CK-12 Chemistry Workbook

CK-12 Foundation

February 5, 2010

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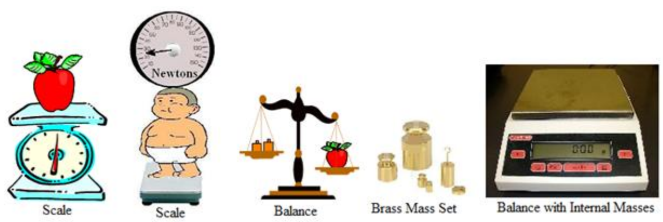
Mass Versus Weight Worksheet

CK-12 Foundation Chemistry

Name____ **Date**________________

The mass of an object is a measure of the amount of matter in it. The mass (amount of matter) of an object remains the same regardless of where the object is placed. For example, moving a brick to the moon does not cause any matter in it to disappear or be removed. The weight of an object is the force of attraction between the object and the earth (or whatever large body it is resting on). We call this force of attraction, **the force of gravity**. The gravitational pull on the object varies depending on where the object is with respect to the earth or other gravity producing object. For example, a man who weighs 180 pounds on earth would weigh only 45 pounds if he were in a stationary position, 4,000 miles above the earth's surface. This same man would weigh only 30 pounds on the moon because the moon's gravity is only one-sixth that of earth. If this man were in outer space with no planet or moon nearby, **his weight would be zero**. There would be gravitational pull on him at all. The mass of this man, however, would be the same in all those situations because the amount of matter in him is constant.

We measure ***weight*** with a ***scale***, which is a spring that compresses when a weight is placed on it. If the gravitational pull is less, the spring compresses less and the scale shows less weight. We measure ***mass*** with a ***balance***. A balance compares the unknown mass to known masses by balancing them on a lever. If we take our balance and known masses to the moon, an object will have the same measured mass that it had on the earth. The weight, of course, would be different on the moon.



On, or near the surface of, the earth, the force of gravity is constant and so we can determine either the mass or the weight of an object if we know one of those two. On or near the surface of the earth, the conversion factor between mass and weight is: 1.00 kg of mass will have a weight of 9.80 Newtons (the standard unit of force in the SI system).

Example: What is the weight in Newtons of a 3.0 kg mass on the surface of the earth?

(gravitational force = (3.00 kg)(9.80 N/kg) = 29.4 N)

Example: If an object weighs 200. N on the surface of the earth, what is its mass?

$$\text{mass} = (200.\ N)\left(\frac{1.00\ kg}{9.80\ N}\right) = 20.4$$

Exercises

1. If an object weighs 400. N on the earth, how much mass does it contain?
2. What is the weight, in Newtons, of a 50 kg mass on the surface of the earth?
3. On the surface of the earth, how much mass is contained in a 600. N weight?
4. If an object weighs 1200 N on the earth, how much will it weigh on the moon?
5. If an object has a mass of 120 kg on the earth, what is its mass on the moon?

1.5 Lesson 1.5 Energy

There are no worksheets for this lesson.

Chapter 2

Chemistry - A Physical Science Worksheets

2.1 Lesson 2.1 Measurements in Chemistry

There are no worksheets for this lesson.

2.2 Lesson 2.2 Using Measurements

Significant Figures Worksheet

Name____ **Date**______________

Working in the field of science almost always involves working with numbers. Some observations in science are qualitative and therefore, do not involve numbers, but in chemistry, most observations are quantitative and so, require numbers. You have been working with numbers for many years in your math classes thus numbers are not new to you. Unfortunately, there are some differences between the numbers you use in math and the numbers you use in science.

The numbers you use in math class are considered to be exact numbers. When you are given the number ***2*** in a math problem, it does not mean ***1.999*** rounded to ***2*** nor does it mean ***2.000001*** rounded to ***2***. In math class, the number ***2*** means exactly ***2.00000000...*** with an infinite number of zeros - ***a perfect 2!*** Such numbers are produced only by definition, not by measurement. That is, we can define 1 foot to contain exactly 12 inches, and these two numbers are perfect numbers, but we cannot measure an object to be exactly 12 inches long. In the case of measurements, we can read our measuring instruments only to a limited number of subdivisions. We are limited by our ability to see smaller and smaller subdivisions, and we are limited by our ability to construct smaller and smaller subdivisions. Even using powerful microscopes to construct and read our measuring devices, we eventually reach a limit, and therefore, even though the actual measurement of an object may be a perfect number of inches, we cannot prove it to be so. ***Measurements do not produce perfect numbers*** and since science is greatly involved with measuring, science does not produce perfect numbers (except in defined numbers such as conversion factors).

It is very important to recognize and report the limitations of measurements along with the magnitude and unit of the measurement. Many times, the analysis of the measurements made in a science experiment is simply the search for regularity in the observations. If the numbers reported show the limits of the measurements, the regularity, or lack thereof, becomes visible.

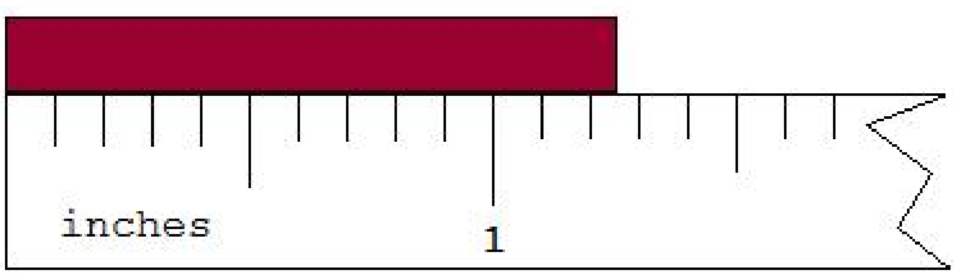
Table 2.1: **Two Sets of Observations**

Observations List A	**Observations List B**
22.41359 m	22.4 m
22.37899 m	22.4 m
22.42333 m	22.4 m
22.39414 m	22.4 m

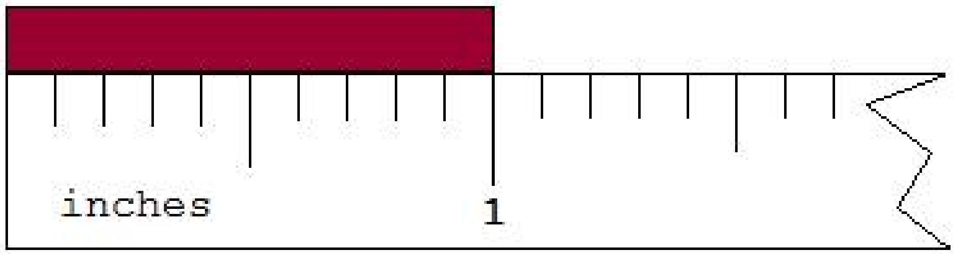
In the lists of observations above, it is difficult to perceive a regularity in List A, but when the numbers are reported showing the limits of the measurements as in List B, the regularity becomes apparent.

One of the methods used to keep track of the limit of a measurement is called ***Significant Figures***. In this system, when you record a measurement, the written number must indicate the limit of the measurement, and when you perform mathematical operations on measurements, the final answer must also indicate the limit of the original measurements.

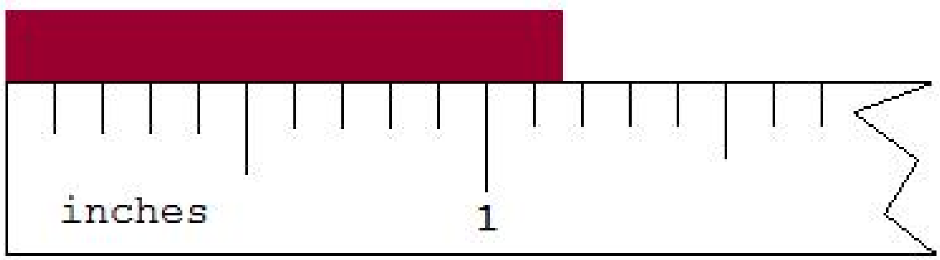
To record a measurement, you must write down all the digits actually measured, including measurements of zero and you must NOT write down any digit not measured. The only real problem that occurs with this system is that zeros are sometimes used as measured numbers and are sometimes used simply to locate the decimal point and ARE NOT measured numbers.



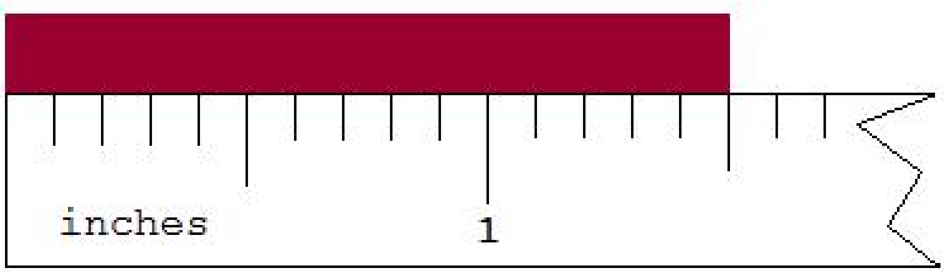
In the case shown above, the correct measurement is greater than 1.3 inches but less than 1.4 inches. ***It is proper to estimate one place beyond the calibrations*** of the measuring instrument. Therefore, this measurement should be reported as either 1.33, 1.34, 1.35, 1.36, or 1.37 inches.



In this second case, it is apparent that the object is, as nearly as we can read, exactly at 1 inch. Since we know the tenths place is zero and can estimate the hundredths place to be zero, the measurement should be reported as 1.00 inch. It is vital that you include the zeros in your measurement report because these are measured places.



This is read as 1.13, 1.14, 1.15, or 1.16 inches.



This is read 1.50 inches.

These readings indicate that the measuring instrument had subdivisions down to the tenths place and the hundredths place is estimated. There is some uncertainty about the last and only the last digit.

In our system of writing significant figures, we must distinguish between measured zeros and place-holding zeros. Here are the rules for determining the number of significant figures in a measurement.

RULES FOR DETERMINING THE NUMBER OF SIGNIFICANT FIGURES

1. All non-zero digits are significant.

2. All zeros between non-zero digits are significant.

3. All beginning zeros are NOT significant.

4. Ending zeros are significant if the decimal point is actually written in but not significant if the decimal point is an understood decimal.

Examples of the Rules

1. All non-zero digits are significant.

543 has 3 significant figures.
22.437 has 5 significant figures.
1.321754 has 7 significant figures.

2. All zeros between non-zero digits are significant.

7,004 has 4 significant figures.
10.3002 has 6 significant figures.
103.0406 has 7 significant figures.

3. All beginning zeros are NOT significant.

00013.25 has 4 significant figures.
0.0000075 has 2 significant figures.
0.000002 has 1 significant figure.

4. Ending zeros are significant if the decimal point is actually written in but not significant if the decimal point is an understood decimal.

37.300 has 5 significant figures.
33.00000 has 7 significant figures.
1.70 has 3 significant figures.
1,000,000 has 1 significant figure.
302,000 has 3 significant figures.
1,050 has 3 significant figures.

1,000,000. has 7 significant figures.
302,000. has 6 significant figures.
1,050. has 4 significant figures.

Exercises

How many significant figures are given in each of the following measurements?

1. 454 g ________
2. 2.2 lbs ________
3. 2.205 lbs ________
4. 0.3937 L ________
5. 0.0353 L ________
6. 1.00800 g ________
7. 500 g ________
8. 480 ft ________
9. 0.0350 kg ________
10. 100. cm ________
11. 1,000 m ________
12. 0.625 L ________
13. 63.4540 mm ________
14. 3,060 m ________
15. 500. g ________
16. 14.0 mL ________
17. 1030 g _________
18. 9,700 g ________
19. 125,000 m ________
20. 12,030.7210 g ________
21. 0.0000000030 cm ________
22. 0.002 m ________
23. 0.0300 cm ________
24. 1.00 L ________

25. 0.025 m/s ________

26. 0.100 kg ________

27. 0.00300 km ________

28. 303.0 g ________

29. 250 g ________

30. 1,000. m ________

Maintaining Significant Figures Through Mathematical Operations

In addition to using significant figures to report measurements, we also use them to report the results of computations made with measurements. The results of mathematical operations with measurements must include an indication of the number of significant figures in the original measurements. There are two rules for determining the number of significant figures after a mathematical operation. One rule is for addition and subtraction, and the other rule is for multiplication and division. (Most of the errors that occur in this area result from using the wrong rule, so always double check that you are using the correct rule for the mathematical operation involved.

Significant Figure Rule for Addition and Subtraction

The answer for an addition or subtraction problem must have digits no further to the right than the shortest addend.

Example:

```
  13.3843 cm
   1.012  cm
+  3.22   cm
-------------
  17.6163 cm = 17.62 cm
```

Note that the vertical column farthest to the right has a **3** in the top number but that this column has blank spaces in the next two numbers in the column. In elementary math classes, you were taught that these blank spaces can be filled in with zeros, and in such a case, the answer would be 17.6163 cm. In science, however, these blank spaces are NOT zeros but are unknown numbers. Since they are unknown numbers, you cannot substitute any numbers into the blank spaces and you cannot claim to know, for sure, the result of adding that column. You can know the sum of adding (or subtracting) any column of numbers that contains an unknown number. Therefore, when you add these three columns of numbers, the only columns for which you are sure of the sum are the columns that have a known number in each space in the column. When you have finished adding these three numbers in the normal mathematical process, you must round off all those columns that contain an unknown number (a blank space). Therefore, the correct answer for this addition is 17.62 cm and has four significant figures.

Example:

$$\begin{array}{lll} & 12 & m \\ + & 0.00045 & m \\ \hline & 12.00045 & m = 12\ m \end{array}$$

In this case, the **12** has no numbers beyond the decimal and therefore, all those columns must be rounded off and we have the seemingly odd result that after adding a number to **12**, the answer is still **12**. This is a common occurrence in science and is absolutely correct.

Example:

$$\begin{array}{lll} & 56.8885 & cm \\ & 8.30 & cm \\ + & 47.0 & cm \\ \hline & 112.1885 & cm = 112.2\ cm \end{array}$$

This answer must be rounded back to the tenths place because that is the last place where all the added numbers have a recorded digit.

Significant Figure Rule for Multiplication and Division

The answer for a multiplication or division operation must have the same number of significant figures as the factor with the least number of significant figures.

Example: $(3.556\ cm)(2.4\ cm)\ =\ 8.5344\ cm^2\ =\ 8.5\ cm^2$

In this case, the factor 2.4 has two significant figures and therefore, the answer must have two significant figures. The mathematical answer is rounded back to two significant figures.

Example: $(20.0\ cm)(5.0000\ cm)\ =\ 100\ cm^2\ =\ 100.\ cm^2$

In this example, the factor 20.0 cm has three significant figures and therefore, the answer must have three significant figures. In order for this answer to have three significant figures, we place an actual decimal after the second zero to indicate three significant figures.

Example: $(5.444\ cm)(22\ cm)\ =\ 119.768 cm^2\ =\ 120\ cm^2$

In this example, the factor 22 cm has two significant figures and therefore, the answer must have two significant figures. The mathematical answer is rounded back to two significant figures. In order to keep the decimal in the correct position, a non-significant zero is used.

Exercises

Add, subtract, multiply, or divide as indicated and report your answer with the proper number of significant figures.

1.
$$\begin{array}{rr} & 703\ \text{g} \\ & 7\ \text{g} \\ + & 0.66\ \text{g} \\ \hline \end{array}$$

2.
$$\begin{array}{rr} & 5.624\ \text{ft} \\ & 0.24\ \text{ft} \\ + & 16.8\ \text{ft} \\ \hline \end{array}$$

3.
$$\begin{array}{rr} & 34\ \text{kg} \\ - & 0.2\ \text{kg} \\ \hline \end{array}$$

4.
$$\begin{array}{rr} & 18.7\ \text{m} \\ + & 0.009\ \text{m} \\ \hline \end{array}$$

5. Add 65.23 cm, 2.666 cm, and 10 cm.
6. Multiply 2.21 cm and 0.3 cm.
7. Multiply: (2.002 cm)(84 cm)
8. Multiply: (107.888 cm)(0.060 cm)
9. Divide 72.4 cm by 0.0000082 cm.
10. Multiply 0.32 cm by 600 cm and then divide the product by 8.21 cm.

Exponential Notation Worksheet

Name____ **Date**______________

Work in science frequently involves very large and very small numbers. The speed of light, for example, is 300,000,000 meters/second; the mass of the earth is 6,000,000,000,000,000,000,000,000 kg; and the mass of an electron is 0.000000000000000000000000000009 kg. It is very inconvenient to write such numbers and even more inconvenient to attempt to carry out mathematical operations with them. Imagine trying to divide the mass of the earth by the mass of an electron! Scientists and mathematicians have designed an easier method for dealing with such numbers. This more convenient system is called **Exponential Notation** by mathematicians and **Scientific Notation** by scientists.

In scientific notation, very large and very small numbers are expressed as the product of a number between 1 and 10 and some power of 10. The number 9,000,000, for example, can be written as the product of 9 times 1,000,000 and 1,000,000 can be written as 10^6. Therefore, 9,000,000 can be written as $9\ x\ 10^6$. In a similar manner, 0.00000004 can be written as 4 times $\frac{1}{10^8}$ or $4\ x\ 10^{-8}$.

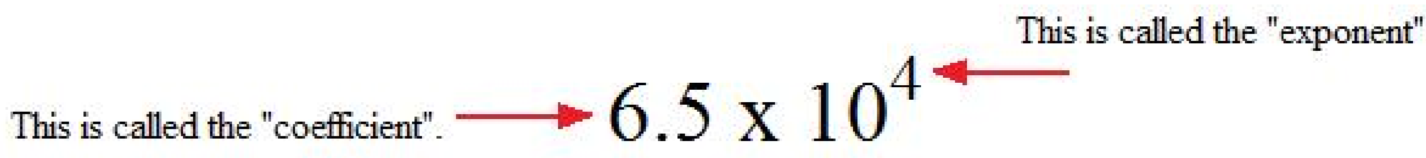


Table 2.2: **Examples**

Decimal Notation	*Scientific Notation*
95,672	$9.5672\ x\ 10^4$
8,340	$8.34\ x\ 10^3$
100	$1\ x\ 10^2$
7.21	$7.21\ x\ 10^0$
0.014	$1.4\ x\ 10^{-2}$
0.0000000080	$8.0\ x\ 10^{-9}$
0.00000000000975	$9.75\ x\ 10^{-12}$

As you can see from the examples above, to convert a number from decimal to exponential form, you count the spaces that you need to move the decimal and that number becomes the exponent of 10. If you are moving the decimal to the left, the exponent is positive, and if you are moving the decimal to the right, the exponent is negative. One and only one non-zero digit exists to the left of the decimal and ALL significant figures are maintained. The value of using exponential notation occurs when there are many non-significant zeros.

Exercises

Express the following decimal numbers in exponential form. The exponential form should have exactly one non-zero digit to the left of the decimal and you must carry all significant figures.

1. 1000
2. 150,000
3. 243
4. 9.3
5. 435,000,000,000
6. 0.0035

7. 0.012567
8. 0.0000000000100
9. 0.000000000000467
10. 0.000200
11. 186,000
12. 9,000,000,000,000
13. 105
14. 77,000
15. 502,000

Carrying Out Mathematical Operations with Exponential Numbers

When numbers in exponential notation are added or subtracted, the exponents ***must*** be the same. If the exponents are the same, the coefficients are added and the exponent remains the same.

Consider the following example.

$4.3\ x\ 10^4\ +\ 1.5\ x\ 10^4\ =\ (4.3\ +\ 1.5)\ x\ 10^4\ =\ 5.8\ x\ 10^4 \qquad (43,000\ +\ 15,000\ =\ 58,000)$

$8.6\ x\ 10^7 - 5.3\ x\ 10^7\ =\ (8.6 - 5.3)\ x\ 10^7\ =\ 3.3\ x\ 10^7 \qquad (86,000,000 - 53,000,000\ =\ 33,000,000)$

If the exponents of the numbers to be added or subtracted are not the same, then one of the numbers must be changed so that the two numbers have the same exponent.

Examples

The two numbers given below, in their present form, cannot be added because they do not have the same exponent. We will change one of the numbers so that it has the same exponent as the other number. In this case, we choose to change $3.0\ x\ 10^4$ to $0.30\ x\ 10^5$. This change is made by moving the decimal one place to the left and increasing the exponent by 1. The two numbers can now be added.

$8.6\ x\ 10^5\ +\ 3.0\ x\ 10^4\ =\ 8.6\ x\ 10^5\ +\ 0.30\ x\ 10^5\ =\ 8.9\ x\ 10^5$

We also could have chosen to alter the other number. Instead of changing the second number to a higher exponent, we could have changed the first number to a lower exponent.

$8.6\ x\ 10^5\ \rightarrow\ 86\ x\ 10^4$

Now, we can add the numbers, $86\ x\ 10^4\ +\ 3.0\ x\ 10^4\ =\ 89\ x\ 10^4$

The answer, in this case, is not in proper exponential form because it has two non-zero digits to the left of the decimal. When we convert the answer to proper exponential form, it is exactly the same answer as before, $89\ x\ 10^4\ \rightarrow\ 8.9\ x\ 10^5$.

Exercises

Add or subtract the following exponential numbers as indicated.

1. $(8.34\ x\ 10^5)\ +\ (1.22\ x\ 10^5)\ =$
2. $(4.88\ x\ 10^3)\ -\ (1.22\ x\ 10^3) =$
3. $(5.6\ x\ 10^{-4})\ +\ (1.2\ x\ 10^{-4})\ =$
4. $(6.38\ x\ 10^5)\ +\ (1.2\ x\ 10^4)\ =$
5. $(8.34\ x\ 10^5)\ -\ (1.2\ x\ 10^4)\ =$
6. $(8.34\ x\ 10^{-5})\ +\ (1.2\ x\ 10^{-6})\ =$
7. $(4.93\ x\ 10^{-1})\ -\ (1.2\ x\ 10^{-2})\ =$
8. $(1.66\ x\ 10^{-5})\ +\ (6.4\ x\ 10^{-6})\ =$
9. $(6.34\ x\ 10^{15})\ +\ (1.2\ x\ 10^{16}) =$
10. $(6.34\ x\ 10^{15})\ -\ (1.2\ x\ 10^1)\ =$

Multiplying or Dividing with Numbers in Exponential Form

When multiplying or dividing numbers in scientific notation, the numbers do not have to have the same exponents. To multiply exponential numbers, ***multiply the coefficients*** and ***add the exponents***. To divide exponential numbers, ***divide the coefficients*** and ***subtract the exponents***.

Examples of Multiplying Exponential Numbers

Multiply: $(4.2\ x\ 10^4)(2.2\ x\ 10^2)\ =\ (4.2\ x\ 2.2)(10^{4+2})\ =\ 9.2\ x\ 10^6$

The coefficient of the answer comes out to be 9.24 but since we can only carry two significant figures in the answer, it has been rounded to 9.2 .

Multiply: $(2\ x\ 10^9)(4\ x\ 10^{14}\ =\ (2\ x\ 4)(10^{9+14})\ =\ 8\ x\ 10^{23}$

Multiply: $(2\ x\ 10^{-9})(4\ x\ 10^4)\ =\ (2\ x\ 4)(10^{-9+4})\ =\ 8\ x\ 10^{-5}$

Multiply: $(2\ x\ 10^{-5})(4\ x\ 10^{-4})\ =\ (2\ x\ 4)(10^{-5-4})\ =\ 8\ x\ 10^{-9}$

Multiply: $(8.2\ x\ 10^{-9})(8.2\ x\ 10^{-4})\ =\ (8.2\ x\ 8.2)(10^{(-9)+(-4)})\ =\ 32.8\ x\ 10^{-13}$

The product in the last example has too many significant figures and is not in proper exponential form, so we must round to two significant figures, $33\ x\ 10^{-13}$, and then move the decimal and correct the exponent, $3.3\ x\ 10^{-12}$.

Examples of Dividing Exponential Numbers

Divide: $\frac{8\ x\ 10^7}{2\ x\ 10^4}\ =\ (\frac{8}{2})(10^{7-4})\ =\ 4\ x\ 10^3$

Divide: $\frac{8\ x\ 10^{-7}}{2\ x\ 10^{-4}}\ =\ (\frac{8}{2})(10^{(-7)-(-4)})\ =\ 4\ x\ 10^{-3}$

Divide: $\frac{4.6\ x\ 10^{3}}{2.3\ x\ 10^{-4}} = (\frac{4.6}{2.3})(10^{(3)-(-4)}) = 2.0\ x\ 10^{7}$

In the final example, since the original coefficients had two significant figures, the answer must have two significant figures and therefore, the zero in the tenths place is carried.

Exercises

1. Multiply: $(2.0\ x\ 10^{7})(2.0\ x\ 10^{7}) =$
2. Multiply: $(5.0\ x\ 10^{7})(4.0\ x\ 10^{7}) =$
3. Multiply: $(4.0\ x\ 10^{-3})(1.2\ x\ 10^{-2}) =$
4. Multiply: $(4\ x\ 10^{-11})(5\ x\ 10^{2}) =$
5. Multiply: $(1.53\ x\ 10^{3})(4.200\ x\ 10^{5}) =$
6. Multiply: $(2\ x\ 10^{-13})(3.00\ x\ 10^{-22}) =$
7. Divide: $\frac{4.0\ x\ 10^{5}}{2.0\ x\ 10^{5}} =$
8. Divide: $\frac{6.2\ x\ 10^{15}}{2.0\ x\ 10^{5}} =$
9. Divide: $\frac{8.6\ x\ 10^{-5}}{3.1\ x\ 10^{3}} =$
10. Divide: $\frac{8.6\ x\ 10^{-5}}{3.1\ x\ 10^{-11}} =$

2.3 Lesson 2.3 Using Mathematics in Chemistry

Measurements Worksheet

Name____ **Date**__________

Measurement makes it possible to obtain more exact observations about the properties of matter such as the size, shape, mass, temperature, or composition. It allows us to make more exact quantitative observations. For example, the balance makes it possible to determine the mass of an object more accurately than we could by lifting the object and a clock gives a better measure of time than we could determine by observing the sun's position in the sky.

