl (CH_3-) or a ethyl (CH_3CH_2-). There are even more complex substituents such as tert-butyl in which a central carbon atom is connected to three different methyl groups. The name of substituents (methyl) comes from the name of the alkane (methane) by replacing the *-ane* suffix with *-yl*.

Naming rules for branched alkanes

Overall, the rules to name branched alkanes are:

1. **Step one:** Look for the longest carbon-carbon chain that will give the ending name of the molecule (e.g. four carbons would be butane).
2. **Step two:** Number the main chain starting at the end closest to the substituents so that the numbers for the substituents are small.
3. **Step three:** Name the substituents with their position and order them alphabetically.

Sample Problem 33

Name the following hydrocarbon:

$$\begin{array}{c} \qquad\qquad\quad CH_3 \\ \qquad\qquad\quad | \\ CH_3-CH_2-C-CH_3 \\ \qquad\qquad\qquad\qquad\quad | \\ \qquad\qquad\qquad\qquad\quad CH_2-CH_2-CH_3 \end{array}$$

SOLUTION

First we locate the longest chain. We have five possible chains, and the longest one has six carbons. Hence the name of the hydrocarbon would be hexane. Now we need to number the carbons so that we start numbering the closes to the substituents the possible.

$$\begin{array}{lllll} & & CH_3 & & \\ & & | & & \\ \underset{1}{CH_3}- & \underset{2}{CH_2}- & C-CH_3 & & \\ & & |3 & & \\ & & \underset{4}{CH_2}- & \underset{5}{CH_2}- & \underset{6}{CH_3} \end{array}$$

We have two methyl connected to carbon number three. Hence the final name will be:

▶Answer: 3-dimethylhexane.

STUDY CHECK

Name the following hydrocarbon:

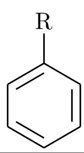
$$\begin{array}{lll} & & CH_2-CH_2-CH_2-CH_3 \\ & & | \\ CH_3-CH_2- & & C-CH_3 \\ & & | \\ & & CH_2-CH_2-CH_3 \end{array}$$

▶Answer: 4-ethyl-4-methyloctane.

Molecular diversity

You have certainly taken painkillers for a headache or over-the-counter drugs to get over a cold. Maybe you drink coffee and perhaps you like tea. All these substances contain active organic molecules. These active molecules are hydrocarbon derivatives and differ from plain hydrocarbons, which are simply made of carbon and hydrogen. Active molecules contain functional groups such as alcohol, ethers, carboxylic acids, amines, amides, or aromatic groups. These groups of atoms have a specific function and give activity to the molecule. The goal of this section is simply to identify the different groups.

Table 12 Names of several functional groups

Functional group	Name	Functional group	Name	Functional group	Name
$R_1R_2C=CR_3R_4$	Alkene	$R-\overset{\underset{\|\|}{O}}{C}-R'$	Ketone	$R-OH$	Alcohol
$R-C\equiv C-R'$	Alkyne	$R-\overset{\underset{\|\|}{O}}{C}-H$	Aldehyde	$R-SH$	Thiol
$R-\overset{\underset{\|\|}{O}}{C}-OH$	Carboxylic acid	$R-\overset{\underset{\|}{R'}}{N}-R''$	Amine	$R-O-R'$	Ether
$R-\overset{\underset{\|\|}{O}}{C}-O-R'$	Ester	$R-\overset{\underset{\|\|}{O}}{C}-\overset{\underset{\|}{R''}}{N}-R'$	Amide		Phenyl

STUDENT INFO

Name: ____________________ Date: ______

Pre-lab Questions

Organic Compounds: Alkanes

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1. Given the following molecular formula, name the following linear alkanes (hydrocarbons):

CH_4	____________	C_2H_6	____________
C_4H_{10}	____________	C_3H_8	____________
C_5H_{12}	____________	C_9H_{20}	____________

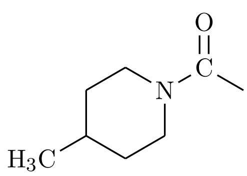
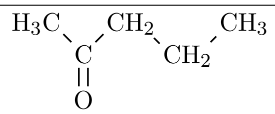
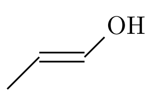
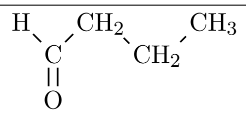
2. Given the following molecular formula, name the following cyclic alkanes (hydrocarbons):

C_3H_6	____________	C_6H_{12}	____________
C_4H_8	____________	C_7H_{14}	____________
C_5H_{10}	____________	C_9H_{18}	____________

3. Indicate the molecular, expanded, condensed and skeletal formula for the following linear alkanes:

	Molecular Formula	Expanded Formula	Condensed Formula	Skeletal Formula
Hexane				
Pentane				

4. Identify the functional groups:

Molecule	Functional Group	Molecule	Functional Group
			
			

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Name: ________________________	Date: ______

Results
EXPERIMENT

Organic Compounds: Alkanes

___ **Linear Alkanes** Use the molecular models set for this experiment. Each sphere represents an element. Carbon is black, hydrogen white, oxygen red and nitrogen blue. Build up the following molecules and complete the table. Show your professor all molecular models before proceeding to next part.

	Expanded Formula	Condensed Formula	Skeletal Formula
Methane			N/A
Ethane			
Propane			
Butane			

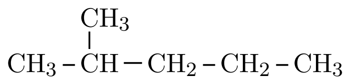
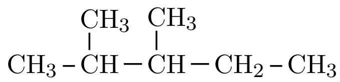
Cyclic Alkanes Use the molecular models set for this experiment. Each sphere represents an element. Carbon is black, hydrogen white, oxygen red and nitrogen blue. Build up the following molecules and complete the table. Show your professor all molecular models before proceeding to next part.

	Expanded Formula	Condensed Formula	Skeletal Formula
Cyclopropane			
Cyclobutane			
Cyclopentane			
Cyclohexane			

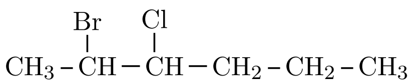
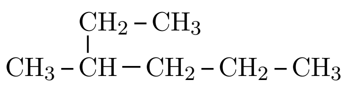
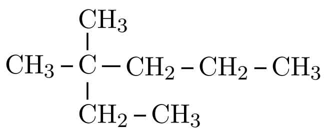
Short alkanes with substituents Use the molecular models set for this experiment. Each sphere represents an element. Carbon is black, hydrogen white, oxygen red and nitrogen blue. Build up the following molecules and complete the table. Show your professor all molecular models before proceeding to next part.

	Expanded Formula	Molecular Formula
Chloromethane		
Dichloromethane		
BromoChloro -Fluoromethane		
Chloroethane		

Long alkanes with substituents Use the molecular models set for this experiment. Each sphere represents an element. Carbon is black, hydrogen white, oxygen red and nitrogen blue. Build up the following molecules and complete the table. Show your professor all molecular models before proceeding to next part.

Name	Condensed Formula	Skeletal Formula
		
		

More alkanes with substituents There is no need to use the molecular models at this point. Now, name the following molecules:

Formula	name
	
	
	

Functional Groups Identity the following functional groups:

Formula	name