Measurements were orginally made by comparing the object being measured to some familiar object. Length was compared to the length of one's foot. Other measures were handspans, elbow to fingertip, and so on. As people's needs increased for more consistent measurements, STANDARD systems of measurement were devised. In a standard system of measurement, some length is chosen to be the standard and copies of this object can then be used by everyone making measurements. With a standard system of measurement, two people measuring the same distance will get the same measurement.

For a time, the standard for length (one meter) was a platinum bar which was marked and stored at constant temperature in a vault. It was stored at constant temperature so that it did not expand or contract. Standard masses are also stored in airtight containers to insure no change due to oxidation. Presently, the standard ***meter*** is the distance light travels in a vacuum in $\frac{1}{299,792,458}$ second and the standard ***second*** is based on the vibrations of a cesium-133 atom.

For any system of measurements, all measurements must include a **unit** term; a word following the number that indicates the standard the measurement is based on. Systems of measurement have several standards such as length, mass, and time, and are based on physical objects such as platinum bars or vibrating atoms. Standards based on physical objects are called **undefined units**. All the other standards are expressed in terms of these object-based standards. For example, length and time are object-based standards and velocity (meters/second) and acceleration (m/s_2) are expressed in terms of length and time. Volume is expressed in terms of the length standard, volume = length x length x length, such as cm^3.

There are two major systems of standards used in the United States. The one commonly used by the public (pounds, feet) and the system used for all scientific and technical work (kilograms, meters). The system used for scientific work is called the **Metric System** in its short form and is called the **International System (SI)** in its complete form. The undefined units in the SI system are the meter, gram, and second. All the sub-divisions in the SI system are in decimal form.

Conversion Factors, English to Metric

1.00 inch = 2.54 centimeters

1.00 quart = 0.946 liter

1.00 pound = 4.54 Newtons (= 454 grams on earth)

Units and Sub-Divisions for the SI System

Basic unit for length = meter

Basic unit for mass = gram

Basic unit for time = second

Unit for volume = liter (lee-ter)

1000 millimeters = 1 meter

100 centimeters = 1 meter

1000 meters = 1 kilometer

10 centimeters = 1 millimeter

1000 milligrams = 1 gram

1000 grams = 1 kilogram

1000 milliliters = 1 liter

1 milliliters = 1 cubic centimeter = 1 cm^3

All the relationships between units are defined numbers and therefore, have an infinite number of significant figures. When converting units, the significant figures of the answer are based on the significant figures of the measurement, not on the conversion factors.

The unit terms for measurements are an integral part of the measurement expression and must be carried through every mathematical operation that the numbers go through. In performing mathematical operations on measurements, the unit terms as well as the numbers obey the algebraic laws of exponents and cancellation.

Examples:

Table 2.3: (continued)

Math Operations	**Unit Term Operations**

Table 2.3: **Unit Terms Follow the Rules of Algebra**

Math Operations	**Unit Term Operations**
$6x + 2x = 8x$	$6\ mL + 2\ mL = 8\ mL$
$(5x)(3x) = 15x^2$	$(5\ cm)(3\ cm) = 15\ cm^2$
$\frac{9x^3}{3x} = 3x^2$	$\frac{9\ cm^3}{3\ cm} = 3\ cm^2$
$\frac{21x}{3a} = 7\ (\frac{x}{a})$	$\frac{21\ grams}{3\ cm^3} = 7\ \frac{grams}{cm^3}$

Converting Units

Frequently, it is necessary to convert units measuring the same quantity from one form to another. For example, it may be necessary to convert a length measurement in meters to millimeters. This process is quite simple if you follow a standard procedure called *unit analysis*. This procedure involves creating a **conversion factor** from equivalencies between various units.

For example, we know that there are 12 inches in 1 foot. Therefore, the conversion factor between inches and feet is 12 inches = 1 foot. If we have a measurement in inches and we wish to convert the measurement to feet, we would generate a conversion factor $(\frac{1\ foot}{12\ inches})$ and multiply the measurement by this conversion factor.

Example: Convert 500. inches to feet.

$$(500.\ \text{inches})(\frac{1\ foot}{12\ inches}) = 41.7\ \text{feet}$$

We design the conversion factor specifically for this problem so that the unit term "inches" will cancel out and the final answer will have the unit "feet". This is how we know to put the unit term "inches" in the denominator and the unit term "foot" in the numerator.

Example: Convert 6.4 nobs to hics given the conversion factor, 5 hics = 1 nob.

$$(6.4\ \text{nobs})(\frac{5\ hics}{1\ nob}) = 32\ \text{hics}$$

Example: Convert 4.5 whees to dats given the conversion factor, 10 whees = 1 dat.

$$(4.5\ \text{whees})(\frac{1\ dat}{10\ whees}) = 0.45\ \text{dats}$$

Sometimes, it is necessary to insert a series of conversion factors.

Example: Convert 5.00 wags to pix given the conversion factors, 10 wags = 1 hat, and 1 hat = 2 pix.

$$(5.00 \text{ wags})(\frac{1\ hat}{10\ wags})(\frac{2\ pix}{1\ hat}) = 1.00 \text{ pix}$$

Solved Conversion Problems

1. Convert 1.22 cm to mm.

$$(1.22 \text{ cm})(\frac{10\ mm}{1\ cm}) = 12.2 \text{ mm}$$

2. Convert 5.00 inches to mm.

$$(5.00 \text{ inches})(\frac{2.54\ cm}{1\ inch})(\frac{10\ mm}{1\ cm}) = 127 \text{ mm}$$

3. Convert 66 lbs to kg. As long as the object is at the surface of the earth, pounds (force) can be converted to grams (mass) with the conversion factor 454 g = 1 lb.

$$(66 \text{ lbs})(\frac{454\ g}{1\ lb})(\frac{1\ kg}{1000\ g}) = 30. \text{ kg}$$

The mathematical answer for this conversion comes out to be 29.964 but must be rounded off to two significant figures since the original measurement has only two significant figures. When 29.964 is rounded to two significant figures, it requires a written in decimal after the zero to make the zero significant. Therefore, the final answer is 30. kg.

4. Convert 340. mg/cm^3 to lbs/ft^3.

$$(\frac{340.\ mg}{1\ cm^3})(\frac{1\ g}{1000\ mg})(\frac{1\ lb}{454\ g})(\frac{16.39\ cm^3}{1\ in^3})(\frac{17.28\ in^3}{1\ ft^3}) = 21.2\ lbs/ft^3$$

You should examine the units yourself to make sure they cancel and leave the correct units for the answer.

Exercises

1. Convert 40. cots to togs given the conversion factor, 10 cots = 1 tog.
2. Convert 8.0 curs to nibbles given the conversion factor, 1 cur = 10 nibbles.
3. Convert 100. gags to bobos given the conversion factor, 5 gags = 1 bobo.
4. Convert 1.0 rat to utes given the conversion factors, 10 rats = 1 gob and 10 gobs = 1 ute.
5. Express 3.69 m in cm.

6. Express 140 mm in cm.

7. Convert 15 inches to mm.

8. Express 32.0 grams in pounds. (Be aware that such a conversion between weight and mass is only reasonable on the surface of the earth.)

9. Express 690 mm in m.

10. Convert 32.0 lbs/qt to g/mL.

11. Convert 240. mm to cm.

12. Convert 14,000 mm to m.

2.4 Lesson 2.4 Using Algebra in Chemistry

There are no worksheets for this lesson.

Chapter 3

Chemistry in the Laboratory Worksheets

3.1 Lesson 3.1 Making Observations

There are no worksheets for this lesson.

3.2 Lesson 3.2 Making Measurements

There are no worksheets for this lesson.

3.3 Lesson 3.3 Using Data

There are no worksheets for this lesson.

3.4 Lesson 3.4 How Scientists Use Data

There are no worksheets for this lesson.

Chapter 4

The Atomic Theory Worksheets

4.1 Lesson 4.1 Early Development of a Theory

There are no worksheets for this lesson.

4.2 Lesson 4.2 Further Understanding of the Atom

There are no worksheets for this lesson.

4.3 Lesson 4.3 Atomic Terminology

There are no worksheets for this lesson.

Chapter 5

The Bohr Model of the Atom Worksheets

5.1 Lesson 5.1 The Wave Form of Light

There are no worksheets for this lesson.

5.2 Lesson 5.2 The Dual Nature of Light

There are no worksheets for this lesson.

5.3 Lesson 5.3 Light and the Atomic Spectra

There are no worksheets for this lesson.

5.4 Lesson 5.4 The Bohr Model

There are no worksheets for this lesson.

Chapter 6

Quantum Mechanics Model of the Atom Worksheets

6.1 Lesson 6.1 The Wave-Particle Duality

There are no worksheets for this lesson.

6.2 Lesson 6.2 Schrodinger's Wave Functions

There are no worksheets for this lesson.

6.3 Lesson 6.3 Heisenberg's Contribution

There are no worksheets for this lesson.

6.4 Lesson 6.4 Quantum Numbers

There are no worksheets for this lesson.

6.5 Lesson 6.5 Shapes of Atomic Orbitals

Quantum Numbers and Orbital Shapes Worksheet

CK-12 Foundation Chemistry . . . . Name___________________________________
Date______________

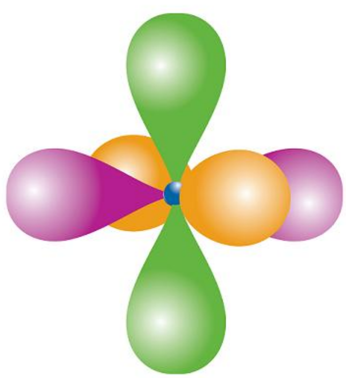
Energy Level	Sub-Level	Number of Sub-Level Orbitals	Maximum Number of Electrons in Sub-Energy Level	Quantum Numbers for this Sub-Level
1	s	1	2	n = 1, l = 0
2	s	1	2	n = 2, l = 0
2	p	3	6	n = 2, l = 1
3	s	1	2	n = 3, l = 0
3	p	3	6	n = 3, l = 1
3	d	5	10	n = 3, l = 2
4	s	1	2	n = 4, l = 0
4	p	3	6	n = 4, l = 1
4	d	5	10	n = 4, l = 2
4	f	7	14	n = 4, l = 3
5	s	1	2	n = 5, l = 0
5	p	3	6	n = 5, l = 1
5	d	5	10	n = 5, l = 2
5	f	7	14	n = 5, l = 3
5	g	9	18	n = 5, l = 4
6	s	1	2	n = 6, l = 0
6	p	3	6	n = 6, l = 1
6	d	5	10	n = 6, l = 2
6	f	7	14	n = 6, l = 3
6	g	9	18	n = 6, l = 4
6	h	11	22	n = 6, l = 5

Mathematically, from Schrodinger's Equation, energy level 5 would have a fifth sub-level named g. It would have 9 orbitals and hold a maximum of 18 electrons. Similarly, energy level 6 would have this g sub-level and another sub-level named h. Sub-level h would have 11 orbitals and would hold a maximum of 22 electrons. This pattern would continue through all the larger energy levels. In terms of usefulness, however, we have no atoms that contain enough electrons to use the 5g, 6g, 6h, 7g, 7h sub-levels. The known atoms never use any energy sub-levels beyond 5f, 6f, and 7f. Therefore, in most listings of energy levels and sub-levels, energy levels 5, 6, and 7 will look exactly like energy level 4, with only s, p, d, and f sub-levels listed.

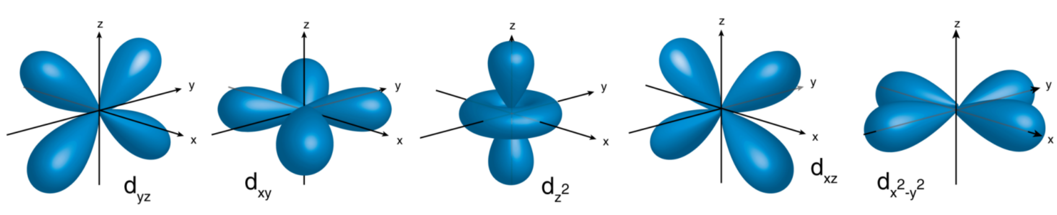
The probability patterns for these sub-levels are shown below.

The s orbitals in every energy level are spherical.

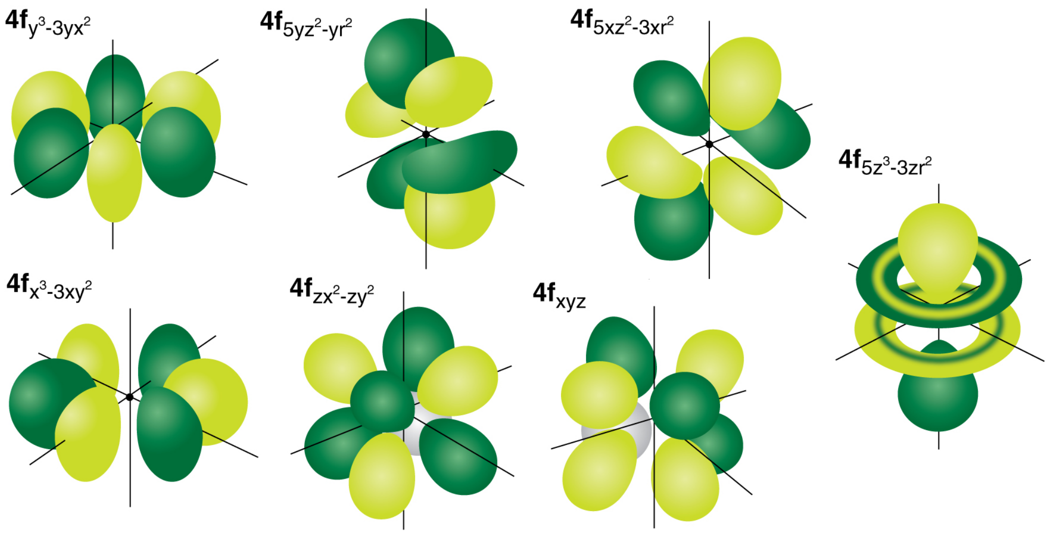
The three p orbitals in energy levels 2 - 7 are dumbbell shaped.



The five d orbitals in energy levels 3 – 7 are sometimes referred to a butterfly shaped.



The seven f orbitals in energy levels 4 – 7 are too complex to describe.



Exercises

True/False

1. All sub-energy levels with $\ell = 1$, regardless of the principal energy level quantum number will have dumbbell shape.

A. True
B. False

2. Theoretically, it is possible for a principal energy level to have n^2 sub-energy levels.

A. True
B. False

3. It is impossible for an electron in an atom to have the quantum numbers n = 3, $\ell = 2$, m$_l$ = 3, m$_s$ = +1/2.

A. True
B. False

Multiple Choice

4. How many sub-energy levels may be present if the principal quantum number is 3?

A. 1
B. 2

C. 3
D. 4
E. None of these.

5. How many possible orbitals are there when n = 3?

A. 1
B. 3
C. 4
D. 5
E. 9

6. How many electrons can be accommodated in the energy level for which n = 3?

A. 2
B. 6
C. 8
D. 10
E. 18

7. How many atomic orbitals are present in the subshell for which n = 3 and $\ell = 2$?

A. 1
B. 3
C. 5
D. 7
E. 9

8. How many orbitals are present in the subshell for which n = 5 and $\ell = 4$?

A. 1
B. 3
C. 5
D. 7
E. 9

9. What is the shape of an orbital in the subshell for which n = 3 and $\ell = 0$?

A. spherical
B. dumbbell

C. butterfly or clover shaped
D. Could be any of these.
E. None of these.

10. What is the shape of an orbital in the subshell for which n = 7 and $\ell = 0$?

A. spherical
B. dumbbell
C. butterfly or clover shaped
D. Could be any of these.
E. None of these.

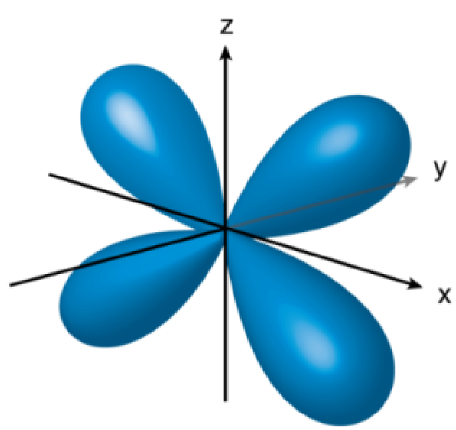
11. Which type of orbital is described by the quantum numbers n = 2, $\ell = 1$?

A. 2s
B. 2p
C. 2d
D. 2f
E. None of these.

12. If the principal quantum number of an atomic orbital is 4, what are the possible values of ℓ?

A. 0, 1, 2, 3, 4
B. 1, 2, 3, 4
C. 0, 1, 2, 3
D. 0, 1, 2
E. None of these.

Use the image below to answers questions 13, 14, and 15.



13. Identify the image above as an s-orbital, p-orbital, d-orbital, f-orbital or none of these.

A. s
B. p
C. d
D. f
E. None of these.

14. What is the ℓ value for the type of orbital pictured above?

A. 0
B. 1
C. 2
D. 3
E. 4

15. Will an orbital of the shape pictured above be found in the n = 2 energy level?

A. Yes
B. No

Chapter 7

Electron Configurations for Atoms Worksheets

7.1 Lesson 7.1 The Electron Spin Quantum Number

Quantum Numbers Worksheet

CK-12 Foundation Chemistry

Name____________________________________ **Date**_______________

1. Which quantum number indicates the electron's energy level?
2. Which quantum number indicates the electron's sub-energy level?
3. Which quantum number indicates the electron's orbital within the sub-energy level?
4. Which quantum number indicates the electron's spin?
5. What is the lowest energy level that has a ***d*** sub-level?
6. What is the total number of electrons that can exist in the 3^{rd} energy level?
7. Which sub-energy level is indicated by $\ell = 1$?
8. which sub-energy level is indicated by $\ell = 2$?
9. What is the maximum number of electrons that can be held in an ***f*** sub-energy level?
10. What does it mean for an electron to be "excited"?
11. What are the ***n*** and ℓ quantum numbers for the last electron in bromine?
12. What are the ***n*** and ℓ quantum numbers for the last electron in iron?

13. What are the $\boldsymbol{n}$ and ℓ quantum numbers for the electron in hydrogen?

14. The three electrons in the ***2p*** sub-energy level of nitrogen have the $\boldsymbol{n}$ and ℓ quantum numbers. What are the m_ℓ quantum numbers for each of these three electrons?

15. What is the basic tenet of the quantum theory?

16. Why are the quantum numbers $\boldsymbol{n} = 2$, $\ell = 2$, $m_\ell = 2$, $s = \frac{1}{2}$, not an acceptable set of quantum numbers for an electron?

17. Sketch a picture of the ***2s*** sub-energy level showing any nodes present.

18. Give the full set of quantum numbers for each of the electrons in a helium atom.

19. What maximum number of electrons in an atom can have the quantum numbers $\boldsymbol{n} = 2$, $\ell = 1$?

20. What maximum number of electrons in an atom can have the quantum numbers $\boldsymbol{n} = 3$, $\ell = 3$?

7.2 Lesson 7.2 Pauli Exclusion

There are no worksheets for this lesson.

7.3 Lesson 7.3 Aufbau Principle

There are no worksheets for this lesson.

7.4 Lesson 7.4 Writing Electron Configurations

Orbital Configuration Worksheet

CK-12 Foundation Chemistry . . . . Name______________________________________
Date________________

Table 7.1: **Draw the Orbital Configuration for these Atoms**

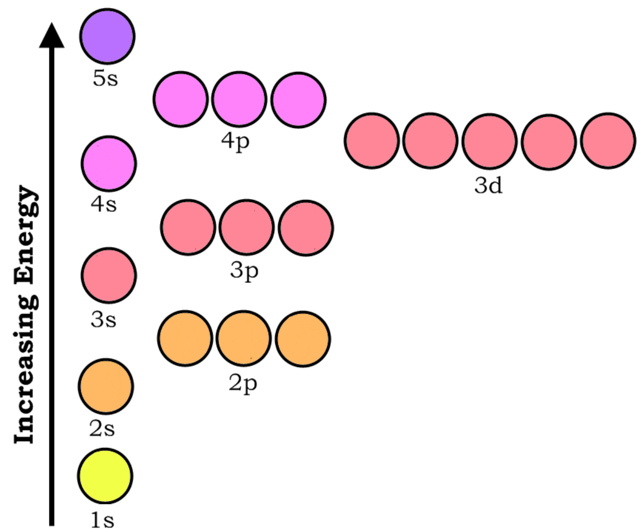
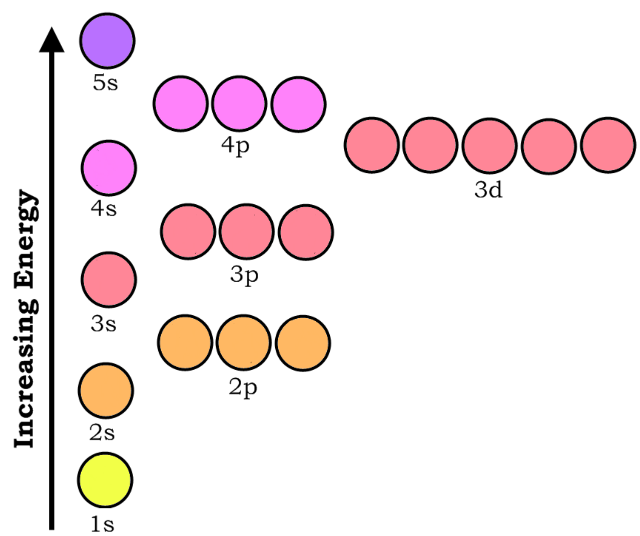
Symbol	**Orbital Diagram**
Mg	
P	

Table 7.1: (continued)

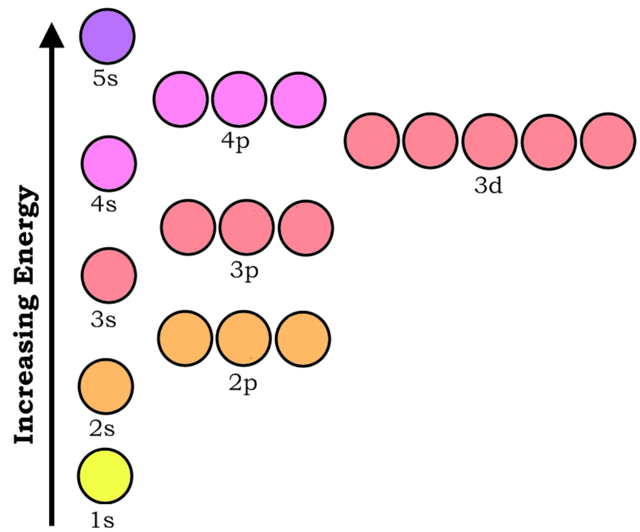
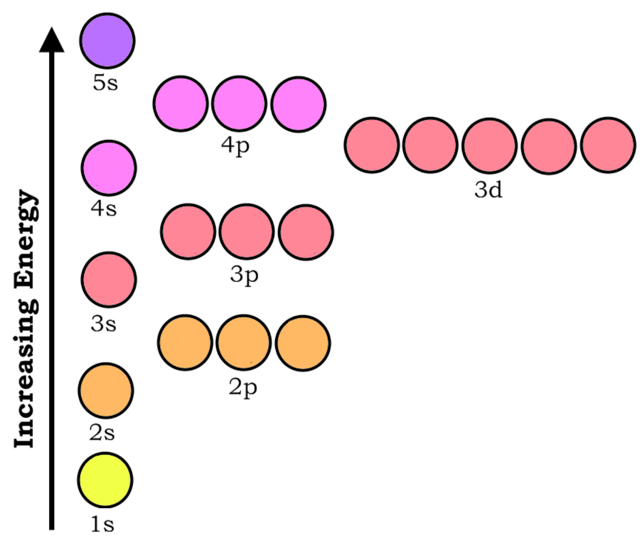
Symbol	Orbital Diagram
Ge	
Kr	

Table 7.1: (continued)

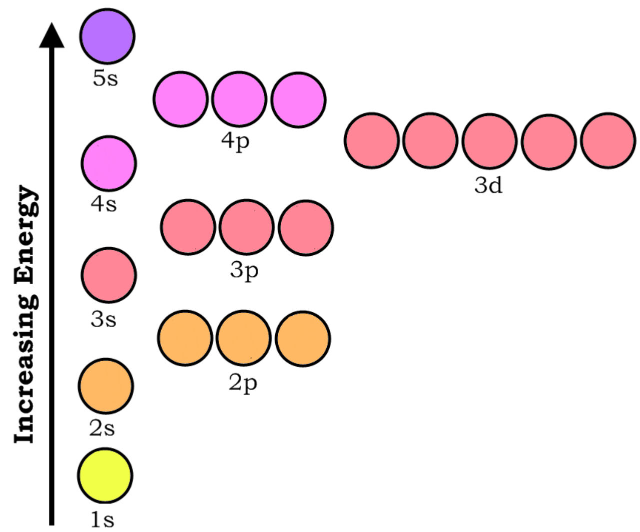
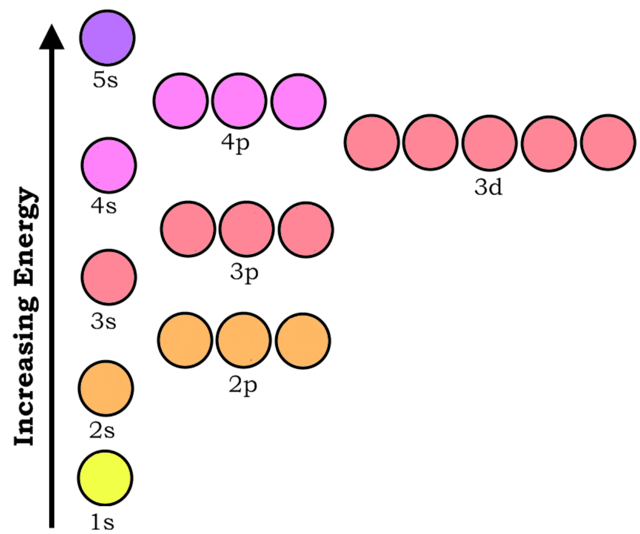
Symbol	Orbital Diagram
O	
F	

Table 7.1: (continued)

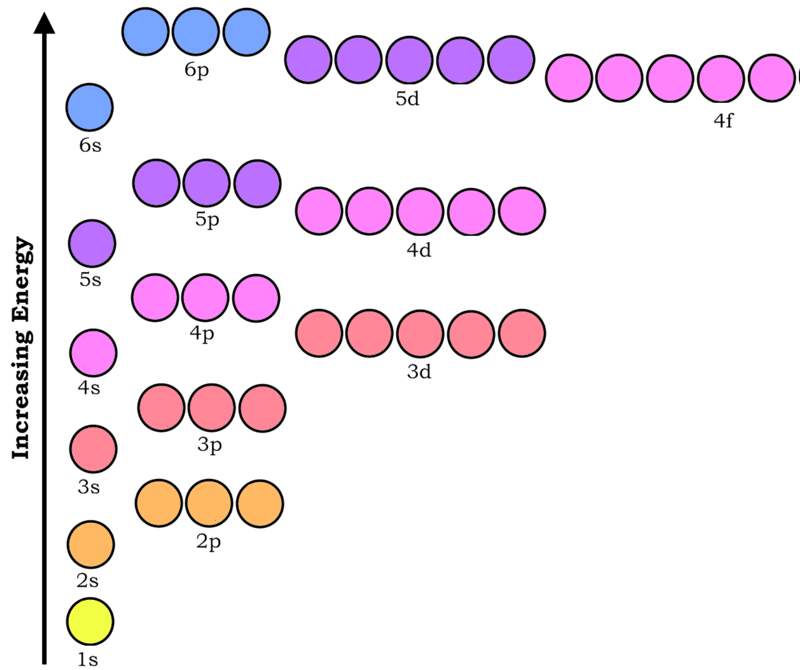
Symbol	Orbital Diagram
Pb	

Table 7.2: **Write the Electron Configuration Code for these Atoms**

Atom	**Electron Configuration Code**
V	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$
Mg	
P	
Ge	
Kr	
O	
F	
Pb	

Chapter 8

Electron Configurations and the Periodic Table Worksheets

8.1 Lesson 8.1 Electron Configurations of Main Group Elements

There are no worksheets for this lesson.

8.2 Lesson 8.2 Orbital Configurations

There are no worksheets for this lesson.

8.3 Lesson 8.3 The Periodic Table and Electron Configurations

The Periodic Table and Electron Configuration Worksheet

When Mendeleev organized the periodic table, he placed the elements in vertical columns according to their chemical behavior. That is, elements were placed in the same vertical columns because they behaved similarly in chemical reactions. All the alkali metals (Li, Na, K, Rb, Cs) react with water to produce heat, hydrogen gas, and the metal hydroxide in solution. Essentially, the only difference in the reactions is that the larger alkali metals react faster than the smaller ones. The vertical columns of elements are frequently referred to chemical "families" because of their similar chemical characteristics.

When quantum theory generated electron configurations which demonstrated that the elements in the same family have the same outer energy level electron configuration, the reason these elements behaved similarly became clear. Since chemical behavior is determined by outer energy level electron configuration, it was clear that elements that behaved similarly should have similar electron configuration.

Table 8.1: **The Electron Configuration of Family 1A Elements**

Element	**Electron Configuration**
Li	$1s^2 2s^1$
Na	$1s^2 2s^2 sp^6 3s^1$
K	$1s^2 2s^2 sp^6 3s^2 3p^6 4s^1$
Rb	$1s^2 2s^2 sp^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1$
Cs	$1s^2 2s^2 sp^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^1$

Table 8.2: **The Electron Configuration of Family 7A Elements**

Element	**Electron Configuration**
F	$1s^2 2s^2 2p^5$
Cl	$1s^2 2s^2 sp^6 3s^2 3p^5$
Br	$1s^2 2s^2 sp^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$
I	$1s^2 2s^2 sp^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^5$

Exercises

1. If the outermost energy level electron configuration of an atom is ns^2np^1,

A. to which family does it belong? ____________________
B. is the atom a metal, metalloid, non-metal, or a noble gas? ___________________________
C. how many valence electrons does it have? _______________________

2. If the outermost energy level electron configuration of an atom is ns^2np^4,

A. to which family does it belong? ____________________
B. is the atom a metal, metalloid, non-metal, or a noble gas? ___________________________
C. how many valence electrons does it have? _______________________

3. If the outermost energy level electron configuration of an atom is ns^2np^6,

A. to which family does it belong? ____________________

B. is the atom a metal, metalloid, non-metal, or a noble gas? ________________________
C. how many valence electrons does it have? _____________________

4. The electron configuration of an element is $[Ar]4s^2 3d^3$.

A. What is the identity of the element? _________________________
B. In what period does the element belong? ___________________
C. In what group does the element belong? ___________________
D. Is the element a main group element, a transition element, a lanthanide, or an actinide? ___________________

5. Write the electron configuration of only the outermost energy level for an element that is in family 5A of the fifth period of the periodic table. ____

6. Write the electron configuration of only the outermost energy level for an element that is in family 8A of the third period of the periodic table. ____

Chapter 9

Relationships Between the Elements Worksheets

9.1 Lesson 9.1 Families on the Periodic Table

There are no worksheets for this lesson.

9.2 Lesson 9.2 Electron Configurations

There are no worksheets for this lesson.

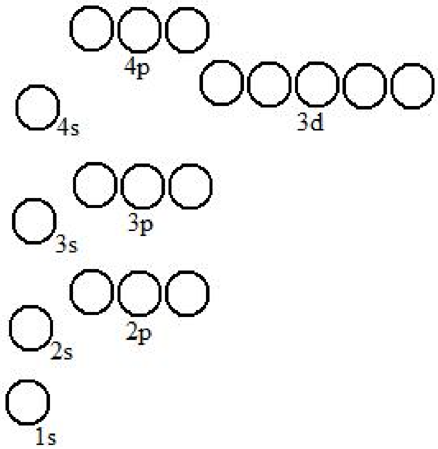
9.3 Lesson 9.3 Lewis Electron Dot Diagrams

Electron Configuration Worksheet

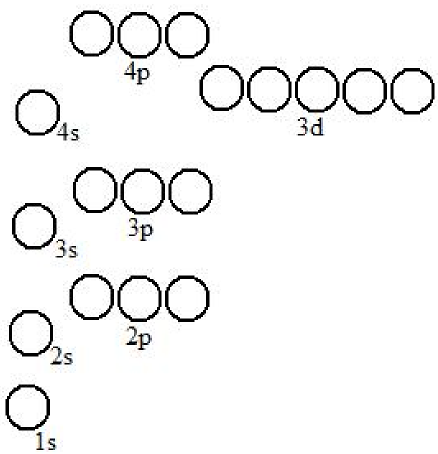
CK-12 Foundation Chemistry

Name____________________________________ **Date**_______________

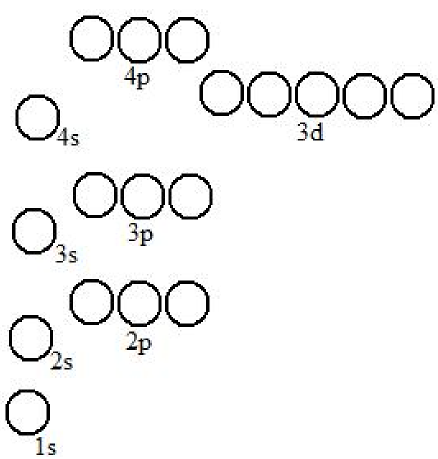
1. Fill in the orbital electron representation for phosphorus.



2. Fill in the electron orbital configuration for cobalt.



3. Fill in the electron orbital configuration for bromine.



4. Write the electron configuration code for phosphorus.
5. Draw the electron-dot formula for phosphorus.
6. How many valence electrons does phosphorus have?
7. Write the electron configuration code for cobalt.
8. How many valence electrons does cobalt have?
9. Write the electron configuration code for bromine.
10. How many valence electrons does bromine have?
11. How many valence electrons does tellurium have?
12. Give the electron dot formula for calcium.
13. What will be the outer energy level electron configuration for element #118?
14. How many valence electrons does silicon have?
15. Draw the electron dot formula for silicon.

9.4 Lesson 9.4 Chemical Family Members Have Similar Properties

There are no worksheets for this lesson.

9.5 Lesson 9.5 Transition Elements

There are no worksheets for this lesson.

9.6 Lesson 9.6 Lanthanide and Actinide Series

There are no worksheets for this lesson.

Chapter 10

Trends on the Periodic Table Worksheets

10.1 Lesson 10.1 Atomic Size

There are no worksheets for this lesson.

10.2 Lesson 10.2 Ionization Energy

There are no worksheets for this lesson.

10.3 Lesson 10.3 Electron Affinity

Trends in the Periodic Table Worksheet

CK-12 Foundation Chemistry

Name____________________________________ **Date**_______________

1. Which atom is larger, by volume, hydrogen or helium?
2. What is the smallest atom, by volume, in the third period?
3. Describe the relationship between atomic volume and ionization energy.
4. Which atom has the greatest electron affinity?
5. What is the most stable number of electrons for an atom's outermost energy level?

6. Which is larger in volume, oxygen or sulfur?
7. Which is chemically more reactive, potassium or cesium?
8. Which is chemically more reactive, oxygen or sulfur?
9. Which atom in period 3 has the greatest electron affinity?
10. Which atom in period 3 has the largest volume?
11. Which atom has greater ionization energy, aluminum or gallium?
12. Which atom has greater second ionization energy, potassium or calcium?
13. What is the outer energy level electron configuration of a noble gas?
14. Which atom in period 3 has the lowest ionization energy?
15. Explain why fluorine, even though it is larger than neon, has a greater electron affinity.

Chapter 11

Ions and the Compounds They Form Worksheets

11.1 Lesson 11.1 The Formation of Ions

Ion Formation Worksheet

Questions 1 - 4 relate to element X whose first six ionization energies are shown in the table below. Element X is a representative element.

Table 11.1: **The First Six Ionization Energies of Element X**

Number of Ionization Energy	Ionization Energy (kJ/mol)
1^{st}	800
2^{nd}	$1,400$
3^{rd}	$15,000$
4^{th}	$18,000$
5^{th}	$21,000$
6^{th}	$25,000$

1. Is element X more likely to be a metal or a non-metal?
2. Which family of elements does element X belong to?
3. How many electrons is element X most likely to gain or lose in a normal chemical reaction?
4. What is the most likely charge for an ion of element X?

Questions 5 - 8 relate to element Y whose first six ionization energies are shown in the table below. Element Y is a representative element.

Table 11.2: **The First Six Ionization Energies of Element Y**

Number of Ionization Energy	Ionization Energy (kJ/mol)
1^{st}	500
2^{nd}	$4,800$
3^{rd}	$6,800$
4^{th}	$9,000$
5^{th}	$13,000$
6^{th}	$15,000$

5. Is element Y more likely to be a metal or a non-metal?

6. Which family of elements does element Y belong to?

7. How many electrons is element Y most likely to gain or lose in a normal chemical reaction?

8. What is the most likely charge for an ion of element Y?

Questions 9 - 12 relate to element M whose first eight ionization energies are shown in the table below. Element M is a representative element.

Table 11.3: **The First Eight Ionization Energies of Element M**

Number of Ionization Energy	Ionization Energy (kJ/mol)
1^{st}	$1,100$
2^{nd}	$1,800$
3^{rd}	$2,800$
4^{th}	$4,000$
5^{th}	$6,000$
6^{th}	$8,000$
7^{th}	$27,000$
8^{th}	$36,000$

9. Is element M more likely to be a metal or a non-metal?

10. Which family of elements does element M belong to?

11. How many electrons is element M most likely to gain or lose in a normal chemical reaction?

12. What is the most likely charge for an ion of element M?

The table below gives the electron affinities for period 3 of the periodic table.

Table 11.4: **The Electron Affinities of Elements in Period Three**

Family	1A	2A	3A	4A	5A	6A	7A	8A
Electron Affinity (kJ/mol)	52	0	42	134	72	200	349	0

The table below gives the electron affinities for period 4 of the periodic table.

Table 11.5: **The Electron Affinities of Elements in Period Four**

Family	1A	2A	3A	4A	5A	6A	7A	8A
Electron Affinity (kJ/mol)	48	2	29	119	78	195	325	0

While family 5A is somewhat anomalous, the general trend is apparent in this data.

13. If a representative element has an electron affinity greater than 150 kJ/mol, would you expect it to be a metal or a non-metal?

14. If all the elements in a family have an electron affinity of 0 kJ/mol, what family is it most likely to be?

15. The first ionization energies (in kJ/mol) of Li, Na, K, Rb, and Cs in random order are 370, 520, 400, 500, and 420.

A. Which first ionization energy do you think belongs to Li?
B. Which first ionization energy do you think belongs to Cs?
C. What knowledge about chemical families did you use to make those choices?

16. Given the electron configuration of the outermost energy level of an atom to be s^2p^4 :

A. is the element a metal or non-metal?
B. is it most likely to gain or lose electrons?
C. how many electrons is it most likely to gain or lose in a normal chemical reaction?
D. what is the most likely charge on an ion of this element?

11.2 Lesson 11.2 Ionic Bonding

There are no worksheets for this lesson.

11.3 Lesson 11.3 Properties of Ionic Compounds

There are no worksheets for this lesson.

Chapter 12

Writing and Naming Ionic Formulas Worksheets

12.1 Lesson 12.1 Predicting Formulas of Ionic Compounds

Formula Writing Worksheet

CK-12 Foundation Chemistry

Name______________________________________ **Date**________________

Fill in the squares with the appropriate formula for the compound formed by the combination of the atoms or ions that intersect.

Table 12.1: **Formula Writing Practice**

	bromine	acetate	sulfate	phosphate	hydroxide	sulfur
potassium						
calcium						
aluminum						
ammonium						
iron (III)						
lead (II)						

12.2 Lesson 12.2 Inorganic Nomenclature

Inorganic Nomenclature Worksheet

CK-12 Foundation Chemistry

Name________________________________ **Date**______________

Table 12.2: **Name the Following Compounds**

Number	Formula	Name
1.	LiF	________________
2.	Na_3PO_4	________________
3.	$Al(OH)_3$	________________
4.	Cl_2O_7	________________
5.	PbO	________________
6.	Fe_2S_3	________________
7.	TeO_2	________________
8.	$CuSO_4$	________________
9.	$Ca_3(PO_4)_2$	________________
10.	HNO_3	________________

Table 12.3:

Number	Name	Formula
1.	copper (I) sulfide	________________
2.	boron trichloride	________________
3.	potassium carbonate	________________
4.	sulfur hexafluoride	________________
5.	chlorine monofluoride	________________
6.	dinitrogen tetraoxide	________________
7.	tin (IV) oxide	________________
8.	silver acetate	________________
9.	diphosphorus pentoxide	________________
10.	lithium nitrate	________________

Chapter 13

Covalent Bonding Worksheets

13.1 Lesson 13.1 The Covalent Bond

There are no worksheets for this lesson.

13.2 Lesson 13.2 Atoms that Form Covalent Bonds

There are no worksheets for this lesson.

13.3 Lesson 13.3 Naming Covalent Compounds

There are no worksheets for this lesson.

Chapter 14

Molecular Architecture Worksheets

14.1 Lesson 14.1 Types of Bonds that Form Between Atoms

There are no worksheets for this lesson.

14.2 Lesson 14.2 The Covalent Molecules of Family 2A-8A

There are no worksheets for this lesson.

14.3 Lesson 14.3 Resonance

There are no worksheets for this lesson.

14.4 Lesson 14.4 Electronic and Molecular Geometry

There are no worksheets for this lesson.

14.5 Lesson 14.5 Molecular Polarity

Molecular Geometry Worksheet

CK-12 Foundation Chemistry

Name__________________________________ **Date**______________

Lewis structures only show how many bonding pairs of electrons, and unshared pairs of electrons, surround a given atom on a flat page. The molecules are actually three dimensional which is not shown by Lewis structures. To convey a sense of three dimensionality, we use "ball and stick" models.

There is a correlation between the number of electron pairs, (sigma bonds plus non-shared pairs) around the central atom of a molecule, and the electronic geometry of that molecule.

The idea that allows us to predict the electronic geometry is that each pair of electrons (shared or unshared) repels all the other electron pairs. The electron pairs move as far apart as possible, but since they are all tied to the central atom, they can only orient themselves in such a way that they make the angles between them as large as possible. This is the essence of the ***Valence Shell Electron Pair Repulsion (VSEPR) Theory*** for predicting molecular shapes.

To use VSEPR theory, we must first be able to determine the number of valence shell electron pairs around the central atom. These pairs consist of all sigma bond pairs and all unshared pairs of electrons. Pi bond electrons are excluded because the electrons are not placed between bonding atoms and therefore, do not contribute to electronic geometry.

Table 14.1: **Visualizing Electron Pairs**

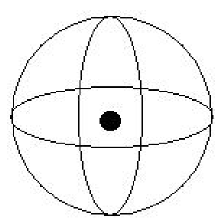
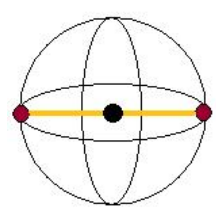
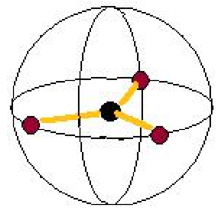
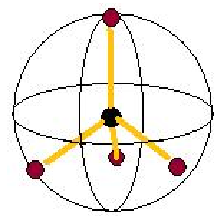
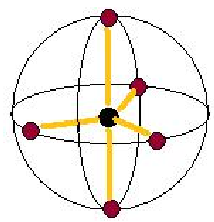
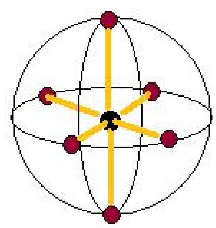
Electron Pairs	**Image**
To visualize the electron pairs that contribute to electronic geometry, imagine them situated on the surface of a sphere with the central atom at the center.	
If there are only two pairs of electrons in the valence shell of the central atom, the two pairs can avoid each other best if they are 180^o apart. This means that the two pairs and the central atom are in a straight line; the arrangement is **linear**.	

Table 14.1: (continued)

Electron Pairs	Image
If a third pair of electrons is added, the three pairs push around to the shape shown at right. The angles between electron pairs would be 120^o and we call the shape **trigonal planar**. The three pairs of electrons and the central atom are all in a single plane.	
A fourth pair of electrons causes the electrons to push around into the shape shown at right, the **tetrahedron**. The angles in this shape are 109.5^o.	
A fifth pair of electrons produces the shape known as **trigonal bipyramidal**. The angles between the three pairs of electrons around the center is 120^o and the angles between the pairs around the center and the pairs on the ends is 90^o.	
Finally, the sixth pair of electrons produces the **octahedral** shape shown at right. All angles in this shape are 90^o.	

Once the number of electron pairs surrounding the central atom is determined, the **electronic geometry** is known.

Table 14.2: **The Relationship Between Number of Electron Pairs and Electronic Geometry**

Electron Pairs Around the Central Atom	Electronic Geometry
1 pair	Linear
2 pairs	Linear
3 pairs	Trigonal Planar
4 pairs	Tetrahedral

Table 14.2: (continued)

Electron Pairs Around the Central Atom	Electronic Geometry
5 pairs	Trigonal Bipyramidal
6 pairs	Octahedral

The **molecular** geometry may be different from the **electronic** geometry because many times, not all the electron pairs are shared. An unshared electron pair will not have an atom in that position of the electronic geometry. In order to determine molecular geometry, we must recognize which pairs of electrons have an atom attached and which are lone pairs. The overall shape of the molecule is determined by how many pairs of electrons are around the central atom, and how many of these have atoms attached.

It is sometimes difficult for students to recognize the difference between the orientation of electron pairs (called electronic geometry) and the overall shape of the molecule (called molecular geometry). We will look at an example that shows the difference between electronic and molecular geometry. Consider the following four molecules: hydrogen chloride, HCl; water, H_2O; ammonia, NH_3; and methane, CH_4.

Table 14.3: **The Relationship Between Shared Pairs and Molecular Geometry**

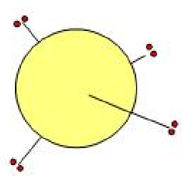
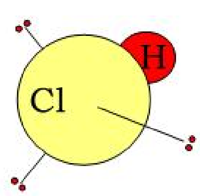
Shared Pairs	Molecular Geometry
The central atom of each of these molecules is surrounded by four pairs of electrons. According to VSEPR theory, these four pairs will be oriented in three-dimensional space to be as far away from each other as possible. The four pairs will point to the corners of the geometrical shape known as a tetrahedron. The angles between the electron pairs will be approximately 109.5^o. In all four cases, the **electronic** geometry is **tetrahedral** but only one of the molecules will have tetrahedral molecular geometry.	
In the case of HCl, even though there are four pairs of electrons around the chlorine atom, three of them are not shared. There is no atom attached to them. These spaces are empty. Since there are only two atoms joined by a bond, the molecular geometry will be **linear**.	

Table 14.3: (continued)

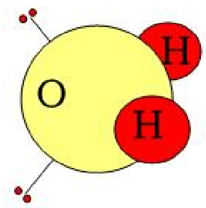
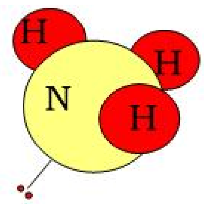
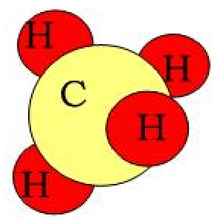
Shared Pairs	**Molecular Geometry**
In the water molecule, two electron pairs are shared and two are unshared. So while the electronic geometry is tetrahedral, the molecular geometry is **bent** (aka angular, aka V-shaped).	
In the ammonia molecule, one pair of electrons is unshared and the other three are shared. This results in a molecular shape called **pyramidal**.	
In the methane molecule, all four pairs of electrons are shared, and so not only is the electronic geometry tetrahedral but the molecular geometry is also **tetrahedral**.	

Table 14.4:

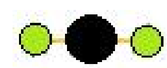
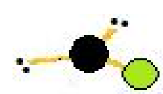
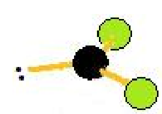
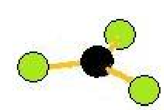
Central Atom Electron Pairs	**Electronic Geometry**	**Bonding Pairs**	**Molecular Geometry**	**Sketch**
2	Linear	2	Linear	
3	Trigonal Planar	1	Linear	
3	Trigonal Planar	2	Bent	
3	Trigonal Planar	3	Trigonal Planar	

Table 14.4: (continued)

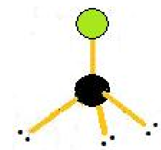
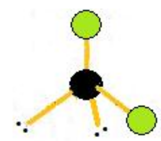
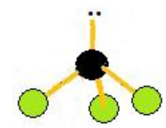
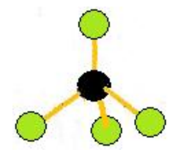
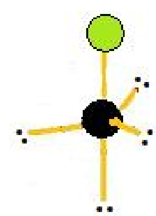
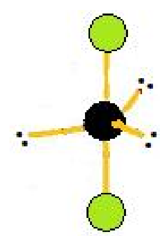
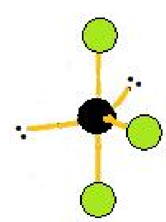
Central Atom Electron Pairs	Electronic Geometry	Bonding Pairs	Molecular Geometry	Sketch
4	Tetrahedral	1	Linear	
4	Tetrahedral	2	Bent	
4	Tetrahedral	3	Pyramidal	
4	Tetrahedral	4	Tetrahedral	
5	Trigonal Bipyramidal	1	Linear	
5	Trigonal Bipyramidal	2	Linear	
5	Trigonal Bipyramidal	3	T-shape	

Table 14.4: (continued)

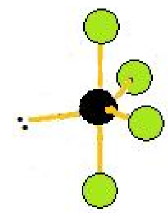
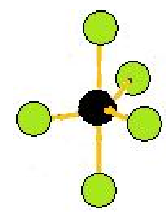
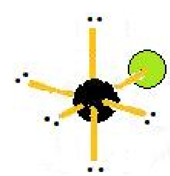
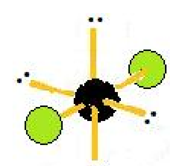
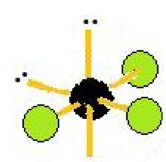
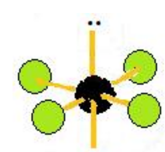
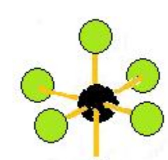
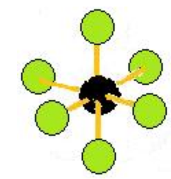
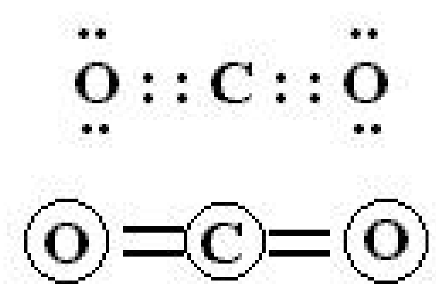
Central Atom Electron Pairs	**Electronic Geometry**	**Bonding Pairs**	**Molecular Geometry**	**Sketch**
5	Trigonal Bipyramidal	4	Distorted Tetrahedron	
5	Trigonal Bipyramidal	5	Trigonal Bipyramidal	
6	Octahedral	1	Linear	
6	Octahedral	2	Linear	
6	Octahedral	3	T-shape	
6	Octahedral	4	Square Planar	
6	Octahedral	5	Square Pyramidal	

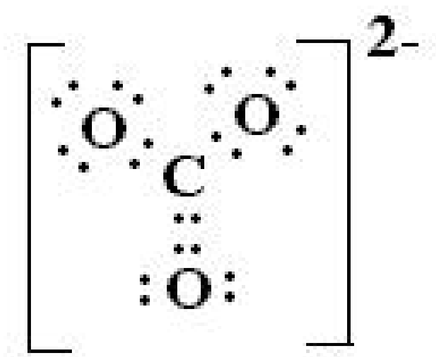
Table 14.4: (continued)

Central Atom Electron Pairs	Electronic Geometry	Bonding Pairs	Molecular Geometry	Sketch
6	Octahedral	6	Octahedral	

In order to choose the correct molecular geometry, you must keep in mind that only electron pairs involved in sigma bonds and unshared pairs contribute to electronic geometry. Pi bonds are not directed bonds, and those electron pairs do not contribute to electronic geometry. In the Lewis structure for the carbon dioxide molecule (shown at right), it is clear that the central atom is carbon, and the carbon atom is surrounded by 4 pairs of electrons. But these fours pairs of electrons are involved in two sigma bonds and two pi bonds. Therefore, the electronic geometry of carbon dioxide is based on **two** pairs of electrons around the central atom, and will be linear. Since both pairs of electrons are shared, the molecular geometry will also be linear.



The Lewis structure for the carbonate ion, shown at right, shows the central atom is carbon and it is surrounded by 4 electron pairs. One of those pairs, however, is a pi bond, and therefore the electronic geometry of the carbonate ion is based on 3 pairs of electrons around the central atom. Thus, the electronic geometry is trigonal planar and since all three pairs are shared, the molecular geometry is also trigon planar.



Polarity

Bonds between atoms that are of the same element are non-polar bonds. Molecules composed of all the same atom such as Cl_2, O_2, H_2, S_8, P_4, have no polar bonds and therefore do not have dipoles. That is, the molecules will be non-polar. A molecule that does have polar bonds can still be non-polar. If the polar bonds are symmetrically distributed, the bond dipoles cancel and do not produce a molecular dipole.

Table 14.5: **Symmetrical and Non-Symmetrical Molecular Shapes**

Molecular Shape	**Symmetry**
Linear	Symmetrical
Bent	Non-Symmetrical
Trigonal Planar	Symmetrical
Pyramidal	Non-Symmetrical
Tetrahedral	Symmetrical
T-shaped	Non-Symmetrical
Distorted Tetrahedron	Non-Symmetrical
Trigonal Bipyramidal	Symmetrical
Square Planar	Symmetrical
Square Pyramidal	Non-Symmetrical
Octahdral	Symmetrical

Exercises

Fill in the table with electronic geometry, molecular geometry, and indicate whether the molecular will be polar or non-polar.

Table 14.6: **Polarity Table**

Formula	**Electronic Geometry**	**Molecular Geometry**	**Polarity**
AsH_3			

Table 14.6: (continued)

Formula	**Electronic Geometry**	**Molecular Geometry**	**Polarity**
BCl_3			
IF_3			
$SiBr_4$			
SeH_4			
XeI_4			
OF_2			
KrF_2			
ICl_5			
CCl_2F_2			

Chapter 15

The Mathematics of Compounds Worksheets

15.1 Lesson 15.1 Determining Formula and Molecular Mass

Calculating Molar Masses Worksheet

CK-12 Foundation Chemistry

Name____________________________________ **Date**_______________

The relative masses of atoms, in units called Daltons, are listed in the periodic table. The relative masses of molecules, in the same units, can be determined by adding up the masses of all the atoms that make up the molecule. For example, the periodic table lists the relative mass of a hydrogen atom as 1.01 Dalton and relative mass of the an oxygen atom to be 16.00 Daltons. Therefore, on this same scale, the relative mass of a water molecule, H_2O, would be the sum of two hydrogen atoms and one oxygen atom, $1.01 + 1.01 + 16.00 = 18.02$ Daltons.

When an Avogadro's number, $6.02\ x\ 10^{23}$, of atoms or molecules are taken, the mass of the group will be the same number as the relative mass, but the units will be grams. That is, the mass in grams, of $6.02\ x\ 10^{23}$ water molecules is 18.02 **grams**. An Avogadro's number of particles is called one **mole** and the mass of that group of particles is called the ***molar mass*** (or mass of one mole) of that substance.

Example: Find the molar mass of calcium phosphate, $Ca_3(PO_4)_2$.

Table 15.1: **Adding Up a Molar Mass**

Atoms of Element	Atoms x Atomic Mass	Total
3 Ca atoms =	3 x 40.1	= 120.3
2 P atoms =	2 x 31.0	= 62.0
8 O atoms =	8 x 16.0	= 128.0
		310.3

Therefore, the molar mass of calcium phosphate is 310.3 grams/mole.

Exercises

Find the molar masses of the following compounds. (Do not fail to include units in your answers.)

1. $NaOH$
2. $NaBr$
3. $PbSO_4$
4. $Ca(OH)_2$
5. AgF
6. $C_6H_{12}O_6$
7. $Ba(C_2H_3O_2)_2$
8. $ZnCl_2$
9. $(NH_4)_2SO_4$
10. $(NH_4)_3PO_4$

15.2 Lesson 15.2 The Mole

Moles Worksheet

CK-12 Foundation Chemistry

Name____________________________________ **Date**_______________

An Avogadro's number of particles of a substance is called ***one mole*** of that substance. When an Avogadro's number, $6.02\ x\ 10^{23}$, of atoms or molecules are taken, the mass of the group will be the same number as the relative molecular mass, but the units will be grams.

The mass of one mole of a substance ($6.02 \ x \ 10^{23}$ particles) is the relative molecular mass **in grams**.

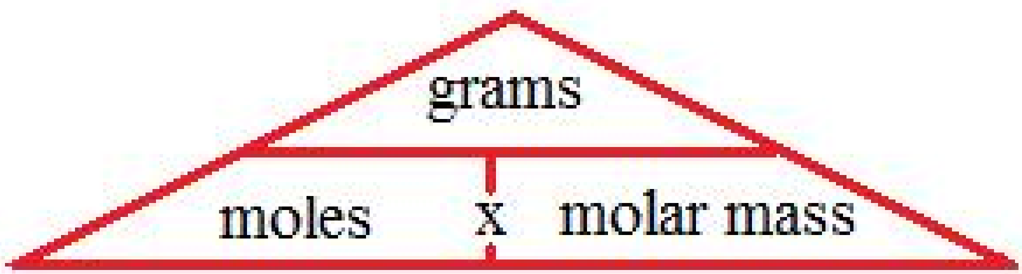
The relationship between the moles and mass of a substance is given by:

$$\text{grams} = (\text{moles})(\text{molar mass})$$

This relationship can be solved for any one of the three variables in the expression.

grams = (moles)(molar mass)	$\text{moles} = \dfrac{\text{grams}}{\text{molar mass}}$	$\text{molar mass} = \dfrac{\text{grams}}{\text{moles}}$

Some students find the triangle below to be a useful crutch. You put your thumb over the quantity you are solving for and the part of the triangle not covered shows the correct formula.



Example 1: How many moles are present in 10.0 grams of sodium hydroxide, $NaOH$?

Solution: The molar mass of $NaOH$ is 40.0 g/mol. ("mol" is the abbreviation of mole.)
$\text{moles} = \frac{grams}{molar \ mass} = \frac{10.0 \ g}{40.0 \ g/mol} = 0.250 \text{ moles}$

Example 2: What is the mass, in grams, of 4.20 moles of $Ca(NO_3)_2$?

Solution: The molar mass of $Ca(NO_3)_2$ is 164.1 g/mol.
grams = (moles)(molar mass) = (4.20 mol)(164.1 g/mol) = 689 grams

Example 3: What is the molar mass of an unknown substance is 0.250 moles of the substance has a mass of 52.6 grams?

Solution:

molar mass = $\frac{grams}{moles} = \frac{56.2\ g}{0.250\ mol} = 225$ g/mol

Example 4: What is the mass of 3.01 x 10^{23} molecules of ammonia, NH_3?

Solution: This problem involves converting the number of molecules to moles (divide by Avogadro's number), and then multiplying the moles by the molar mass.

mass = (3.01 x 10^{23} molecules)$\left(\frac{1.00\ mol}{6.02\ x\ 10^{23}\ molecules}\right)\left(\frac{17.0\ g}{1.00\ mol}\right)$ = 8.50 grams

Exercises

1. How many moles are present in 5.00 grams of $NaOH$?
2. How many grams are present in 2.50 moles of NH_3?
3. How many moles are present in 100. g of $Ca(NO_3)_2$?
4. What is the mass of 0.468 moles of $C_6H_{12}O_6$?
5. How many moles are present in 1.00 x 10^{24} molecules of water?
6. What is the mass, in grams, of one molecule of water?
7. What is the molar mass of a substance if 0.336 moles of it has a mass of 70.0 grams?
8. Convert 4.00 grams of CH_4 to moles.
9. Convert 4.00 moles of CH_4 to grams.
10. How many molecules are present in 1.00 g of $Al(C_2H_3O_2)_2$

15.3 Lesson 15.3 Percent Composition

Percent Composition Worksheet

CK-12 Foundation Chemistry

Name__________________________________ **Date**______________

The percent composition (or percentage composition) of a compound is a measure of the percentage of each different element present in the compound. To calculate the percent composition, the mass of each individual element is divided by the total mass of the compound and then multiplied by 100 (to get its percentage). The percent composition of a compound can be calculated either from the known masses of the elements in the compound (determined in the lab) or from the formula of the compound.

Example: The composition of a compound is determined in the laboratory to be 5.748 grams of sodium and 8.862 grams of chlorine. What is the percent composition of the compound?

Solution: The total mass of this sample of the compound is 14.61 grams.
% sodium = $\frac{5.748\ g}{14.61\ g}$ x 100 = 39.34%
% chlorine = $\frac{8.862\ g}{14.61\ g}$ x 100 = 60.66%

When you add up all the percentages of elements, you should get 100%, although on many occasions, rounding may cause the last digit of the total to be off by 1. That is, on occasion, you get a total of 99.9% or 100.1% due to several individual percentages all being rounded up or all being rounded down.

Example: Calculate the percent composition of all the elements in $(NH_4)_3PO_4$.

Solution:
3 N atoms = 3 x 14.01 = 42.03
12 H atoms = 12 x 1.01 = 12.12
1 P atom = 1 x 30.97 = 30.97
4 O atoms = 4 x 16.00 = 64.00
Formula weight for $(NH_4)_3PO_4$ = 149.12
% N =$\frac{42.03}{149.12}$ x 100 = 28.19% % P = $\frac{30.97}{149.12}$ x 100 = 20.77%
% H =$\frac{12.12}{149.12}$ x 100 = 8.13% % O = $\frac{64.00}{149.12}$ x 100 = 42.92%

When the four percentages are added in this case, the total is 100.01%. The extra 0.01% is due to the fact that all four of these percentages were rounded up.

Exercises

1. Determine the percent composition of Na_2SO_4.
2. Determine the percent composition of $NaOH$.
3. Determine the percent composition of $AlCl_3$.
4. Determine the percent composition of $Ca(C_2H_3O_2)_2$.
5. Determine the percent composition of $C_6H_12O_6$.

15.4 Lesson 15.4 Empirical and Molecular Formulas

Empirical Formulas Worksheet

CK-12 Foundation Chemistry

Name______________________________________ **Date**_______________

Empirical formulas represent the simplest whole number ratio of the atoms that make up a compound. In some cases, such as CO_2, the empirical formula is exactly the same as the actual molecular formula. In other cases such as benzene, C_6H_6, whose empirical formula is CH, the molecular formula is some multiple of the empirical formula.

Empirical formulas can be determined either from the masses of the elements making up the compound or from the percent composition.

Example 1: What is the empirical formula of a compound that contains 0.0134 grams of iron, 0.00769 grams of sulfur, and 0.0115 grams of oxygen?

Step 1: Convert each of the masses into moles of atoms of that element. This is accomplished by dividing the grams of each element by the atomic mass of the element.

$$\text{moles Fe} = \frac{0.0134\ g}{55.8\ g/mol} = 0.000240\ \text{mol}$$
$$\text{moles S} = \frac{0.00769\ g}{32.1\ g/mol} = 0.000240\ \text{mol}$$
$$\text{moles O} = \frac{0.0115\ g}{16.0\ g/mol} = 0.000719\ \text{mol}$$

It is important to note that we are determining the number of moles of each atom that exists in the compound and therefore, for the diatomic gases, we use the atomic mass of a single atom of the element (not the diatomic molar mass).

Step 2: The ratio of moles that we determined in step 1 is the correct ratio for the compound. We are not allowed, however, to write a formula in the form, $Fe_{0.000230}S_{0.000240}O_{0.000719}$. Before we can write the formula, we must get the ratio into a simplest whole number ratio. This is often accomplished by dividing each of the moles by the smallest of them.

$$\text{moles Fe} = \frac{0.000240}{0.000240} = 1.00$$
$$\text{moles S} = \frac{0.000240}{0.000240} = 1.00$$
$$\text{moles O} = \frac{0.000719}{0.000240} = 3.00$$

Therefore, the empirical formula for this compound is $FeSO_3$.

Example 2: Find the empirical formula of a compound that contains 48.78% carbon, 2.439% hydrogen, 26.02% oxygen, and 22.77% nitrogen.

Solution: When the empirical formula is to be determined from percent composition, it is easiest to assume a 100. gram sample, take each percentage of the 100. grams to get grams for each element, and then proceed as in **Example 1**. Using this technique, each of the percentages in the problem becomes the mass of the element in grams.

grams C = 48.78 g, grams H = 2.439 g grams O = 26.02 g, grams N = 22.77 g

Step 1:

$$\text{moles C} = \frac{48.78\ g}{12.01\ g/mol} = 4.062 \text{ mols}$$
$$\text{moles H} = \frac{2.439\ g}{1.01\ g/mol} = 2.415 \text{ mols}$$
$$\text{moles O} = \frac{26.02\ g}{16.00\ g/mol} = 1.626 \text{ mols}$$
$$\text{moles N} = \frac{22.77\ g}{14.01\ g/mol} = 1.625 \text{ mols}$$

Step 2: Divide each of the moles by the smallest.

$$\text{moles C} = \frac{4.062}{1.625} = 2.50$$
$$\text{moles H} = \frac{2.415}{1.625} = 1.49$$
$$\text{moles O} = \frac{1.626}{1.625} = 1.00$$
$$\text{moles N} = \frac{1.625}{1.625} = 1.00$$

Step 3: In a case, such as this one, where step 2 does NOT produce a simple whole number ratio, we then choose a multiplier with which to multiply each of the final numbers such that we do get a simple whole number ratio. This is usually an integer between 2 and 5 but could possible be a larger integer. In this case, the multiplier is 2.

$$\text{moles C} = 5, \text{ moles H} = 3, \text{ moles O} = 2, \text{ moles N} = 2$$

Therefore, the empirical formula for this compound is $C_5H_3O_2N_2$.

Exercises

1. Find the empirical formula for a compound that is 75.0% carbon and 25.0% hydrogen.

2. Find the empirical formula for a compound that is 32.8% chromium and 67.2% chlorine.

3. Find the empirical formula for a compound that is 67.1% zinc and the rest oxygen.

4. A sample of a compound was found to contain 0.62069 g of carbon, 0.10345 g of hydrogen, and 0.27586 g of oxygen. What is the empirical formula?

5. A sample of a compound was found to contain 48.65% carbon, 8.11% hydrogen, and 43.24% oxygen. What is its empirical formula?

Molecular Formulas Worksheet

CK-12 Foundation Chemistry

Name____________________________________ **Date**_______________

Empirical formulas show the simplest whole number ratio of the atoms of the elements that make up a compound. Molecular formulas show the actual number of atoms of each element that make up the compound. The molecular formula for benzene is C_6H_6 but the empirical formula for benzene would be the simplest whole number ratio for these atoms, which would be CH. The empirical formula can be determined from either the masses of the elements in a compound or from percent composition. In order to determine the molecular formula, we also need the molar mass of the compound. The molecular formula will always be some whole number multiple of the empirical formula. In the case of benzene, the multiplier is 6.

The molecules C_2H_4, C_3H_6, C_4H_8, and C_5H_{10} all have the same empirical formula, namely CH_2. If we have the empirical formula CH_2 and the molar mass of 56 g/mol for a compound, we can determine the molecular formula by dividing the formula mass of CH_2 into the molar mass to find the multiplier. The formula mass of CH_2 is 14 g/mol. When we divide the formula mass, 14 g/mol, into the molar mass, 56 g/mol, we get the multiplier 4. Therefore, the molecular formula for this compound is 4 times the empirical formula. $CH_2\ x\ 4\ =\ C_4H_8$.

Example: What is the molecular formula for a compound with the empirical formula HCO_2 and a molar mass of 90. g/mol?

Solution: The formula mass of HCO_2 is 45 g/mol. Dividing 45 g/mol into 90. g/mol yields a multiplier of 2. Therefore, the molecular formula for this compound is $2\ x\ CHO_2\ =\ H_2C_2O_4$.

Exercises

1. A compound has the empirical formula C_2OH_4 and a molar mass of 88 g/mol. What is its molecular formula?

2. A compound has the empirical formula C_4H_4O and a molar mass of 136 g/mol. What is its molecular formula?

3. A compound has the empirical formula $CFBrO$ and a molar mass of 254.7 g/mol. What is its molecular formula?

4. A compound is 7.692% hydrogen and 93.308% carbon. Its molar mass is 104 g/mol. What is its molecular formula?

5. A compound is 47.0% potassium, 14.5% carbon, and 38.5% oxygen. Its molar mass is 166.2 g/mol. What is its molecular formula?

Chapter 16

Chemical Reactions Worksheets

16.1 Lesson 16.1 Chemical Equations

There are no worksheets for this lesson.

16.2 Lesson 16.2 Balancing Equations

Balancing Equations Worksheet

CK-12 Foundation Chemistry

Name____________________________________ **Date**_______________

Balance the following equations by inserting the smallest whole number coefficients.

1. $\_\_ \ CuCl \ + \ \_\_ \ H_2S \ \rightarrow \ \_\_ \ Cu_2S \ + \ \_\_ \ HCl$
2. $\_\_ \ Na \ + \ \_\_ \ H_2O \ \rightarrow \ \_\_ \ NaOH \ + \ \_\_ \ H_2$
3. $\_\_ \ Mg \ + \ \_\_ \ O_2 \ \rightarrow \ \_\_ \ MgO$
4. $\_\_ \ Fe \ + \ \_\_ \ O_2 \ \rightarrow \ \_\_ \ Fe_2O_3$
5. $\_\_ \ H_2O \ + \ \_\_ \ N_2O_3 \ \rightarrow \ \_\_ \ HNO_2$
6. $\_\_ \ Fe \ + \ \_\_ \ H_2O \ \rightarrow \ \_\_ \ Fe_3O_4 \ + \ \_\_ \ H_2$
7. $\_\_ \ Al \ + \ \_\_ \ Pb(NO_3)_2 \ \rightarrow \ \_\_ \ Al(NO_3)_3 \ + \ \_\_ \ Pb$
8. $\_\_ \ KOH \ + \ \_\_ \ H_3PO_4 \ \rightarrow \ \_\_ \ K_3PO_4 \ + \ \_\_ \ H_2O$
9. $\_\_ \ C_2H_6 \ + \ \_\_ \ O_2 \ \rightarrow \ \_\_ \ CO_2 \ + \ \_\_ \ H_2O$

10. ___ C_2H_5OH + ___ O_2 → ___ CO_2 + ___ H_2O

11. ___ N_2 + ___ H_2 → ___ NH_3

12. ___ $Al(OH)_3$ + ___ H_2SO_4 → ___ $Al_2(SO_4)_3$ + ___ H_2O

13. ___ $SbCl_3$ + ___ H_2S → ___ Al_2S_3 + ___ HCl

14. ___ C_5H_{12} + ___ O_2 → ___ CO_2 + ___ H_2O

15. ___ NH_4Cl + ___ $Ca(OH)_2$ → ___ $CaCl_2$ + ___ NH_3 + ___ H_2O

Convert the following word equations into formula equations and then balance them.

16. Iron + oxygen yields iron (III) oxide.

17. Antimony + chlorine yields antimony (III) chloride.

18. Sodium chlorate ($NaClO_3$) yields sodium chloride + oxygen.

19. Lead (II) nitrate + hydrogen sulfide yields lead (II) sulfide + nitric acid (HNO_3).

20. Aluminum + sulfuric acid (H_2SO_4) yields aluminum sulfate + hydrogen gas.

16.3 Lesson 16.3 Types of Reactions

Types of Chemical Reactions Worksheet

There are millions of different compounds and therefore, there must be millions of different chemical reactions to form these compounds. When chemists are confronted with an overwhelming number of things, they tend to classify them into groups in order to make them easier to study and discuss. One popular system of classification for chemical reactions places them in five major categories. Some of the categories have different names in different books and you should become familiar with all the names.

Types of Chemical Reactions

1. Synthesis (also called Direct Combination)

A synthesis reaction occurs when two or more substances combine to make a single, more complex substance. The reactants may be elements or compounds but the product will always be a compound. The general formula for this type of reaction can be shown as:

$$A + B \rightarrow AB$$

Some examples of synthesis reactions are shown below.

$$2\ H_{2(g)} + O_{2(g)} \rightarrow 2\ H_2O_{(g)}$$

$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$
$CaO_{(s)} + H_2O_{(L)} \rightarrow Ca(OH)_{2(s)}$

You should note in each case above, there are two or more substances in the reactants and only one substance as the product.

2. Decomposition (also called Analysis)

A decomposition reaction occurs when one substance is broken down into two or more simpler substances. This type of reaction is the opposite of a synthesis reaction, as shown by the general formula below:

$$AB \rightarrow A + B$$

Some examples of decomposition reactions are shown below.

$C_{12}H_{22}O_{11(s)} \rightarrow 12\ C_{(s)} + 11\ H_2O_{(g)}$
$Pb(OH)_{2(s)} \rightarrow PbO_{(s)} + H_2O_{(g)}$
$2\ Ag_2O_{(s)} \rightarrow 4\ Ag_{(s)} + O_{2(g)}$

3. Single Displacement (also called Single Replacement)

In this type of reaction, a neutral element becomes as ion as it replaces another ion in a compound. The general form of this equation can be written as:

$$A + BC \rightarrow B + AC \qquad \text{(positive ion replaced)}$$

Or

$$A + BC \rightarrow C + BA \qquad \text{(negative ion replaced)}$$

In either case, the equation is element + compound → element + compound.

Some examples of single displacement reactions are shown below.

$Zn_{(s)} + H_2SO_{4(aq)} \rightarrow ZnSO_{4(aq)} + H_{2(g)}$
$2\ Al_{(s)} + 3\ CuCl_{2(aq)} \rightarrow 2\ AlCl_{2(aq)} + 3\ Cu_{(s)}$
$Cl_{2(g)} + KBr_{(aq)} \rightarrow KCl_{(aq)} + Br_{2(L)}$

4. Double Displacement (also called Double Replacement and Metathesis)

In this reaction type, pairs of ionic compounds exchange partners. The basic form for this type of reaction is shown below.

$$AB + CD \rightarrow CB + AD$$

The reaction is Compound + Compound → Compound + Compound

Some examples of double displacement reactions are shown below.

$$AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$$
$$ZnBr_{2(aq)} + 2\ AgNO_{3(aq)} \rightarrow Zn(NO_3)_{2(aq)} + 2\ AgBr_{(s)}$$
$$H_2SO_{4(aq)} + 2\ NaOH_{(aq)} \rightarrow Na_2SO_{4(aq)} + 2\ H_2O_{(L)}$$

5. Combustion

When organic compounds are burned, they react with oxygen in the air to form carbon dioxide and water. The basic form of the combustion reaction is shown below.

hydrocarbon + oxygen → carbon dioxide + water

Some examples of combustion reactions are shown below.

$$CH_{4(g)} + 2\ O_{2(g)} \rightarrow 2\ H_2O_{(g)} + CO_{2(g)}$$
$$2\ C_2H_{6(g)} + 7\ O_{2(g)} \rightarrow 6\ H_2O_{(g)} + 4\ CO_{2(g)}$$
$$C_3H_{8(g)} + 5\ O_{2(g)} \rightarrow 4\ H_2O_{(g)} + 3\ CO_{2(g)}$$

Exercises

Fill in the reaction type on the line following the balanced equation.

1. $3\ NaBr + H_3PO_4 \rightarrow Na_3PO_4 + 3\ HBr$ ____________________
2. $3\ Ca(OH)_2 + Al_2(SO_4)_3 \rightarrow 3\ CaSO_4 + 2\ Al(OH)_3$ ____________________
3. $3\ Mg + Fe_2O_3 \rightarrow 2\ Fe + 3\ MgO$ ____________________
4. $C_2H_4 + 3\ O_2 \rightarrow 2\ CO_2 + 2\ H_2O$ ____________________
5. $2\ PbSO_4 \rightarrow 2\ PbSO_3 + O_2$ ____________________
6. $2\ NH_3 + 3\ I_2 \rightarrow N_2I_6 + 3\ H_2$ ____________________
7. $H_2O + SO_3 \rightarrow H_2SO_4$ ____________________
8. $2\ NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4$ ____________________
9. $4\ C_5H_9O + 27\ O_2 \rightarrow 20\ CO_2 + 18\ H_2O$ ____________________
10. $Li_3N + 3\ NH_4NO_3 \rightarrow 3\ LiNO_3 + (NH_4)_3N$ ____________________

Chapter 17

Mathematics and Chemical Equations Worksheets

17.1 Lesson 17.1 The Mole Concept and Equations

There are no worksheets for this lesson.

17.2 Lesson 17.2 Mass-Mass Calculations

Stoichiometry Worksheet

CK-12 Foundation Chemistry

Name____________________________________ **Date**_______________

1. How many moles are present in 58.6 grams of lead (II) oxide?

A. 0.113 moles
B. 0.158 moles
C. 0.263 moles
D. 0.300 moles
E. None of these.

2. According to the following balanced equation, how many moles of oxygen can be produced by the complete reaction of 10.0 moles of potassium chlorate, $KClO_3$?

$2\ KClO_3 \rightarrow 2\ KCl + 3\ O_2$

A. 10.0 moles
B. 6.67 moles
C. 15.0 moles
D. 4.00 moles
E. None of these.

3. Balance the following equation and determine how many moles of water will be produced by the complete reaction of 0.600 moles of aluminum hydroxide?

$$\_\_\_\ Al(OH)_3\ +\ \_\_\_\ H_2SO_4\ \rightarrow \_\_\_\ Al_2(SO_4)_3\ +\ \_\_\_\ H_2O$$

A. 1.80 moles
B. 0.200 moles
C. 20.0 moles
D. 0.600 moles
E. None of these.

4. Using the balanced equation, $2\ KClO_3\ \rightarrow\ 2\ KCl\ +\ 3\ O_2$, how many moles of O_2 can be produced by the complete reaction of 100. grams of $KClO_3$?

A. 0.326 moles
B. 0.544 moles
C. 0.816 moles
D. 1.22 moles
E. None of these.

5. If hydrogen is completely reacted with oxygen and produces 180. grams of water, how many grams of hydrogen was consumed? The following equation for the reaction is not yet balanced.

$$\_\_\_\ H_2\ +\ \_\_\_\ O_2\ \rightarrow\ \_\_\_\ H_2O$$

A. 2.02 g
B. 20.2 g
C. 10.1 g
D. 90.0 g
E. 180. g

6. How many grams of calcium can be produced by the complete reaction of 9.35 grams of calcium oxide, according to following, as yet unbalanced, equation?

$\_\_ CaO + \_\_ C \rightarrow \_\_ Ca + \_\_ CO_2$

A. 6.70 g
B. 3.34 g
C. 12.4 g
D. 7.19 g
E. None of these.

7. In a particular reaction, iron (III) oxide and carbon solid reacted to produce iron metal and carbon monoxide. How many grams of iron (III) oxide are required to produce 150. grams of carbon monoxide?

A. 160. g
B. 222 g
C. 286 g
D. 480. g
E. None of these.

8. How many grams of octane, C_8H_{18}, when burned in oxygen gas are required to produce 272 grams of carbon dioxide? The other product is water.

A. 136 g
B. 121 g
C. 100. g
D. 94.6 g
E. 88.2 g

9. How many grams of bromine gas would be liberated when 25.0 grams of gallium bromide were heated and decomposed to form gallium metal and bromine gas?

A. 16.4 g
B. 19.4 g
C. 21.8 g
D. 27.1 g
E. None of these.

10. 2000. g of potassium carbonate react completely with barium phosphate to produce potassium phosphate and barium carbonate. How many grams of barium carbonate will be formed?

A. 1240 g

B. 1680 g
C. 2220 g
D. 2860 g
E. None of these.

17.3 Lesson 17.3 Limiting Reactant

Limiting Reactant Worksheet

CK-12 Foundation Chemistry

Name________________________________ **Date**______________

1. If 2.5 moles of copper and 5.5 moles of silver nitrate are available to react in the following equation, what is the limiting reactant? (The equation is not yet balanced.)

$$\_\_\ Cu\ +\ \_\_\ AgNO_3\ \rightarrow\ \_\_\ Cu(NO_3)_2\ +\ \_\_\ Ag$$

A. copper
B. silver nitrate
C. copper (II) nitrate
D. silver
E. None of these.

2. How many grams of calcium hydroxide will be formed in the following reaction when 4.44 g of calcium oxide and 7.77 g of water are available to react? (The equation is not yet balanced.)

$$\_\_\ CaO\ +\ \_\_\ H_2O\ \rightarrow\ \_\_\ Ca(OH)_2$$

A. 12.2 g
B. 7.77 g
C. 5.86 g
D. 4.11 g
E. None of these.

3. Magnesium undergoes a single replacement reaction with nitric acid, HNO_3. Write the balance equation for the reaction and determine how many grams of hydrogen gas will be formed from the reaction of 3.00 grams of magnesium with 18.00 grams of nitric acid.

A. 0.695 g

B. 0.572 g
C. 0.540 g
D. 0.492 g
E. None of these.

4. Sulfur reacts with oxygen gas to produce sulfur trioxide. Write the balanced equation for the reaction and determine how many grams of sulfur trioxide will be produced when 6.30 g of S and 10.0 g of O_2 are available for reaction.

A. 16.3 g
B. 15.7 g
C. 13.2 g
D. 11.9 g
E. None of these.

5. Some of the acid in acid rain is produced from the following reaction:

$$3\ NO_2\ +\ H_2O\ \rightarrow\ NO\ +\ 2HNO_3$$

A falling raindrop with a mass of 0.0500 gram comes into contact with 0.200 gram of NO_2. What mass of HNO_3 can be produced?

A. 0.183 g
B. 0.250 g
C. 0.350 g
D. 0.146 g
E. None of these.

6. In problem #5, how many grams of the excess reactant remains after the reaction?

A. 0.0415 g
B. 0.0388 g
C. 0.0264 g
D. 0.0239 g
E. None of these.

7. Consider the following reaction: $2\ Al\ +\ 6\ HBr\ \rightarrow\ 2\ AlBr_3\ +\ 3\ H_2$. When 87.0 g of Al is combined with 401 g of HBr, how many grams of H_2 are formed?

A. 3.89 g
B. 5.01 g
C. 7.11 g
D. 12.4 g
E. None of these.

17.4 Lesson 17.4 Percent Yield

Percent Yield Worksheet

CK-12 Foundation Chemistry

Name__________________________________ **Date**______________

1. Methanol, CH_3OH can be produced by the following reaction.

$$2\ H_2\ +\ CO\ \rightarrow\ CH_3OH$$

Assume CO is the limiting reactant and 2.00 mols of CO are used in the reaction. If 0.780 mols of CH_3OH are produced by the reaction, what is the percent yield?

2. Consider the following reaction.

$$3\ Si\ +\ 2\ N_2\ \rightarrow\ Si_3N_4$$

A. What is the theoretical yield of Si_3N_4 from this reaction when 21.45 mols of Si are reacted with excess N_2?

B. If 5.92 mols of Si_3N_4 are actually produced, what is the percent yield?

3. Part of the SO_2 that is introduced into the atmosphere by the combustion of sulfur containing compounds ends up being converted to sulfuric acid, H_2SO_4 by the following reaction.

$$2\ SO_2\ +\ O_2\ +\ 2\ H_2O\ \rightarrow\ 2\ H_2SO_4$$

A. What is the theoretical yield of H_2SO_4 if 100. g of SO_2 is completely consumed?

B. If the actual yield from the reaction in A is 100. g of H_2SO_4, what is the percent yield?

4. Consider the reaction: $4\ FeS_2\ +\ 11\ O_2\ \rightarrow\ 2\ Fe_2O_3\ +\ 8\ SO_2$

A. If 20.0 moles of FeS_2 react with 60.0 moles of O_2, what is the limiting reactant?

B. How many moles of SO_2 are formed?

C. How many moles of the reactant in excess will be left over at the end of the reaction?

D. If the actual yield of SO_2 is 25.0 moles, what is the percent yield?

17.5 Lesson 17.5 Energy Calculations

There are no worksheets for this lesson.

Chapter 18

The Kinetic Molecular Theory Worksheets

18.1 Lesson 18.1 The Three States of Matter

There are no worksheets for this lesson.

18.2 Lesson 18.2 Gases

There are no worksheets for this lesson.

18.3 Lesson 18.3 Gases and Pressure

There are no worksheets for this lesson.

18.4 Lesson 18.4 Gas Laws

The Kinetic Molecular Theory and Gas Laws Worksheet

CK-12 Foundation Chemistry

Name__________________________________ **Date**______________

True or False

________ 1. The mass of a gas is the sum of the masses of the individual molecules.

________ 2. The volume of a gas is the sum of the volumes of the individual molecules.

________ 3. Molecules of different substances move at different velocities when they are at the same temperature.

________ 4. Molecules of the same substance move at the same velocity when they are at the same temperature.

________ 5. Molecules are in motion at all temperatures above absolute zero.

________ 6. Gases are more compressible than solids and liquids because they have more space between the molecules.

________ 7. Molecules of liquid water and molecules of solid water at the same temperature have the same velocity.

________ 8. A liquid has its own shape and volume regardless of the container.

________ 9. All molecules at the same temperature have the same velocity.

________ 10. All molecules at the same temperature have the same average kinetic energy.

________ 11. Molecules of different substances, at the same temperature, exert different forces when they collide with the walls of their container.

________ 12. In a mixture of gases, the partial pressure of a gas has the same ratio to the total pressure as the mole fraction of that gas.

________ 13. Small molecules diffuse faster than large molecules as the same temperature.

________ 14. One mole of He gas will occupy a smaller volume than one mole of UF_6 gas under the same conditions of temperature and pressure.

________ 15. The density of a gas under standard conditions can be found by dividing the molar mass by 22.4 L.

Multiple Choice

16. A sample of gas is held at constant volume. When the temperature of the gas is 100. K, the pressure is 1.00 atm. What must the temperature become in order for the pressure to become 3.00 atm?

A. 27 K
B. 100. K
C. 300. K
D. None of these.
E. Cannot be determined from this data.

17. A sample of gas occupies 100. mL at 1520 Torr and 323 K. What volume will this sample

occupy under standard conditions?

A. 100. mL
B. 116 mL
C. 232 mL
D. 169 mL
E. None of these.

18. 10.0 liters of oxygen gas is held at 3800. mm of Hg pressure and 27.0^oC. What volume will this gas occupy if it is at -23.0^oC and 380. mm of Hg pressure?

A. 8.33 L
B. 83.3 L
C. 833 L
D. 50.0 L
E. None of these.

19. 1.00 g of H_2 gas is placed in a flask with 1.00 g of He gas. The total pressure in the flask is 900. Torr. What is the partial pressure of the H_2?

A. 100. Torr
B. 300. Torr
C. 450. Torr
D. 600. Torr
E. 800. Torr

20. 10.0 atm of pressure is applied to 0.250 mole of methane gas. What must the temperature be if the volume is to be 1400. mL?

A. 409 K
B. 682 K
C. 955 K
D. 0 K
E. None of these.

21. Given a sample of gas at 1.0 atm pressure, what would the pressure become if the amount of gas is doubled, the volume decreased to half, and the absolute temperature quadrupled?

A. 1.0 atm
B. 2.0 atm

C. 4.0 atm
D. 8.0 atm
E. 16 atm

22. How many mols of gas are required to fill a 1.0 liter container to 5.00 atm pressure at 27.0^oC?

A. 0.13 moles
B. 0.20 moles
C. 0.29 moles
D. 0.38 moles
E. None of these.

23. What is the molar mass of a gas if 0.500 g of it occupies 0.250 liters at 1.00 atm and $100.^oC$?

A. 32.0 g/mol
B. 44.0 g/mol
C. 61.3 g/mol
D. 77.2 g/mol
E. 104 g/mol

24. 10.0 liters of gas at 27.0^oC and 0.15 atm has a mass of 10.0 grams. What is the molar mass of the gas?

A. 40. g/mol
B. 80. g/mol
C. 100. g/mol
D. 120 g/mol
E. 164 g/mol

25. What is the mass of 100. L of Br_2 gas under standard conditions?

A. 22.4 g
B. 357 g
C. 560. g
D. 714 g
E. Insufficient data to determine.

18.5 Lesson 18.5 Universal Gas Law

There are no worksheets for this lesson.

18.6 Lesson 18.6 Molar Volume

There are no worksheets for this lesson.

18.7 Lesson 18.7 Stoichiometry Involving Gases

There are no worksheets for this lesson.

Chapter 19

The Liquid State Worksheets

19.1 Lesson 19.1 The Properties of Liquids

There are no worksheets for this lesson.

19.2 Lesson 19.2 Forces of Attraction

There are no worksheets for this lesson.

19.3 Lesson 19.3 Vapor Pressure

There are no worksheets for this lesson.

19.4 Lesson 19.4 Boiling Point

There are no worksheets for this lesson.

19.5 Lesson 19.5 Heat of Vaporization

There are no worksheets for this lesson.

Chapter 20

The Solid State Worksheets-HSC

20.1 Lesson 20.1 The Molecular Arrangement in Solids Controls Solid Characteristics

There are no worksheets for this lesson.

20.2 Lesson 20.2 Melting

Heat Transfer Worksheet

CK-12 Foundation Chemistry

Name__________________________________ **Date**______________

Temperature is defined as the **average** kinetic energy of all the molecules in a body, while ***heat*** is defined as the **total** kinetic energy of all the molecules in a body. A sample of matter will contain kinetic energy due to the motion of its molecules, and it also contains potential energy due to its phase (solid, liquid, gas). When two objects come into contact with each other, heat always flows from the one with higher temperature to the one with lower temperature. This transfer of KE is accomplished by the collision of molecules and continues until the two objects are at the same temperature.

Every chemical change and many physical changes involve the gain or loss of energy. In most cases, this energy gain or loss occurs in the form of heat, but light and electricity are also possible. Heat gains and losses are measured in units called **Joules**. It requires 4.18 Joules of energy to raise the temperature of 1.00 gram of water by 1.00^oC. When heat energy is added to a substance, it produces one or both of the following effects: 1) it may increase the temperature of the object, which means it increases the average kinetic energy of the molecules or, 2) it may cause a phase change in that substance, which means it increases the potential energy of the substance.

When heat is absorbed by a substance as kinetic energy, the temperature of the substance increases because temperature is a measure of the average kinetic energy of the molecules of the substance. Different substances have a different amount of increase in temperature when they absorb the same amount of energy. The quantity of heat 1.00 gram of the substance must absorb to raise its temperature by 1.00^{o}C is called the ***specific heat*** of the substance. The symbol, **C**, is often used for specific heat. The specific heat of water is 4.18 $J/g \cdot^o C$. This means that 1.00 gram of water requires 4.18 J of heat to raise its temperature by 1.00^{o}C. The specific heats of most substances are considerably less than that of water.

Table 20.1: **Specific Heat of Various Substances**

Substance	Specific Heat
Aluminum, Al	0.900 $J/g \cdot^o C$
Copper, Cu	0.386 $J/g \cdot^o C$
Gold, Au	0.126 $J/g \cdot^o C$
Silver, Ag	0.235 $J/g \cdot^o C$
Ethanol, C_2H_5OH	2.40 $J/g \cdot^o C$
Butane, C_4H_{10}	2.34 $J/g \cdot^o C$
Water, H_2O	4.18 $J/g \cdot^o C$

Energy is also absorbed or given off by substances when they undergo a phase change. The energy gained or lost during a phase change is potential energy. This energy gain or loss does not change the temperature of the substance. When substances undergo a phase change, the average distance between the molecules changes and this requires an input or output of potential energy. When a substance changes from solid to liquid, the energy that must be absorbed is called **heat of melting**. The reverse process, changing from liquid to solid, gives off exactly the same amount of energy but for this phase change, the amount of energy is known as the **heat of fusion**. The phase change from liquid to gas requires an input of the **heat of vaporization**. The reverse process, gas condensing to liquid, gives off the same amount of potential energy but it is called the **heat of condensation**. Like specific heat, each substance has its own heat of melting and heat of vaporization.

Table 20.2: **Thermodynamic Data of Various Substances**

Substance	Specific Heat	Heat of Fusion, ΔH_{fusion}	Heat of Vaporization, ΔH_{vap}
Aluminum, Al	0.900 $J/g \cdot^o C$	400. J/g	10,900 J/g
Copper, Cu	0.386 $J/g \cdot^o C$	205 J/g	5,069 J/g
Gold, Au	0.126 $J/g \cdot^o C$	64.5 J/g	1,578 J/g
Silver, Ag	0.235 $J/g \cdot^o C$	111 J/g	2,320 J/g

Table 20.2: (continued)

Substance	Specific Heat	Heat of Fusion, ΔH_{fusion}	Heat of Vaporization, ΔH_{vap}
Ethanol, C_2H_5OH	$2.40\ J/g \cdot^o C$	109 J/g	841 J/g
Butane, C_4H_{10}	$2.34\ J/g \cdot^o C$	80.1 J/g	385 J/g
Water, H_2O	$4.18\ J/g \cdot^o C$	334 J/g	2,260 J/g

The energy absorbed or given off by a substance during a temperature change (with no phase change) can be calculated with the equation, $Q = mC\Delta t$, where Q is the amount of heat in Joules, m is the mass in grams, C is the specific heat, and Δt is the temperature change.

Example: How many Joules are given off when 52.5 g of water cools from 67.5^{o}C to 23.2^{o}C?

Solution: $Q = mC\Delta t = (52.5g)(4.18J/g \cdot^o C)(44.3^oC) = 9720$ J

The specific heat is taken from the table above and the units cancel appropriately to yield Joules.

Example: If 4490 J of heat are added to 50.0 g of solid silver at 25.0^{o}C, what would the final temperature be?

Solution: $Q = mC\Delta t$ so $\Delta t = \frac{Q}{mC}$

$\Delta t = \frac{Q}{mC} = \frac{4490\ J}{(50.0\ g)(0.235J/g \cdot^o C)} = 382^oC$
Final temperature = initial temperature + $\Delta t = 25^oC + 382^oC = 407^oC$

The energy absorbed or given off by a substance during a phase change (with no temperature change) can be calculated with the equations, $Q = m\Delta H_{fusion}$ or $Q = m\Delta H_{vap}$, where Q is the amount of heat in Joules, m is the mass of the substance in grams, and ΔH_{fusion} or ΔH_{vap} is the heat of fusion or vaporization.

Example: How many Joules are required to melt 17.7 grams of solid aluminum at its normal melting point with no temperature change?

Solution: $Q = m\Delta H_{fusion}$ = (17.7 g)(400. J/g) = 7080 J

When heat is added to a substance such that the substance undergoes both a temperature change and a phase change, the problem is solved separately for each process. For example, if sufficient heat is added to solid water (ice) at -20^oC to raise the temperature and cause the necessary phase changes, the solid water will go through five processes; 1) the temperature of the ice will be raised to the melting point, 2) the solid water will be melted, 3) the temperature of the liquid water will be raised to the boiling point, 3) the liquid will be vaporized, and 5) the temperature of the gaseous water will be raised to the final temperature. To do calculations for this entire process, many bits of thermodynamic data will be required.

We would need to know the specific heat of solid water (not the same as liquid water), the heat of fusion for water, the specific heat of liquid water, the heat of vaporization, and the specific heat of gaseous water.

Example: Calculate the heat necessary to raise 100. g of iron at 25.0^oC to liquid iron at $2000.^oC$. The necessary thermodynamic data are: melting point of iron = $1540.^oC$, specific heat of solid iron = 0.450 $J/g \cdot^o C$, specific heat of liquid iron = 0.770 $J/g \cdot^o C$, heat of fusion of iron = 280. J/g.

Solution:

Step 1: Heat the solid iron from 25.0^oC to its melting point at $1540.^oC$ $(\Delta t = 1515^oC)$.

$$Q = mC\Delta t = (100.g)(0.450J/g \cdot^o C)(1515^oC) = 68{,}200 \text{ J}$$

Step 2: Melt the solid iron to liquid.

$$Q = m\Delta H_{fusion} = (100.\text{ g})(280.\text{ J/g}) = 28{,}000 \text{ J}$$

Step 3: Heat the liquid iron from the melting point $(1540.^oC)$ to the final temperature $(2000.^oC)$ $\Delta t = 460.^oC$.

$$Q = mC\Delta t = (100.\ g)(0.770\ J/g \cdot^o C)(460^oC) = 35{,}400 \text{ J}$$

Step 4: Add up the heat added for each step to get the total.

$$Q_{TOTAL} = 68{,}200 \text{ J} + 28{,}000 \text{ J} + 35{,}400 \text{ J} = 131{,}600 \text{ J} = 131.6 \text{ kJ} = 132 \text{ kJ}$$

Example: Calculate the heat necessary to raise 40.00 g of ice at -50.0^oC to water vapor at 180.0^oC.

Necessary Thermodynamic Data

- $C_{ice} = 2.09\ J/g \cdot^o C$
- $C_{water} = 4.18\ J/g \cdot^o C$
- $C_{water\ vapor} = 2.01\ J/g \cdot^o C$
- Melting Point = 0^oC
- Boiling Point = $100.^oC$
- $\Delta H_{fusion} = 334\ J/g$
- $\Delta H_{vap} = 2260\ J/g$

Solution:

Step 1: Raise the temperature of the ice from -50.0^oC to the melting point 0^oC.

$$Q = mC\Delta t = (40.00\ g)(2.09\ J/g\cdot^{o}C)(50.00^{o}C) = 4,180\ J$$

Step 2: Melt the ice to liquid water.

$$Q = m\ \Delta H)fusion = (40.00\ g)(334\ J/g) = 13,360\ J$$

Step 3: Raise the temperature of the liquid water from the m.p. to the b.p. ($\Delta t = 100.^{o}$C).

$$Q = mC\Delta t = (40.00g)(4.18\ J/g\cdot^{o}C)(100.^{o}C) = 16,720\ J$$

Step 4: Vaporize the liquid water.

$$Q = m\Delta H_{vap} = (40.00\ g)(2260\ J/g) = 90,400\ J$$

Step 5: Raise the temperature of the gaseous water from the b.p. to the final temperature ($\Delta t = 100.^{o}$C).

$$Q = mC\Delta t = (40.00\ g)(2.01\ J/g\cdot^{o}C)(80.^{o}C) = 6,400\ J$$

Step 6: Add up the results of each step.

$$Q_{TOTAL} = 4180 + 13360 + 16720 + 90400 + 6400 = 131{,}000\text{ J} = 131\text{ kJ}$$

Questions and Exercises

The thermodynamic data necessary for these problems can be found in the preceding pages.

1. Assuming no phase change occurs, what happens to the temperature of a substance when it absorbs heat?

2. What happens when two objects at different temperatures are brought into contact?

3. How many Joules of heat must be added to 5000. g of water to change its temperature from $20.^{o}C$ to $80.^{o}C$?

4. If 500. g of water at $25.^{o}C$ loses 10,000. J of heat, what will its final temperature be?

5. What does the temperature of an object actually measure?

6. At what temperature do molecules have zero kinetic energy?

7. Describe a situation where heat can enter a body without causing an increase in temperature?

8. How much heat is released when 44.8 g of solid gold are cooled from $80.^oC$ to 62^oC?

9. How much heat is needed to melt 25.0 g of silver at its normal melting point?

10. How much heat is absorbed when 24.5 g of ice at -10.0^oC is warmed to liquid water at 42.5^oC?

11. Calculate the amount of heat necessary to raise 45.0 g of cesium metal from 24.0^oC to 880.0^oC. Use the data given below.

Necessary Thermodynamic Data

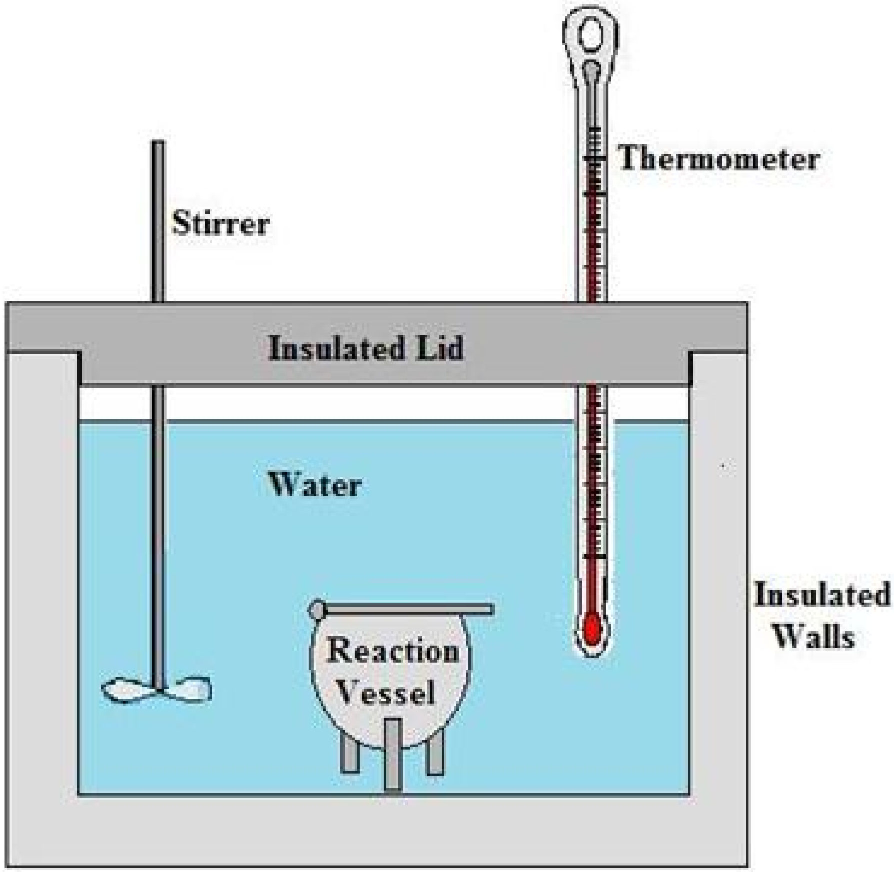
- $C_{solid\ Cs} = 0.251\ J/g \cdot^o C$
- $C_{liquid\ Cs} = 0.209\ J/g \cdot^o C$
- $C_{gaseous\ Cs} = 0.167\ J/g \cdot^o C$
- Melting Point = 29.0^oC
- Boiling Point = 690.0^oC
- $\Delta H_{fusion} = 16.3\ J/g$
- $\Delta H_{vap} = 669\ J/g$

Calorimetry Worksheet

CK-12 Foundation Chemistry

Name____________________________________ **Date**_______________

The laboratory process for measuring the amount of heat gained or during a chemical reaction or other energy exchange involves the use of an instrument called a ***calorimeter***. The basic idea of a calorimeter is sketched below.



The calorimeter has an insulated container to eliminate heat exchange with the outside, a reaction vessel where the reaction to be measured will occur, a quantity of water to absorb from or give up to the heat from the reaction, a thermometer to accurately measure the temperature of the water, and a stirring rod to assure that all the water is the same temperature. Since the heat will come out of or go into the reaction vessel, it is likely that the water touching the vessel would be warmer or colder than the remainder of the water. The stirring rod is used to keep the water circulating and thus all the water will be the same temperature.

At an earlier time, the unit chemists used to measure heat was the calorie. The words *calorimeter* (the name of the instrument) and *calorimetry* (the name of the process) came from the unit, calorie. When scientists decided to use the same units in all branches of science, chemists changed their unit for heat (and all other forms of energy) from calories to **Joules**. The old unit *calorie* is equal to 4.18 Joules. Even though chemists don't use the calorie unit anymore, the words calorimeter and calorimetry remain with us.

Extremely accurate calorimeters are calibrated before each use. A precisely known amount of heat is added to the calorimeter and the temperature change is noted. In this way, the scientist can determine exactly how much heat is required to raise the temperature of the calorimeter by 1.00^{o}C. This allows the scientist to measure not only the heat absorbed by the water in the calorimeter but also the heat absorbed by the reaction vessel, the stirrer, the thermometer, and the inside walls of the calorimeter. For a less precise calorimeter, the scientist assumes all the heat added to the calorimeter is absorbed by the water, ignoring the small amount of absorbed by other components.

To use a calorimeter of the less precise type, the scientist measures the amount of water inside very carefully, measures the temperature of the water before the reaction begins, and measures the maximum or minimum temperature the water reaches after the reaction. Since it is assumed that all the heat absorbed or given off by the reaction went into the water, knowing the amount of water and the temperature change of the water, the scientist can then calculate the amount of heat that the water absorbed or gave off, and that is the heat input or output by the reaction. The equation used to calculate the change in heat content of the water is the same one used before, namely $Q = mC\Delta t$.

Example: How much heat was absorbed by 1000. g of water in a calorimeter if the temperature of the water was raised from 23.5^{o}C to 44.8^{o}C?

Solution: $Q = mC\Delta t = (1000.\ g)(4.18\ J/g \cdot^o C)(21.3^o C) = 89,000\ J = 89\ kJ$

Example: How much heat was absorbed by 500. g of water in a calorimeter if the water temperature changed from 25.0^{o}C to 17.2^{o}C?

Solution: $Q = mC\Delta t = mC(t_2 - t_1) = (500.\ g)(4.18\ J/g \cdot^o C)(17.2^o C - 25.0^o C)$

$$Q = (500.\ g)(4.18\ J/g \cdot^o C)(-7.8^o C) = -16,300\ J = -16.3\ kJ$$

The negative sign of this result indicates the water in the calorimeter lost heat to the reaction, so the reaction was endothermic.

Calorimeters are used by scientists to measure many types of heat exchanges, such as finding the specific heat of substances, the heat value of fuels, and the heat of chemical reactions. Coal mined in different areas is of different quality. When coal is purchased by users from producers, the price paid is based not only on the mass of coal purchased but also on the amount of heat produced by burning a unit quantity of the coal. When a trainload of coal is delivered, there is a scientist on hand to take samples of the coal and burn them in a calorimeter to determine the average Joules/gram of heat produced by that particular load of coal and the price is adjusted accordingly.

Physicists use calorimeters to determine the specific heat of substances. Suppose we wished to determine the specific heat of brass. We use a calorimeter containing 250. g of water at 25.0^{o}C and into it we place a 100. g piece of brass whose temperature we have raised to 91.0^{o}C. When the heat transfer is complete, the final temperature of the water and the piece of brass are 27.3^{o}C. (Since they are in contact, they must eventually reach the same temperature.) The amount of heat lost by the brass will equal the amount of heat gained by the water. We can use the following equation to find the specific heat of the brass.

$$m_{water}C_{water}\Delta t_{water} = -m_{brass}C_{brass}\Delta t_{brass}$$

The negative sign on the brass side of the equation is present because the heat is being gained by the water and lost by the brass. Therefore, the Δt for the water will be positive but the Δt for the brass will be negative. The heat calculated on the two sides of the equation can only be equal if we change the sign of one of them.

Substituting from the problem yields

$(250.\ g)((4.18\ J/g \cdot^o C)(27.3^oC - 25.0^oC) = -(100.\ g)((x\ J/g \cdot^o C)(27.3^oC - 91.0^oC)$.

Solving for x yields, $x = 0.377\ J/g \cdot^o$C

The heat of reaction, ΔH, for a chemical reaction is commonly expressed in J/mole or kJ/mole of product. It is also standard to express the ΔH for an endothermic reaction as a positive number (the reaction is gaining energy) and the ΔH for an exothermic reaction as a negative number (the reaction is losing energy). For the reaction between hydrochloric acid and sodium hydroxide, $HCl + NaOH \rightarrow NaCl + H_2O$, the amount of materials necessary to produce one mole of water would be too large for the calorimeter. That is, we can't actually use molar quantities of these materials. Therefore, we use a fraction of a mole and calculate what the heat transfer would have been for an entire mole.

Example: Suppose we carry out the above reaction in a calorimeter. We use 4.00 g of $NaOH$ with excess HCl solution. That means the $NaOH$ will be the limiting reactant. The 4.00 g of $NaOH$ is 0.100 mole and will produce 0.100 mole of H_2O. We use 250. g of water in the calorimeter and the temperature change during the reaction is from 22.4^{o}C to 28.4^{o}C. Calculate the the heat of reaction for the reaction between hydrochloric acid and sodium hydroxide.

Solution: We can calculate the heat absorbed by the water in the calorimeter in the usual way.

$$Q = (250.\ g)(4.18\ J/g \cdot^o C)(6.0^oC) = 6,270\ J = 6.27\ kJ$$

We can then calculate the ΔH for the reaction by dividing the heat transferred to the water in the calorimeter by the moles of water produced during the reaction. Since the temperature of the water in the calorimeter increased, we know this is an exothermic reaction and therefore, we provide for making the ΔH a negative value . . . required by the definition of ΔH. We can use the following equation.

$$\Delta H = \frac{-\ \Delta\ Q}{moles\ product} = \frac{-6.27\ kJ}{0.100\ mol} = \text{- 62.7 kJ/mol}$$

Exercises

1. How much heat is absorbed by 1.00 g of water when its temperature changes from 20.0^{o}C to 25.0^{o}C.

2. What was the heat transfer if 800. g of water in a calorimeter underwent a temperature change from 25.0^{o}C to 22.0^{o}C?

3. A 7.38 g sample of coal is burned in a calorimeter and raises the temperature of 1000. g of water in calorimeter form 22.0^{o}C to 68.8^{o}C. What is the heat content of this coal in J/g?

4. A reaction that formed 10.0 g of magnesium oxide, MgO, was carried out in a calorimeter. The calorimeter contained 800. g of water and the temperature of the water increased 44.6^{o}C. What was the ΔH for this reaction in kJ/mol?

5. Using the ΔH you found in problem #4, suppose you had carried out exactly this same reaction except that you had used a calorimeter than container 250. g of water instead of 800. g of water. What would the temperature change have been? Give a reason that this reaction wouldn't be carried out with 250. g of water.

20.3 Lesson 20.3 Types of Forces of Attraction for Solids

There are no worksheets for this lesson.

20.4 Lesson 20.4 Phase Diagrams

There are no worksheets for this lesson.

Chapter 21

The Solution Process Worksheets

21.1 Lesson 21.1 The Solution Process

There are no worksheets for this lesson.

21.2 Lesson 21.2 Why Solutions Occur

There are no worksheets for this lesson.

21.3 Lesson 21.3 Solution Terminology

There are no worksheets for this lesson.

21.4 Lesson 21.4 Measuring Concentration

Concentration by Percent Mass Worksheet

CK-12 Foundation Chemistry

Name____________________________________ **Date**_______________

The definition of percent mass concentration is the ratio of the mass of solute divided by the total mass of the solution and multiplied by 100 to convert to a percentage.

$$percent\ by\ mass\ =\ \frac{mass\ of\ solute}{mass\ of\ solution}\ x\ 100$$

Example: What is the percent concentration by mass of a solution formed by dissolving 100. grams of ethanol, C_2H_5OH, in 100. grams of water?

Solution: percent by mass $= \frac{mass\ of\ solute}{mass\ of\ solution}\ x\ 100\ =\ \frac{100.\ g}{200.\ g}\ x\ 100\ =\ 50.0\%$

Example: If the density of a 10.0% by mass KNO_3 solution in water is 1.19 g/mL, how many grams of KNO_3 are present in 100. mL of the solution?

Solution: We can multiply the volume times the density to the mass of the 100. mL of solution and then take 10.0% of the mass of the solution to get the mass of the potassium nitrate.

grams of solution = (100. mL)(1.19 g/mL) = 119 grams
grams of KNO_3 = (0.10)(119 grams) = 11.9 grams

Exercises

1. If 30.0 grams of $AgNO_3$ are dissolved in 275 grams of water, what is the concentration of the silver nitrate by mass percent?

2. How many grams of MgF_2 are present in 100.0 g of a 20.0% MgF_2 in water solution?

3. How many grams of water are present in the solution in question #2?

4. The density of a 30.0% by mass solution of $NaOH$ in water is 1.33 g/mL. How many grams of $NaOH$ are required to prepare 500. mL of this solution?

5. The density of pure water is 1.00 g/mL. What is the concentration gy percent mass of a solution prepared by dissolving 85.0 grams of $NaOH$ in 750. mL of water?

6. A solution is prepared by dissolving 66.0 grams of acetone, C_3H_6O, in 146.0 grams of water. The density of the solution is 0.926 g/mL. What is the percent concentration of acetone by mass?

7. A 35.4% solution of H_3PO_4 in water has a density of 1.20 g/mL. How many grams of phosphoric acid are present in 300. mL of this solution?

Mole Fraction and Molality Worksheet

CK-12 Foundation Chemistry

Name____________________________________ **Date**_______________

Mole Fraction

The definition of mole fraction is the ratio of the moles of solute divided by the total moles of the solution.

$$mole\ fraction\ =\ \frac{moles\ of\ solute}{moles\ of\ solution}$$

Example: What is the mole fraction of ethanol in a solution prepared by dissolving 100. g of ethanol, C_2H_5OH, in 100. g of water?

Solution:

moles ethanol $= \frac{100.\ g}{46.0\ g/mol} = 2.17$ moles
moles water $= \frac{100.\ g}{18.0\ g/mol} = 5.56$ moles
mole fraction of ethanol $= \frac{2.17\ mols}{7.73\ mols} = 0.281$

Molality

The definition of molality is the ratio of the moles of solute divided by the kilograms of solvent.

$$molality\ =\ \frac{moles\ of\ solute}{kilograms\ of\ solvent}$$

Example: What is the molality of a solution prepared by dissolving 100. g of ethanol, C_2H_5OH, in 100. g of water?

moles ethanol $= \frac{100.\ g}{46.0\ g/mol} = 2.17$ moles
molality of ethanol $= \frac{2.17\ mols}{0.100\ kg} = 21.7$ m

Example: A 35.4% solution of H_3PO_4 in water has a density of 1.20 g/mL. What is the mole fraction of H_3PO_4 in this solution and what is the molality?

Solution: We can choose a sample volume of this solution and get the mass of it by multiplying the volume times the density. Suppose we choose a 1.00 L sample.

mass of solution = (1000. mL)(1.20 g/mL) = 1200. grams
mass of H_3PO_4 in the solution = (0.354)(1200. grams) = 425 grams
mass of H_2O = 1200. grams - 425 grams = 775 grams

moles $H_3PO_4 = \frac{425\ g}{98.0\ g/mol} = 4.34$ moles
moles $H_2O\ =\ \frac{775\ g}{18.0\ g/mol} = 43.1$ moles
mole fraction of $H_3PO_4\ =\ \frac{4.34\ mol}{47.4\ mol} = 0.0916$
molality $= \frac{4.34\ mol}{0.775\ kg} = 5.60$ m

Exercises

1. What is the mole fraction of MgF_2 in a solution that has 20.0 g of MgF_2 dissolved in 80.0 grams of water?

2. What is the molality of the solution in question 1?

3. The density of a 30.0% by mass solution of $NaOH$ in water is 1.33 g/mL. What is the mole fraction of $NaOH$ in this solution?

4. What is the molality of the solution in problem 3?

5. What is the molality of a solution prepared by dissolving 4.00 g of $NaCl$ in 100. g of water?

6. How many grams of beryllium chloride would you need to add to 125 g of water to make a 0.500 m solution?

7. What would be the mole fraction of $BeCl_2$ in the solution in problem 6?

8. A solution is prepared by dissolving 66.0 g of acetone, C_3H_6O, in 146.0 g of water. The density of the solution is 0.926 g/mL. What is the molality of this solution?

9. What is the mole fraction of acetone in the solution in problem 8?

Molarity Worksheet

CK-12 Foundation Chemistry

Name____________________________________ **Date**_______________

The definition of molarity is the ratio of the mols of solute divided by the volume of the solution.

$$molarity = \frac{moles\ of\ solute}{liters\ of\ solution}$$

Example: What is the molarity of a solution prepared by dissolving 60.0 grams of NaOH in sufficient water to produce 2.00 liters of solution?

Solution:

$$\text{moles } NaOH = \frac{60.0\ g}{40.0\ g/mol} = 1.50 \text{ moles}$$
$$\text{molarity} = \frac{1.50\ mol}{2.00\ L} = 0.750 \text{ M}$$

Example: What volume of 0.750 M $NaOH$ solution will contain 10.0 gram of $NaOH$?

$$\text{moles } NaOH = \frac{10.0\ g}{40.0\ g/mol} = 0.250 \text{ moles}$$
$$\text{volume} = \frac{mol}{M} = \frac{0.250\ mol}{0.750\ mol/L} = 0.333 \text{ L}$$

Exercises

1. What is the molarity of a solution in which 4.50 g of $NaNO_3$ is dissolved in 265 mL of solution?

2. How many grams of ammonia, NH_3 are present in 5.0 L of 0.100 M solution?

3. How many milliliters of 0.200 M $NaOH$ solution is necessary to contain 6.00 grams of $NaOH$?

4. How many liters of 0.500 M CaF_2 solution is required to contain 78.0 g of CaF_2?

5. What mass of ammonium phosphate is needed to make 100. mL of 0.500 M $(NH_4)_3PO_4$ solution?

6. What is the molarity of a solution prepared by dissolving 198 g of $BaBr_2$ in 2.00 liters of solution?

7. How many grams of glycerine, $C_3H_8O_3$, are needed to make 100. mL of 2.60 M solution?

8. A test tube contains 10.0 mL of 3.00 M $CaCO_3$ solution. How many grams of calcium carbonate are in the tube?

21.5 Lesson 21.5 Solubility Graphs

There are no worksheets for this lesson.

21.6 Lesson 21.6 Factors Affecting Solubility

There are no worksheets for this lesson.

21.7 Lesson 21.7 Colligative Properties

Dilution Worksheet

CK-12 Foundation Chemistry

Name____________________________________ **Date**_______________

The process of dilution involves increasing the amount of solvent in a solution without changing the amount of solute. For example, you could dilute 50. mL of 0.250 M HCl solution by placing the solution in a 100. mL graduated cylinder and adding water until the solution reached the 100. mL line in the graduate. The original solution contained 0.0125 moles of HCl before it was diluted and therefore, it also contains 0.0125 moles of HCl after the dilution. In the process of dilution, the amount of solute never changes. The amount of solvent, the total volume of the solution, and the concentration change but the amount of solute remains the same.

For a solution whose concentration is expressed in molarity, the moles of solute can be calculated by multiplying the volume in liters times the molarity.

moles solute = (molarity)(liters)

For the moles of solute in the original solution, $moles_{initial} = molarity_{initial}\ x\ liters_{initial}$ or $mols_i = M_i\ x\ V_i$. After the solution has been diluted, the moles in the final solution can be calculated with $mols_f = M_f\ x\ V_f$. Since the mols do not change during dilution,

$$mols_i = mols_f \quad and \quad \mathbf{M_i\ x\ V_i = M_f\ x\ V_f}.$$

In the dilution problems you will be given, for the most part, three of the four variables or ways to find three of the four variables and you will asked to calculate the fourth variable.

Example: How many milliliters of 6.00 M $NaOH$ solution are necessary to prepare 300. mL of 1.20 M $NaOH$ solution?

Solution:

$$(M_i)(V_i) = (M_f)(V_f)$$
$$V_i = \frac{(M_f)(V_f)}{(M_i)} = \frac{(1.20\ M)(0.300\ L)}{(6.00\ M)} = 0.0600\ L = 60.0\ mL$$

Exercises

1. 200. mL of 3.00 M $NaCl$ solution is diluted to a final volume of 500. mL. What is the concentration of the final solution?

2. 100. mL of concentrated hydrochloric acid was diluted to 1.20 liters of 1.00 M solution. What was the concentration of the original concentrated solution?

3. What volume of 6.00 M $NaOH$ is needed to prepare 250. mL of 0.600 M $NaOH$?

4. If 25.0 mL of 16.0 M HNO_3 is diluted to 500. mL, what is the final concentration?

5. To what volume must you dilute 10.0 mL of 6.00 M H_2SO_4 to produce a solution that is 1.00 M H_2SO_4?

6. Solution A is 5.00 mL of 12.0 M HCl. Solution B is prepared by diluting solution A to a new volume of 100. mL. Solution C is produced by taking 5.00 mL of solution B and diluting it to 100. mL. What is the molarity of solution C?

Colligative Properties: Solution Vapor Pressure Worksheet

Colligative properties are those properties of a solution that depend on the number of particles of solute present in the solution, and not on the chemistry nor the mass of the particles. That is, the chemical behavior and the molar masses of urea, $(NH_2)_2CO$, and glucose, $C_6H_{12}O_6$, are very different, but the colligative properties of a 1.0 M solution of urea will be exactly the same as the colligative properties of a 1.0 M solution of glucose.

The colligative properties of solutions include vapor pressure lowering, boiling point elevation, freezing point depression, and changes in osmotic pressure. The changes in these properties are dependent entirely on the concentration of particles of solute in the solution. It must be noted that ionic solutes dissociate when dissolved in water and therefore, add more particles to the solution than a substance that does not dissociate in water.

Vapor Pressure Lowering

The vapor pressure of a solution can be calculated from the individual vapor pressures of the components (solute and solvent) and the mole fractions of each component. Raoult's Law is an expression of the relationship.

$$\textbf{Vapor Pressure}_{\textbf{solution}} = (\textbf{X}_{\textbf{mol fraction solvent}})(\textbf{Vapor Pressure}_{\textbf{solvent}}) + (\textbf{X}_{\textbf{mol fraction solute}})(\textbf{Vapor Pressure}_{\textbf{solute}})$$

Example: What is the vapor pressure, at 25°C, of a solution produced by dissolving 50.0 of acetone, C_3H_6O, in 50.0 grams of water? The vapor pressure of pure acetone at 25°C is 230. mm of Hg and the vapor pressure of pure water at 25°C is 23.7 mm of Hg.

Solution: 50.0 g of acetone is 0.86 moles and 50.0 g of water is 2.78 moles.

Therefore, the mole fractions in this solution are 0.236 acetone and 0.764 water.
VP_{SOLUTION} = (0.764)(23.7 mm of Hg) + (0.236)(230. mm of Hg) = 18.1 mm of Hg + 54.3 mm of Hg = 72.4 mm of Hg

In this case, the vapor pressure of the solution is higher than the vapor pressure of the solvent. That is due to the fact that acetone is a volatile (weak intermolecular forces of attraction) and therefore, evaporates readily. When we refer to ***vapor pressure lowering***, we are referring to solutions in which the solute is ***non-volatile***. When the solute is a solid, it can be generally be assumed that the solute is non-volatile.

Suppose we are making a solution of glucose in water. Glucose is a non-volatile, solid solute whose vapor pressure at room conditions is so small that it is negligible compared to the vapor pressure of water. When we substitute the values for a glucose solution into Raoult's Law, the second term (the one for the solute) is essentially zero because the vapor pressure of the pure solute is essentially zero.

$$\textbf{Vapor Pressure}_{\textbf{Solution}} = (\textbf{X}_{\textbf{Mol fraction solvent}})(\textbf{Vapor Pressure}_{\textbf{Solvent}}) + (\textbf{X}_{\textbf{Mol fraction solute}})(\textbf{Vapor Pressure}_{\textbf{Solute}})$$

If the second term in this equation, $(\mathrm{X}_{\text{Mol fraction solute}})(\text{Vapor Pressure}_{\text{Solute}})$, becomes zero, then for a solution with a non-volatile solute, Raoult's Law becomes:

$$\textbf{Vapor Pressure}_{\textbf{Solution}} = (\mathbf{X}_{\textbf{Mol fraction solvent}})(\textbf{Vapor Pressure}_{\textbf{Solvent}})$$

This is Raoult's Law for solutions whose solute is a non-volatile.

$$\mathbf{VP}_{\textbf{Solution}} = (\mathbf{X}_{\textbf{Solvent}})(\mathbf{VP}_{\textbf{Solvent}})$$

Example: What is the vapor pressure, at 25^oC, of a solution produced by dissolving 50.0 of glucose, 25^oC, in 50.0 grams of water? Glucose is non-volatile and the vapor pressure of pure water at 25^oC is 23.7 mm of Hg.

Solution: 50.0 g of water is 2.78 moles and 50.0 g of glucose is 0.278 moles.

Therefore, the mole fraction of water in this solution is 0.909. We do not need to calculate the mole fraction of glucose because it isn't needed in Raoult's Law for non-volatile solutes.

$\mathrm{VP}_{\text{Solution}} = (\mathrm{X}_{\text{Solvent}})(\mathrm{VP}_{\text{Solvent}} = (0.909)(23.7 \text{ mm of Hg}) = 21.5 \text{ mm of Hg}$

In this case, and in all cases of non-volatile solutes, the vapor pressure of the solution is less than the vapor pressure of the pure solvent.

Exercises

1. If 25.0 grams of sodium chloride is added to 500. grams of water at 25^oC, what will be the vapor pressure of the resulting solution in kPa? The vapor pressure of pure water at 25^oC is 3.17 kPa.

2. 125 g of the non-volatile solute glucose, $C_6H_{12}O_6$, is dissolved in 125 g of water at 25^oC. IF the vapor pressure of water at 25^oC is 23.7 Torr, what is the vapor pressure of the solution?

3. Glycerin, $C_3H_8O_3$, is a non-volatile, non-electrolyte solute. If 53.6 g of glycerin is dissolved in 133.7 g of ethanol at $40.^oC$, C_2H_5OH, what is the vapor pressure of the solution? The vapor pressure of pure ethanol is 113 Torr at $40.^oC$.

4. The vapor pressure of hexane, C_6H_{14}, at 60.0^oC is 573 Torr. The vapor pressure of benzene at the same temperature is 391 Torr. What will be the vapor pressure of a solution of 58.9 g of hexane with 44.0 g of benzene?

Colligative Properties: B.P. Elevation and M.P. Depression Worksheet

When a non-volatile, solid solute is added to a solvent, the boiling point of the solution will be higher than the boiling point of the solvent, and the melting point of the solution will be lower than the melting point of the solvent. The size of the boiling point elevation and the melting point depression are colligative properties, that is, they are dependent not on the chemistry of the solute but only on the number of solute particles present in the solution.

The formula used to calculate boiling point elevation is $\Delta T_b = imK_b$, where ΔT_b is the increase in the boiling point, **m** is the molality of the solute, K_b is the boiling point elevation constant, and **i** is the van't Hoff factor.

The boiling point elevation constant, K_b, is an experimentally determined constant for the solvent. Each solvent will have its own K_b and these values are determined in the laboratory and listed in reference tables. For example, the boiling point elevation constant for water is 0.512^oC/m. As the molality of the solution increases, the boiling point of the solution increases by 0.512^oC for each increase of 1.00 in the molality.

The van't Hoff factor is the ratio between the actual concentration of **particles** produced when the substance is dissolved, and the concentration of the molecules dissolved. For most non-electrolytes dissolved in water, the van't Hoff factor is essentially 1. For most ionic compounds dissolved in water, the van't Hoff factor is equal to the number of discrete ions in a formula unit of the substance. For example, a glucose solution that is 1.00 molal will have a particle concentration that is also 1.00 molal because glucose molecules do not dissociate. A 1.00 molal sodium chloride solution, on the other hand, since it dissociates into two ions will have a ***particle molality*** of 2.00 m. The van't Hoff factor, i, is the number of ions that the molecule will dissociate into when dissolved. Sometimes, in concentrated solutions, an ionic substance does not dissociate 100% and therefore, the value of ***i*** will not be exactly equal to the apparent number of ions produced. In such cases, the value of ***i*** must also be determined experimentally. If you are not given an actual value for ***i*** in the problem, assume that ***i*** is the number of ions apparently produced per molecule. This is true in most dilute solutions.

The formula used to calculate melting point depression is $\Delta T_f = imK_f$, where ΔT_f is the decrease in the melting point, **m** is the molality of the solute, K_f is the melting point depression constant, and **i** is the van't Hoff factor.

The melting point depression constant, K_f, is an experimentally determined constant for the solvent. Each solvent will have its own K_f and these values are determined in the laboratory and listed in reference tables. For example, the freezing point depression constant for water is 1.86^oC/m. As the molality of the solution increases, the melting point of the solution decreases by 1.86^oC for each increase of 1.00 in the molality.

Example: What is the boiling point of a 5.00 m glucose solution in water? Glucose is a non-volatile, non-electrolyte solute. K_b for water = 0.512^oC/m.

Solution: $\Delta T_b = imK_b = (1)(5.00m)(0.512^oC/m) = 2.56^o$C

Since the boiling point of the pure solvent was 100.00^oC, the b.p. of the solution is $100.00^oC + 2.56^oC = 102.56^oC$

Example: What is the melting point of a 5.00 m <mathNaCl</math> solution in water? Sodium chloride is a non-volatile solute that dissociates 100% in water. K_f for water = 1.86^{o}C/m.

Solution: $\Delta T_f = imK_f = (2)(5.00\ m)(1.86^oC/m) = 18.6^o$C (Since $NaCl$ produces two ions in solution, i = 2.)

Since the melting point of the pure solvent was 0.00^{o}C, the m.p. of the solution is $0.00^oC - 18.6^oC = -18.6^oC$

Exercises

1. What is the melting point of a solution produced by dissolving 45.0 g of $NaCl$ in 500. g of water. K_f for water = 1.86^{o}C/m.

2. What is the boiling point of a solution produced by dissolving 45.0 g of NaCl in 500. g of water. K_b for water = 0.512^{o}C/m.

3. Which solution will have higher boiling point: a solution containing 105 g of $C_12H_{22}O_{11}$ in 500. g of water or a solution containing 35.0 g of $NaCl$ in 500. g of water?

4. When 25.0 g of an unknown, non-volatile, non-electrolyte is dissolved in 130. g of water, the boiling point of the solution is 102.5^{o}C. What is the molar mass of the unknown?

5. How many grams of $C_2H_6O_2$ (anti-freeze, a non-electrolyte) must be added to 4,000. grams of water to reduce the melting point to $-40.^o$C?

6. The melting point constant for benzene is 4.90^{o}C/m. The normal melting point of benzene is 5.50^{o}C. What is the melting point of a solution of 9.30 g of $C_{12}H_{25}OH$ (a non-electrolyte) in 250. g of benzene?

7. Assuming 100% dissociation, what is the boiling point of a solution of 200. g of AlF_3 in 500. g of water?

21.8 Lesson 21.8 Colloids

There are no worksheets for this lesson.

21.9 Lesson 21.9 Separating Mixtures

There are no worksheets for this lesson.

Chapter 22

Ions in Solution Worksheets

22.1 Lesson 22.1 Ions in Solution

There are no worksheets for this lesson.

22.2 Lesson 22.2 Covalent Compounds in Solution

There are no worksheets for this lesson.

22.3 Lesson 22.3 Reactions Between Ions in Solutions

Reactions Between Ions in Solution Worksheet

CK-12 Foundation Chemistry

Name___________________________________ **Date**_______________

For the following five reactions (all reactants are in water solution):

- Write and balance the molecular equation indicating the state of each reactant and product.
- Write the total ionic equation.
- Identify the precipitate.
- Identify the spectator ions.
- Write the net ionic equation.

1. iron (III) chloride + sodium hydroxide

Balanced molecular equation ______________________________

Total ionic equation ______________________________

Precipitate = ______________ Spectator ions = ______________

Net ionic equation ______________________________

2. barium chloride + silver nitrate

Balanced molecular equation ______________________________

Total ionic equation ______________________________

Precipitate = ______________ Spectator ions = ______________

Net ionic equation ______________________________

3. magnesium sulfate + potassium phosphate

Balanced molecular equation ______________________________

Total ionic equation ______________________________

Precipitate = ______________ Spectator ions = ______________

Net ionic equation ______________________________

4. copper (II) nitrate + calcium hydroxide

Balanced molecular equation ______________________________

Total ionic equation ______________________________

Precipitate = ______________ Spectator ions = ______________

Net ionic equation ______________________________

5. sodium chromate + strontium nitrate

Balanced molecular equation ______________________________

Total ionic equation ______________________________

Precipitate = ______________ Spectator ions = ______________

Chapter 23

Chemical Kinetics Worksheets

23.1 Lesson 23.1 Rate of Reactions

There are no worksheets for this lesson.

23.2 Lesson 23.2 Collision Theory

There are no worksheets for this lesson.

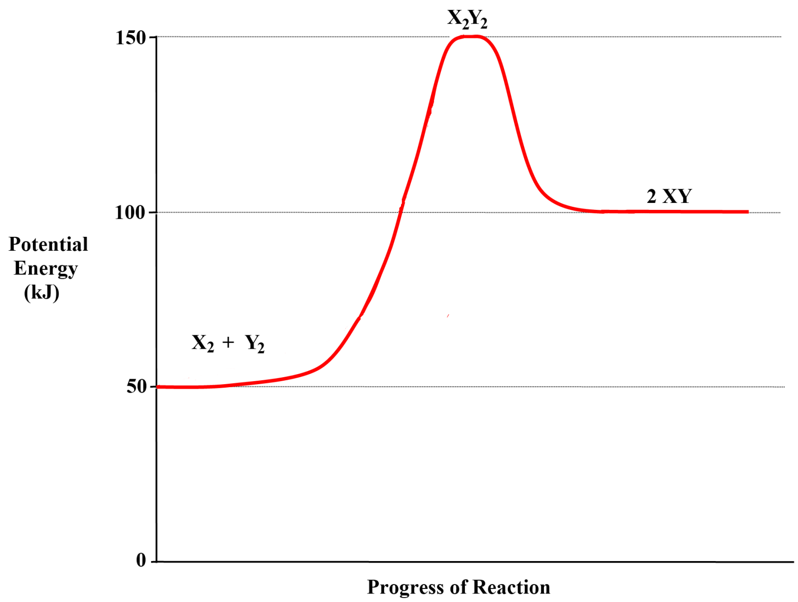
23.3 Lesson 23.3 Potential Energy Diagrams

Potential Energy Diagrams Worksheet

CK-12 Foundation Chemistry

Name___________________________________ **Date**_______________

Use the following Potential Energy Diagram to answer questions 1 - 12.



1. Is the overall reaction as shown exothermic or endothermic? ____________________

2. What is the activation energy for the forward reaction? ____________________

3. What is the activation energy for the reverse reaction? ____________________

4. What is the enthalpy change for (ΔH) for the forward reaction? ____________________

5. What is the ΔH for the reverse reaction? ____________________

6. Is the reverse reaction exothermic or endothermic? ____________________

7. Which species is the activated complex? ____________________

8. Which species or group of species has the *highest* potential energy? ____________________

9. Which species or group of species has the *weakest* bonds? ____________________

10. Which species or group of species has the *strongest* bonds? ____________________

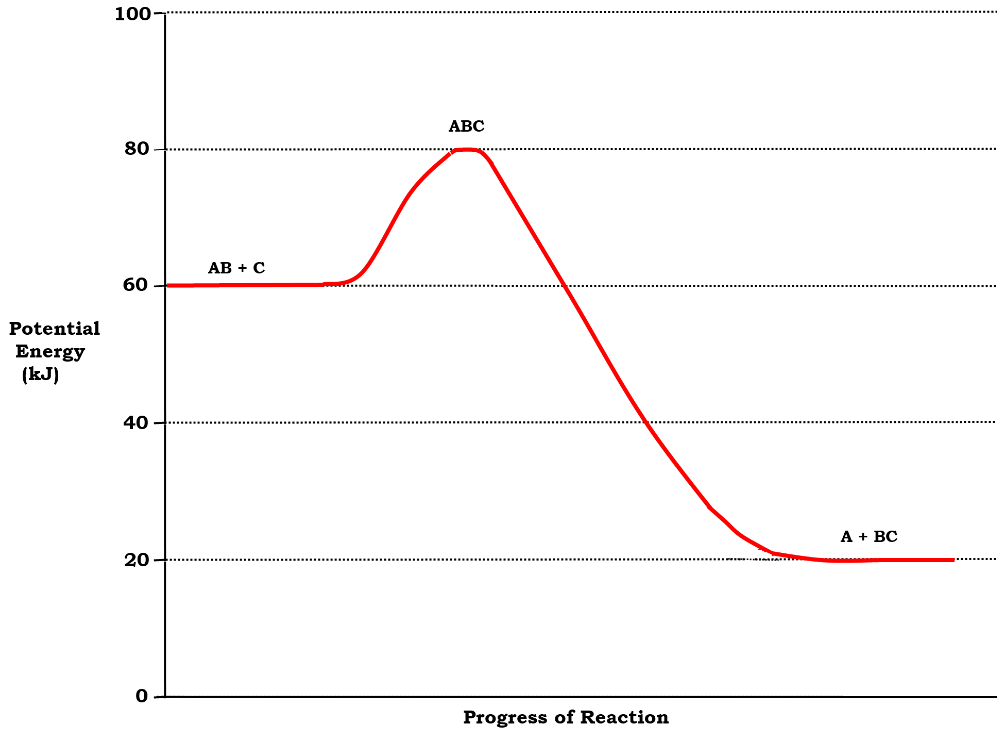
11. Which do you think would be *faster* at that the same temperature, the forward or reverse reaction? ____________________

12. What is the threshold energy for the forward reaction? ____________________

13. In general, as reactant particles begin a collision, the potential energy ____________________ (increases, decreases, stays the same) and the kinetic energy ____________________ (increases, decreases, stays the same).

14. Describe what happens to two reactant particles that collide with less than the activation energy?

Use the following Potential Energy Diagram to answer questions 15 - 22.



15. What is the activation energy for the forward reaction? ____________________
16. What is the activation energy for the reverse reaction? ____________________
17. What is the ΔH for the forward reaction? ______________________
18. What is the ΔH for the reverse reaction? ______________________
19. Is the forward reaction exothermic or endothermic? ______________________
20. What is the threshold energy for the forward reaction? ______________________
21. Which bond is stronger, A-B or B-C? ______________________
22. Give a reason for your answer in question 21.

23.4 Lesson 23.4 Factors That Affect Reaction Rates

There are no worksheets for this lesson.

23.5 Lesson 23.5 Reaction Mechanism

There are no worksheets for this lesson.

Chapter 24

Chemical Equilibrium Worksheets

24.1 Lesson 24.1 Introduction to Equilibrium

There are no worksheets for this lesson.

24.2 Lesson 24.2 Equilibrium Constant

Equilibrium Worksheet

CK-12 Foundation Chemistry

Name____________________________________ **Date**_______________

Questions 1 - 20 relate to the following reaction at equilibrium in a closed container.

$$\mathbf{P_{(s)} + 2\ O_{2(g)} \leftrightarrows PO_{4(g)} \qquad H = -794\ kJ/mol}$$

1. What is the instantaneous effect on the FORWARD REACTION RATE of adding some solid phosphorus with no change in surface area?

A. Increase.
B. Decrease.
C. No change.

2. What is the instantaneous effect on the FORWARD REACTION RATE of adding some oxygen gas with no change in pressure?

A. Increase.
B. Decrease.
C. No change.

3. What is the instantaneous effect on the FORWARD REACTION RATE of adding some PO_4 gas with no change in pressure?

A. Increase.
B. Decrease.
C. No change.

4. What is the instantaneous effect on the FORWARD REACTION RATE of increasing the temperature?

A. Increase.
B. Decrease.
C. No change.

5. What is the instantaneous effect on the FORWARD REACTION RATE of increasing the pressure by reducing the volume?

A. Increase.
B. Decrease.
C. No change.

6. What is the instantaneous effect on the FORWARD REACTION RATE of adding a catalyst?

A. Increase.
B. Decrease.
C. No change.

7. What is the instantaneous effect on the REVERSE REACTION RATE of adding some solid phosphorus with no change in surface area?

A. Increase.
B. Decrease.
C. No change.

8. What is the instantaneous effect on the REVERSE REACTION RATE of adding some oxygen gas with no change in pressure?

A. Increase.
B. Decrease.
C. No change.

9. What is the instantaneous effect on the REVERSE REACTION RATE of adding some PO_4 gas with no change in pressure?

A. Increase.
B. Decrease.
C. No change.

10. What is the instantaneous effect on the REVERSE REACTION RATE of increasing the temperature?

A. Increase.
B. Decrease.
C. No change.

11. What is the instantaneous effect on the REVERSE REACTION RATE of increasing the pressure by reducing the volume?

A. Increase.
B. Decrease.
C. No change.

12. What is the instantaneous effect on the REVERSE REACTION RATE of adding a catalyst?

A. Increase.
B. Decrease.
C. No change.

13. Which direction will the equilibrium shift when solid phosphorus is added with no change in surface area?

A. Forward.

B. Reverse.
C. No shift.

14. Which direction will the equilibrium shift when oxygen gas is added with no change in pressure?

A. Forward.
B. Reverse.
C. No shift.

15. Which direction will the equilibrium shift when gaseous PO_4 is added with no change in pressure?

A. Forward.
B. Reverse.
C. No shift.

16. Which direction will the equilibrium shift when the temperature is increased?

A. Forward.
B. Reverse.
C. No shift.

17. Which direction will the equilibrium shift when the pressure is increased by reducing the volume?

A. Forward.
B. Reverse.
C. No shift.

18. Which direction will the equilibrium shift when a catalyst is added?

A. Forward.
B. Reverse.
C. No shift.

19. Which of the following changes to the sytem at equilibrium will change the value of the equilibrium constant?

I. Adding some solid phosphorus.

II. Adding some oxygen gas.
III. Increasing the pressure by reducing the volume.
IV. Increasing the temperature.
V. Adding a catalyst.

A. I, II, and IV.
B. III, IV, and V.
C. IV and V.
D. IV only.
E. V only.

20. If oxygen gas is added to the system at equilibrium, the equilibrium will shift forward until a new equilibrium is established. When the new equilibrium is established, how will the concentration of oxygen gas in the new equilibrium compare to the original concentration of oxygen gas before the stress was applied?

A. higher
B. lower
C. the same

21. Here are four equations with their equilibrium constant values. Which of these reactions will have the greatest proportion of material in the form of products?

Table 24.1:

Choice	**Equation**	**Equilibrium Constant**
A.	$AB_{(aq)} \leftrightarrows A^+_{(aq)} + B^-_{(aq)}$	$K_e = 2\ x\ 10^{-2}$
B.	$CD_{(aq)} \leftrightarrows C^+_{(aq)} + D^-_{(aq)}$	$K_e = 3\ x\ 10^{-2}$
C.	$EF_{(aq)} \leftrightarrows E^+_{(aq)} + F^-_{(aq)}$	$K_e = 3\ x\ 10^{-3}$
D.	$GH_{(aq)} \leftrightarrows G^+_{(aq)} + H^-_{(aq)}$	$K_e = 6\ x\ 10^{-3}$

22. Solid sulfur reacts with oxygen gas to form $SO_{2(g)}$ according to the following equation.

$$S_{(s)} + O_{2(g)} \leftrightarrows SO_{2(g)}$$

Given that the equilibrium constant for the reaction is 5.00 and that the reaction begins with 60.0 M sulfur and 3.00 M O_2, calculate the equilibrium concentration of SO_2.

A. 15.0 M
B. 5.55 M
C. 2.50 M
D. 1.25 M
E. None of these.

23. For the reaction, $N_{2(g)} + O_{2(g)} \leftrightarrows 2\ NO_{2(g)}$, the equilibrium constant is 1.0 x 10^{-6}. Find the equilibrium concentration of NO_2 if the beginning concentration of N_2 and O_2 are both 2.0 M?

A. 0.0020 M
B. 2.0 x 10^{-6} M
C. 4.0 x 10^{-6} M
D. 0.020 M
E. None of these.

24. For the reaction, $H_{2(g)} + CO_{2(g)} \leftrightarrows H_2O_{(g)} + CO_{(g)}$, the two reactants begin the reaction at 1.0 M and at equilibrium, the concentration of CO is found to be 0.80 M. What is the equilibrium constant value?

A. 1.7
B. 2.0
C. 4.0
D. 16
E. None of these.

25. $K_e = 4.00$ for the reaction, $H_{2(g)} + CO_{2(g)} \leftrightarrows H_2O_{(g)} + CO_{(g)}$. If all four species begin at 1.00 M, what will be the equilibrium concentration of H_2?

A. 0.33 M
B. 0.67 M
C. 1.3 M
D. 1.0 M
E. None of these.

24.3 Lesson 24.3 The Effect of Applying Stress to Reactions at Equilibrium

Le Chatelier's Principle Worksheet

CK-12 Foundation Chemistry

Name___________________________________ **Date**______________

Le Chatelier's Principle is useful in predicting how a system at equilibrium will respond when certain changes are imposed. Le Chatelier's Principle does NOT explain why the system changes, and is not an acceptable explanation for the change. It merely allows you to determine quickly how the system will change when a disturbance is imposed. The explanation for why the system changes can be found in your textbook.

There are three common ways a stress may be applied to a chemical system at equilibrium:

- changing the concentration (or partial pressure) of a reactant or product.
- changing the temperature.
- changing the volume of the container (which changes partial pressure of all gases in the reaction).

You should be aware that adding a gaseous substance that is not involved in the reaction changes the total pressure in the system but does not change the partial pressure of any of the reactants or products and therefore does not affect the equilibrium.

Le Chatelier's Principle states *when a system at equilibrium is disturbed, the equilibrium shifts so as to* ***partially*** *undo (counteract) the effect of the disturbance.*

Changes in Concentration or Partial Pressure

If a system at equilibrium is disturbed by adding a reactant or removing a product, Le Chatelier's Principle predicts that the equilibrium will shift forward, thus using up some of the added reactant or producing more of the removed product. In this way, the equilibrium shift partially counteracts the disturbance. Similarly, if the disturbance is the removal of a reactant or the addition of a product, the equilibrium will shift backward, thus producing more of the removed reactant or using up some of the added product. Once again, the shift tends to "undo" the disturbance. It should be noted that when the disturbance is an increase or decrease of concentration of reactant or product, the equilibrium shift tends to **partially** return the concentration to its former value but **it never gets all the way back to the former value**.

The equilibrium constant value, K_e is not changed by the addition or removal of reactants or products. Since the concentration of solids are constant, they do not appear in the equilibrium constant expression and their concentrations do not change when disturbances cause equilibrium shifts, however, the **amount** of the solid present most certainly does change. The amount of solid can increase or decrease but the *concentration* does not change.

Changes in Temperature

Increasing the temperature of a system at equilibrium increases both forward and reverse reaction rate, but it increases the endothermic reaction more that the exothermic. Therefore, in an exothermic reaction, the reverse reaction is endothermic and so increasing the temperature will increase the reverse reaction more than the forward reaction, and the equilibrium will shift backwards. Since the forward reaction produces heat and the reverse reaction consumes heat, Le Chatelier's Principle predicts that when heat is added, the equilibrium will shift backward, consuming heat, and thus partially countering the disturbance. Cooling an exothermic reaction slows both reactions but it slows the reverse more than the forward, hence the equilibrium will shift forward producing more heat, thus partially undoing the stress.

For an endothermic reaction, all the same logic is involved except that the forward reaction is endothermic and the reverse reaction is exothermic. Therefore, heating an endothermic reaction causes the equilibrium to shift forward, and cooling an endothermic reaction causes the equilibrium to shift backward.

When an equilibrium shifts due to a temperature change all the substances on one side of the equation move in the same direction, that is, they all increase or they all decrease. Therefore, the equilibrium constant value will also change when the temperature is changed.

Table 24.2: **Summary of K changes due to Temperature Changes**

Reaction Type	**Increase Temperature**	**Decrease Temperature**
Endothermic ($\Delta H > 0$)	K increases	K decreases
Exothermic ($\Delta H < 0$)	K decreases	K increases

Changes in Volume

When the volume of a reaction vessel is decreased, the partial pressure (and concentration) of all gases in the container increase. The total pressure in the vessel will also increase. Le Chatelier's Principle predicts that the equilibrium will shift in a direction that tends to counteract the disturbance. Therefore, the equilibrium will shift to produce fewer moles of gaseous substances so that the pressure will decrease. Thus, decreasing the volume will cause the equilibrium to shift toward the side with fewer moles of gaseous substances. The reverse is true if the volume of the vessel is increased. The partial pressure of all gases will decrease, and the total pressure will decrease, so the equilibrium shift will be toward the side that contains more moles of gas, thus increasing pressure and partially counteracting the change.

The Addition of a Catalyst

The addition of a catalyst will increase both forward and reverse reaction rates. In the case of a catalyst, both reaction rates are increased by the same amount and therefore there will be no equilibrium shift.

Exercises

Consider the following reaction.

$$5\ CO_{(g)} + I_2O_{5(s)} \rightleftharpoons I_{2(g)} + 5\ CO_{2(g)} \qquad \Delta H^o = -1175\ kJ$$

1. If some $CO_{2(g)}$ is added to this sytem at equilibrium, which way will the equilibrium shift?

A. Toward the products.
B. Toward the reactants.
C. No shift.

2. When equilibrium is re-established after the $CO_{2(g)}$ is added, how will the concentration of $I)2(g)$ compare to the original concentration?

A. Increased.
B. Decreased.
C. No change.

3. When equilibrium is re-established after the $CO_{2(g)}$ is added, how will the concentration of I_2O_5 compare to the original concentration?

A. Increased.
B. Decreased.
C. No change.

4. When equilibrium is re-established after the $CO_{2(g)}$ is added, how will the amount of I_2O_5 compare to the original amount?

A. Increased.
B. Decreased.
C. No change.

5. When equilibrium is re-established after the $CO_{2(g)}$ is added, how will the value of K compare to the original value of K?

A. Higher.
B. Lower.
C. No change.

6. If some $I_{2(g)}$ is removed from this sytem at equilibrium, which way will the equilibrium shift?

A. Toward the products.
B. Toward the reactants.

C. No shift.

7. When equilibrium is re-established after the $I_2(g)$ is removed, how will the concentration of $CO_2(g)$ compare to the original concentration?

A. Increased.
B. Decreased.
C. No change.

8. When equilibrium is re-established after the $I_{2(g)}$ is removed, how will the concentration of $I_{2(g)}$ compare to the original concentration?

A. Increased.
B. Decreased.
C. No change.

9. 5. When equilibrium is re-established after the $I_{2(g)}$ is removed, how will the value of K compare to the original value of K?

A. Higher.
B. Lower.
C. No change.

10. If the temperature of this system at equilibrium is lowered, which way will the equilibrium shift?

A. Toward the products.
B. Toward the reactants.
C. No shift.

11. When equilibrium is re-established after the temperature was lowered, how will the concentration of $CO_{(g)}$ compare to its original concentration?

A. Increased.
B. Decreased.
C. No change.

12. When equilibrium is re-established after the temperature was lowered, how will the value of K compare to the original value of K?

A. Higher.
B. Lower.
C. No change.

13. If the volume of the reaction vessel for this system at equilibrium is decreased, which way will the equilibrium shift?

A. Toward the products.
B. Toward the reactants.
C. No shift.

14. When equilibrium is re-established after the volume was decreased, how will the concentration of $CO_{(g)}$ compare to its original concentration?

A. Higher.
B. Lower.
C. No change.

15. When equilibrium is re-established after the volume was decreased, how will the value of K compare to the original value of K?

A. Higher.
B. Lower.
C. No change.

Consider the following reaction.

$$4\ NO_{(g)}\ +\ 6\ H_2O_{(g)}\ \rightleftharpoons\ 4\ NH_{3(g)}\ +\ 5\ O_{2(g)} \qquad \Delta H\ =\ +1532\ kJ$$

16. If some $NO_{(g)}$ is added to this sytem at equilibrium, which way will the equilibrium shift?

A. Toward the products.
B. Toward the reactants.
C. No shift.

17. When equilibrium is re-established after the $NO_{(g)}$ is added, how will the concentration of $NH_{3(g)}$ compare to the original concentration?

A. Increased.

B. Decreased.
C. No change.

18. If the temperature of this system at equilibrium is raised, which way will the equilibrium shift?

A. Toward the products.
B. Toward the reactants.
C. No shift.

19. When equilibrium is re-established after the temperature was raised, how will the concentration of $NO_{(g)}$ compare to its original concentration?

A. Increased.
B. Decreased.
C. No change.

20. When equilibrium is re-established after the temperature was raised, how will the value of K compare to the original value of K?

A. Higher.
B. Lower.
C. No change.

24.4 Lesson 24.4 Slightly Soluble Salts

Solubility and Solubility Product Constant Worksheet

CK-12 Foundation Chemistry

Name____________________________________ **Date**_______________

1. When excess solid $SrCrO_4$ is shaken with water at 25^oC, it is found that 6.00 x 10^{-3} moles dissolve per liter of solution. Use this information to calculate the K_{sp} for $SrCrO_4$.

2. The solubility of $PbCl_2$ is 1.6 x 10^{-2} mol/L. What is the K_{sp} for $PbCl_2$?

3. The solubility of $AgC_2H_3O_2$ is 11.11 g/L at 25^oC. What is the K_{sp} for silver acetate at this temperature?

4. The solubility of $Ag_2Cr_2O_7$ is 0.083 g/L at 25^oC. What is the K_{sp} for silver dichromate at this temperature?

5. What is the solubility of AgI in grams/liter given the $K_{sp} = 8.3\ x\ 10^{-17}$?

6. What is the solubility of $Ca(OH)_2$ in grams/liter given the $K_{sp} = 6.0\ x\ 10^{-6}$?

7. Write balanced net ionic equations for the precipitation reactions that occur when the following pairs of solutions are mixed. If no reaction occurs, write ”no reaction”. Use the solubility table in your textbook if you need it.

A. Lead nitrate and hydrochloric acid.
B. Silver nitrate and lithium hydroxide.
C. Ammonium sulfide and cobalt (II) bromide.
D. Copper (II) sulfate and potassium carbonate.
E. Barium nitrate and copper (II) sulfate.

8. Lead (II) chloride has a K_{sp} value of 1.7 $x\ 10^{-5}$. Will a precipitate form when 140.0 mL of 0.0100 M $Pb_3(PO_4)_2$ is mixed with 550.0 mL of 0.0550 M NaCl?

9. A solution contains 1.0 $x\ 10^{-4}$ M Pb^{2+} ions and 2.0 $x\ 10^{-3}$ M Sr^{2+} ions. If a source of SO_4^{2-} ions is very slowly added to this solution, will $PbSO_4$, ($K_{sp} = 1.8\ x\ 10^{-8}$) or $SrSO_4$, ($K_{sp} = 3.4\ x\ 10^{-7}$) precipitate first? Calculate the concentration of SO_4^{2-} ions that will begin to precipitate each cation.

Chapter 25

Acids and Bases Worksheets

25.1 Lesson 25.1 Arrhenius Acids

There are no worksheets for this lesson.

25.2 Lesson 25.2 Strong and Weak Acids

Strong Acids and Bases Worksheet

CK-12 Foundation Chemistry

Name____________________________________ **Date**_______________

1. If the hydrogen ion concentration in a solution is 1.00 x 10^{-4} M, what is the hydroxide ion concentration?

2. What is the hydroxide ion concentration in a solution whose pH is 11?

3. What is the hydrogen ion concentration in a solution prepared by dissolving 0.400 grams of $NaOH$ in enough water to make 2.00 liters of solution?

4. How many mL of 0.100 M potassium hydroxide are required to neutralize 75.0 mL of 0.500 M HNO_3?

5. If 50.0 mL of H_2SO_4 are neutralized by 100. mL of 0.200 M LiOH, what is the molarity of the H_2SO_4?

6. What volume of 6.00 M HCl would be necessary to neutralize 400. mL of 3.00 M $Ba(OH)_2$?

7. 200. mL of 0.0150 M $NaOH$ is mixed with 300. mL of 0.00100 M HCl. What is the final

$[H^+]$ and $[OH^-]$?

8. What is the pH of the final solution in problem 7?

9. 700. mL of 1.00 x 10^{-4} M H_2SO_4 is mixed with 300. mL of 1.00 x 10^{-3} M $Ba(OH)_2$. What is the final $[H^+]$ and $[OH^-]$?

10. What is the pH of the final solution in problem 9?

11. 25.0 mL of 0.0100 M HCl is mixed with 35.0 mL of 0.0300 M $NaOH$. What is the final $[H^+]$ and $[OH^-]$?

12. What is the pH of the final solution in problem 11?

13. What is the final $[H^+]$ and $[OH^-]$ in a solution made by adding 100. mL of 0.000200 M HNO_3 to 100. mL of 0.0000990 M $Ba(OH)_2$?

14. What is the pH of the final solution in problem 13?

15. What is the molar mass of a solid monoprotic acid if 0.300 grams of the acid requires 30.0 mL of 0.200 M $NaOH$ to neutralize it?

25.3 Lesson 25.3 Arrhenius Bases

There are no worksheets for this lesson.

25.4 Lesson 24.4 Salts

There are no worksheets for this lesson.

25.5 Lesson 25.5 pH

pH Worksheet

CK-12 Foundation Chemistry

Name______________________________________ **Date**_______________

1. Calculate the pH of a solution with $[H^+]$ = 7.0 x 10^{-5} M.

2. Calculate the pH of a solution that is 0.050 M $NaOH$.

3. Calculate the pH of a solution that is 7.0 x 10^{-5} M $Mg(OH)_2$.

4. What is the $[H^+]$ in a solution with pH = 4.4?

5. What is the $[OH^-]$ in a solution with pH = 3.0?

6. 10.0 g of KOH is added to enough water to make 400. mL of solution. What is the pH?

7. A 1.0 liter solution has a pH = 2. How many liters of water must be added to change the pH to 3?

8. If you do the regular calculations to determine the pH of a 1.0 x 10^{-12} M HBr solution, you will get the pH = 12. You should have a feeling that something is wrong with this situation because this indicates that a solution of acid has a basic pH. What do you think is wrong with this calculation?

Complete the following table.

Table 25.1: **Acid, Base, or Neutral**

	pH	$[H^+]$	$[OH^-]$	**A, B, or N**
9.		6.2 x 10^{-4} M		
10.			8.5 x 10^{-10} M	
11.	10.75			
12.			4.0 x 10^{-2} M	

25.6 Lesson 25.6 Weak Acid/Base Equilibria

Weak Acids and Bases Worksheet

CK-12 Foundation Chemistry

Name____________________________________ **Date**_______________

1. Explain the difference between the designations "strong" acid and "weak" acid.

2. The K_a of acid A is 6.4 x 10^{-4} and the K_a of acid B is 1.7 x 10^{-5}. Which acid is the stronger acid?

3. Explain what happens to the pH of a solution of acetic acid when a solution of sodium acetate is added to it.

4. Explain why a solution of sodium acetate will be basic.

5. What is the pH of a 0.0100 M solution of a weak acid, HX, if the K_a for HX is 8.1 x 10^{-7}.

6. The pH of a 0.100 M solution of a weak acid, HQ, is 4.0. What is the K_a of this acid?

7. What is the pH of a 0.150 M solution of NH_4OH? The K_b for NH_4OH is 1.80 x 10^{-5}.

8. The pH of a 1.00 M solution of the weak base methylamine is 12.3. The equation for the

reaction of methylamine in water is

$$CH_3NH_{2(aq)} + H_2O \leftrightharpoons CH_3NH^+_{3(aq)} + OH^-_{(aq)}.$$

What is the K_b for methylamine?

9. Will a 1.00 M solution of potassium acetate be acidic, basic, or neutral?

10. Will a 1.00 M solution of NH_4NO_2 be acidic, basic, or neutral? Use 1.8 x 10^{-5} as the K_b for NH_4OH and 7.1 x 10^{-4} as the K_a for HNO_2.

25.7 Lesson 25.7 Bronsted Lowry Acids-Bases

There are no worksheets for this lesson.

25.8 Lesson 25.8 Lewis Acids and Bases

There are no worksheets for this lesson.

Chapter 26

Water, pH and Titration Worksheets

26.1 Lesson 26.1 Water Ionizes

There are no worksheets for this lesson.

26.2 Lesson 26.2 Indicators

There are no worksheets for this lesson.

26.3 Lesson 26.3 Titrations

There are no worksheets for this lesson.

26.4 Lesson 26.4 Buffers

There are no worksheets for this lesson.

Chapter 27

Thermodynamics Worksheets - HS Chemistry

27.1 Lesson 27.1 Energy Change in Reactions

There are no worksheets for this lesson.

27.2 Lesson 27.2 Enthalpy

Enthalpy Worksheet

1. The combustion of methane, CH_4, releases 890.4 kJ/mol of heat. That is, when one mole of methane is burned, 890.4 kJ are given off to the surroundings. This means that the products have 890.4 kJ less energy stored in the bonds than the reactants. Thus, ΔH for the reaction = - 890.4 kJ. A negative symbol for ΔH indicates an exothermic reaction.

$$CH_{4(g)} \ + \ 2\ O_{2(g)} \ \rightarrow \ CO_{2(g)} \ + \ 2\ H_2O_{(L)} \qquad \Delta H \ = \ -890.4\ kJ$$

A. How much energy is given off when 2.00 mol of CH_4 are burned?

B. How much energy is released when 22.4 g of CH_4 are burned?

C. If you were to attempt to make 45.0 g of methane from CO_2 and H_2O (with O_2 also being produced), how much heat would be absorbed during the reaction?

Use the following heat of formation table in questions 2 – 6.

Table 27.1: **The Standard Enthalpy and Entropy of Various Substances**

Substance	ΔH_f^o **(kJ/mol)**	S^o $(J/K \cdot mol)$
$C_4H_{10(g)}$	-126	310
$CaC_{2(s)}$	-63	70.
$Ca(OH)_{2(s)}$	-987	83
$C_2H_{2(g)}$	227	201
$CO_{2(g)}$	-394	214
$H_{2(g)}$	0	131
$H_2O_{(g)}$	-242	189
$H_2O_{(L)}$	-286	70.
$NH_{3(g)}$	-46	193
$NO_{(g)}$	90.	211
$NO_{2(g)}$	34	240.
$N_2O_{(g)}$	82	220.
$O_{2(g)}$	0	205
$O_{3(g)}$	143	239

2. Using data from the heat of formation table above, calculate the enthalpy of reaction for

$$3\ H_{2(g)}\ +\ O_{3(g)}\ \rightarrow\ 3\ H_2O_{(g)}\ .$$

3. Using data from the heat of formation table above, calculate the heat of reaction for

$$2\ NO_{(g)}\ +\ O_{2(g)}\ \rightarrow\ 2\ NO_{2(g)}\ .$$

4. Using data from the heat of formation table above, calculate the heat of reaction for

$$N_2O_{(g)}\ +\ NO_{2(g)}\ \rightarrow\ 3\ NO_{(g)}\ .$$

5. Using data from the heat of formation table above, calculate the heat of reaction for

$$CaC_{2(s)}\ +\ 2\ H_2O_{(L)}\ \rightarrow\ Ca(OH)_{2(s)}\ +\ C_2H_{2(g)}\ .$$

6. Many cigarette lighters contain liquid butane, C_4H_{10}. Using the heat of formation table above, calculate the quantity of heat produced when 1.0 g of gaseous butane is completely combusted in air.

Hess's Law Worksheet

CK-12 Foundation Chemistry

Name________________________________ **Date**______________

Example Problem

Find the ΔH for the reaction below, using the following reactions and their ΔH values.

$$N_2H_{4(L)} \ + \ H_{2(g)} \ \rightarrow \ 2\ NH_{3(g)}$$

Table 27.2: **Given Equations and**

Equation	ΔH **Value**
$N_2H_{4(L)} + CH_4O_{(L)} \ \rightarrow \ CH_2O_{(g)} + N_{2(g)} + 3\ H_{2(g)}$	$\Delta H \ = \ -37\ kJ$
$N_{2(g)} \ + \ 3\ H_{2(g)} \ \rightarrow \ 2\ NH_{3(g)}$	$\Delta H \ = \ -46\ kJ$
$CH_4O_{(L)} \ \rightarrow \ CH_2O_{(g)} \ + \ H_{2(g)}$	$\Delta H \ = \ -65\ kJ$

Solution

Table 27.3: **Solution Arrangement**

Changes	**Equation**	ΔH **Value**
Keep Same	$N_2H_{4(L)} + CH_4O_{(L)} \ \rightarrow \ CH_2O_{(g)} + N_{2(g)} + 3\ H_{2(g)}$	$\Delta H \ = \ -37\ kJ$
Keep Same	$N_{2(g)} + 3\ H_{2(g)} \ \rightarrow \ 2\ NH_{3(g)}$	$\Delta H \ = \ -46\ kJ$
Reverse	$CH_2O_{(g)} + H_{2(g)} \ \rightarrow \ CH_4O_{(L)}$	$\Delta H \ = \ +65\ kJ$

Sum $\quad N_2H_{4(L)} \ + \ H_{2(g)} \ \rightarrow \ 2\ NH_{3(g)} \qquad \Delta H \ = \ -18\ kJ$

Exercises

1. Find the ΔH for the reaction below, using the following reactions and their ΔH values.

$$H_2SO_{4(L)} \ \rightarrow \ SO_{3(g)} \ + \ H_2O_{(g)}$$

Table 27.4: **Given Equations and**

Equation	ΔH **Value**
$H_2S_{(g)} \ + \ 2\ O_{2(g)} \ \rightarrow \ H_2SO_{4(L)}$	$\Delta H \ = \ -235\ kJ$

Table 27.4: (continued)

Equation	ΔH Value
$H_2S_{(g)} + 2\ O_{2(g)} \rightarrow SO_{3(g)} + H_2O_{(L)}$	$\Delta H = -207\ kJ$
$H_2O_{(L)} \rightarrow H_2O_{(g)}$	$\Delta H = +44\ kJ$

2. Find the ΔH for the reaction below, using the following reactions and their ΔH values.

$$4\ NH_{3(g)} + 5\ O_{2(g)} \rightarrow 4\ NO_{(g)} + 6\ H_2O_{(g)}$$

Table 27.5: **Given Equations and**

Equation	ΔH Value
$N_{2(g)} + O_{2(g)} \rightarrow 2\ NO_{(g)}$	$\Delta H = -180.5\ kJ$
$N_{2(g)} + 3\ H_{2(g)} \rightarrow 2\ NH_{3(g)}$	$\Delta H = -91.8\ kJ$
$2\ H_{2(g)} + O_{2(g)} \rightarrow 2\ H_2O_{(g)}$	$\Delta H = -483.6\ kJ$

3. Find the ΔH for the reaction below, using the following reactions and their ΔH values.

$$PCl_{5(g)} \rightarrow PCl_{3(g)} + Cl_{2(g)}$$

Table 27.6:

Equation	ΔH Value
$P_{4(s)} + 6\ Cl_{2(g)} \rightarrow 4\ PCl_{3(g)}$	$\Delta H = -2439\ kJ$
$4\ PCl_{5(g)} \rightarrow P_{4(s)} + 10\ Cl_{2(g)}$	$\Delta H = +3438\ kJ$

4. Find the ΔH for the reaction below, using the following reactions and their ΔH values.

$$3\ H_{2(g)} + 2\ C_{(s)} + \frac{1}{2}\ O_{2(g)} \rightarrow C_2H_5OH_{(L)}$$

Table 27.7: **Given Equations and**

Equation	ΔH Value
$C_2H_5OH_{(L)} + 3\ O_{2(g)} \rightarrow 2\ CO_{2(g)} + 3\ H_2O_{(L)}$	$\Delta H = -875.0\ kJ$
$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$	$\Delta H = -394.5\ kJ$
$H_{2(g)} + \frac{1}{2}\ O_{2(g)} \rightarrow H_2O_{(L)}$	$\Delta H = -285.8\ kJ$

5. Find the ΔH for the reaction below, using the following reactions and their ΔH values.

$$2\ CO_{2(g)}\ +\ H_2O_{(g)}\ \rightarrow\ C_2H_{2(g)}\ +\ \frac{5}{2}\ O_{2(g)}$$

Table 27.8: **Given Equations and**

Equation	**ΔH Value**
$C_2H_{2(g)}\ +\ 2\ H_{2(g)}\ \rightarrow\ C_2H_{6(g)}$	$\Delta H\ =\ -94.5\ kJ$
$H_2O_{(g)}\ \rightarrow\ H_{2(g)}\ +\ \frac{1}{2}\ O_{2(g)}$	$\Delta H\ =\ +71.2\ kJ$
$C_2H_{6(g)}\ +\ \frac{7}{2}\ O_{2(g)}\ \rightarrow\ 2\ CO_{2(g)}\ +\ 3\ H_2O_{(g)}$	$\Delta H\ =\ -283.0\ kJ$

6. Find the ΔH for the reaction below, using the following reactions and their ΔH values.

$$\frac{1}{2}\ H_{2(g)}\ +\ \frac{1}{2}\ Cl_{2(g)}\ \rightarrow\ HCl_{(g)}$$

Table 27.9: **Given Equations and**

Equation	**ΔH Value**
$COCl_{2(g)}\ +\ H_2O_{(L)}\ \rightarrow\ CH_2Cl_{2(L)}\ +\ O_{2(g)}$	$\Delta H\ =\ +48\ kJ$
$2\ HCl_{(g)}\ +\ \frac{1}{2}\ O_{2(g)}\ \rightarrow\ H_2O_{(L)}\ +\ Cl_{2(g)}$	$\Delta H\ =\ +105\ kJ$
$CH_2Cl_{2(L)} + H_{2(g)} + \frac{3}{2}\ O_{2(g)}\ \rightarrow\ COCl_{2(g)} + 2\ H_2O_{(L)}$	$\Delta H\ =\ -403\ kJ$

27.3 Lesson 27.3 Spontaneous Processes

There are no worksheets for this lesson.

27.4 Lesson 27.4 Entropy

Entropy Worksheet

Use the following entropy of formation table in questions 1 – 5.

Table 27.10: **The Standard Enthalpy and Entropy of Various Substances**

Substance	**ΔH_f^o (kJ/mol)**	$S^o\ (J/K \cdot mol)$
$C_4H_{10(g)}$	-126	310

Table 27.10: (continued)

Substance	ΔH_f^o **(kJ/mol)**	S^o $(J/K \cdot mol)$
$CaC_{2(s)}$	-63	70.
$Ca(OH)_{2(s)}$	-987	83
$C_2H_{2(g)}$	227	201
$CO_{2(g)}$	-394	214
$H_{2(g)}$	0	131
$H_2O_{(g)}$	-242	189
$H_2O_{(L)}$	-286	70.
$NH_{3(g)}$	-46	193
$NO_{(g)}$	90.	211
$NO_{2(g)}$	34	240.
$N_2O_{(g)}$	82	220.
$O_{2(g)}$	0	205
$O_{3(g)}$	143	239

1. Using data from the entropy of formation table above, calculate the entropy of reaction for

$$3\ H_{2(g)}\ +\ O_{3(g)}\ \rightarrow\ 3\ H_2O_{(g)}.$$

2. Using data from the entropy of formation table above, calculate the change in entropy for

$$2\ NO_{(g)}\ +\ O_{2(g)}\ \rightarrow\ 2\ NO_{2(g)}\ .$$

3. Using data from the heat of formation table above, calculate the ΔS^o for

$$N_2O_{(g)}\ +\ NO_{2(g)}\ \rightarrow\ 3\ NO_{(g)}.$$

4. Using data from the entropy of formation table above, calculate the heat of reaction for

$$CaC_{2(s)}\ +\ 2\ H_2O_{(L)}\ \rightarrow\ Ca(OH)_{2(s)}\ +\ C_2H_{2(g)}.$$

5. Using the entropy of formation table above, calculate the change in entropy for the following reaction.

$$C_4H_{10(g)}\ +\ \tfrac{13}{2}\ O_{2(g)}\ \rightarrow\ 4\ CO_{2(g)}\ +\ 5\ H_2O_{(g)}$$

27.5 Lesson 27.5 Gibb's Free Energy

Enthalpy, Entropy, and Free Energy Worksheet

CK-12 Foundation Chemistry

Name____________________________________ **Date**_______________

1. As the amount of energy required to decompose a compound increases, the thermodynamic stability of the compound ____________________.

A. increases
B. decreases
C. remains constant
D. varies randomly

2. The enthalpy of formation for a free element is

A. 0 kJ/mol.
B. 1 kJ/mol.
C. 10 kJ/mol.
D. -100 kJ/mol.
E. variable.

Questions 3 and 4 relate to the following equation and ΔH_R value.

$$2\ HgO_{(s)} \ \rightarrow \ 2\ Hg_{(L)} \ + \ O_{2(g)} \qquad \Delta H_R \ = \ +181.7\ kJ$$

3. Which of the following can definitely be concluded from the equation and heat of reaction above?

A. The reaction is spontaneous.
B. The reaction is non-spontaneous.
C. The reaction is endothermic.
D. The reaction is exothermic.
E. None of these.

4. From the equation and heat of reaction above, what is the ΔH_f of HgO?

A. 181.7 kJ/mol

B. -181.7 kJ/mol
C. 0 kJ/mol
D. 90.9 kJ/mol
E. -90.9 kJ/mol

5. Which of the following four substances is the most thermodynamically stable? Use the data in the Thermodynamic Data Table at the bottom of the worksheet.

A. $NH_{3(g)}$
B. $CO_{2(g)}$
C. $H_2O_{(L)}$
D. $NO_{(g)}$

6. The free energy of a reaction is the combination of ____________________ and ____________________.

A. heat and work
B. pressure and volume
C. enthalpy and entropy
D. internal energy and PV
E. None of these.

7. All reactions that occur spontaneously must have a negative _________________.

A. $T\Delta S$
B. ΔG
C. ΔH
D. ΔS
E. All of these.

Questions 8, 9, 10, and 11, relate to the equation shown below.

$$4\ NH_{3(g)}\ +\ 5\ CO_{2(g)}\ \rightarrow\ 6H_2O_{(L)}\ +\ 4\ NO_{(g)}$$

8. Use the data in the Thermodynamic Data Table at the bottom of this worksheet to find the ΔH_R for the reaction above?

A. +92.8 kJ
B. -92.8 kJ
C. -806.3 kJ
D. +806.3 kJ

E. None of these.

9. Use the data in the Thermodynamic Data Table at the bottom of this worksheet to find the ΔG_R for the reaction above?

A. -981.6 kJ
B. +981.6 kJ
C. -269.0 kJ
D. +269.0 kJ
E. None of these.

10. Use the data in the Thermodynamic Data Table at the bottom of this worksheet to find the ΔS_R for the reaction above?

A. $-575.9\ J/^o$
B. $+575.9\ J/^o$
C. $-1419.1\ J/^o$
D. $+1419.1\ J/^o$
E. None of these.

11. Use the ΔH_R you found in question 6 and the ΔS_R you found in question 8 to calculate ΔG_R for this reaction.

A. 634.7 kJ
B. -634.7 kJ
C. 977.9 kJ
D. -977.9 kJ
E. None of these.

12. Find ΔS for the reaction, $2\ NO_{(g)}\ +\ O_{2(g)}\ \rightarrow\ 2\ NO_{2(g)}$.

A. -146.5 J/K
B. +146.5 J/K
C. -16.5 J/K
D. +16.5 J/K
E. None of these.

13. Find ΔG_R for the reaction, $2\ H_2O_{(g)}\ +\ 2\ F_{2(g)}\ \rightarrow\ O_{2(g)}\ +\ 4HF_{(g)}$.

A. -1550.0 kJ

B. +1550.0 kJ
C. -635.6 kJ
D. +635.6 kJ
E. None of these.

14. What is the change in enthalpy for 4 $Al_{(s)}\ +\ 3\ O_{2(g)}\ \rightarrow\ 2\ Al_2O_{3(s)}$?

A. 0 kJ
B. -1657.7 kJ
C. +1657.7 kJ
D. +3351.4 kJ
E. -3351.4 kJ

15. What is the change in entropy for 4 $Al_{(s)}\ +\ 3\ O_{2(g)}\ \rightarrow\ 2\ Al_2O_{3(s)}$?

A. 0 J/K
B. -626.7 J/K
C. +626.7 J/K
D. -500.0 J/K
E. +500.0 J/K

16. Use the results from questions 14 and 15 to determine under what conditions this reaction will be spontaneous.

A. This reaction will be spontaneous at all temperatures.
B. This reaction will never be spontaneous at any temperature.
C. This reaction will be spontaneous at high temperatures.
D. This reaction will be spontaneous at low temperatures.

Table 27.11: **Thermodynamic Properties of Some Substances (at**

Substance	ΔH_f^o **(kJ/mol)**	ΔG_f^o **(kJ/mol)**	$S^o\ (J/mol \cdot K)$
$Al_{(s)}$	0	0	+28.3
$Al_2O_{3(s)}$	-1675.7	-1582.3	+50.9
$CO_{(g)}$	-110.5	-137.2	+197.7
$CO_{2(g)}$	-393.5	-394.4	+213.7
$F_{2(g)}$	0	0	+202.8
$HF_{(g)}$	-271.1	-273.2	+173.8
$H_2O_{(L)}$	-285.8	-237.1	+69.9
$H_2O_{(g)}$	-241.8	-228.6	+188.8

Table 27.11: (continued)

Substance	ΔH_f^o **(kJ/mol)**	ΔG_f^o **(kJ/mol)**	S^o $(J/mol \cdot K)$
$NH_{3(g)}$	-46.1	-16.5	+192.5
$NO_{(g)}$	+90.3	+86.6	+210.8
$NO_{2(g)}$	+33.2	+51.3	+240.1
$O_{2(g)}$	0	0	+205.1

Chapter 28

Electrochemistry Worksheets

28.1 Lesson 28.1 Origin of the Term Oxidation

There are no worksheets for this lesson.

28.2 Lesson 28.2 Oxidation-Reduction

There are no worksheets for this lesson.

28.3 Lesson 28.3 Balancing Redox Equations Using the Oxidation Number Method

Balancing Redox Equations Worksheet

CK-12 Foundation Chemistry

Name_______________________________ **Date**____________

Steps in the balancing redox equations process.

1. Determine the oxidation number for all atoms in the reaction.
2. Determine which atom is being oxidized and which is being reduced.
3. Write a half-reaction for the reduction process, showing the species containing the atom being reduced and the product containing that atom.
4. Write a half-reaction for the oxidation process, showing the species containing the atom being oxidized and the product containing that atom.

5. If the atoms being oxidized and reduced are not already balanced in the half-reactions, balance them.
6. Add the appropriate number of electrons to each half-reaction needed to bring about the reduction and oxidation.
7. Balance all other atoms in each half-reaction except H and O.
8. Balance the H and O according to either (a) or (b) depending on whether the reaction is acidic or basic.
   (a) If the reaction is acidic, add H_2O and H^+. Balance O first by adding H_2O, then balance H by adding H^+. Charge should now be balanced.
   (b) If the reaction is basic, add OH^- and H_2O. Balance charge first by adding OH^-, then balance O by adding H_2O. The H should now be balanced.
9. Once the half-reactions are balanced, find the lowest common multiple (LCM) for the electrons in the two half-reactions.
10. Multiply each half-reaction by a whole number so that the total number of electrons in the reduction half-reaction equals the total number of electrons in the oxidation half-reaction, and they each equal the LCM.
11. Add the two half-reactions and cancel those species that are common to both sides.
12. Check the equation to be sure that it is balanced by both atoms and charge.

Example of an acidic redox reaction balancing.

Given skeleton: $MnO_4^- \ + \ C_2O_4^{2-} \ \rightarrow \ Mn^{2+} \ + \ CO_2$ (in acid)

Step 1: $\underset{+7\ -2}{MnO_4^-} \ + \ \underset{+3\ -2}{C_2O_4^{2-}} \ \rightarrow \ \underset{+2}{Mn^{2+}} \ + \ \underset{+4\ -2}{CO_2}$ (in acid solution)

Step 2: Mn^{+7} is being reduced to Mn^{+2} and C^{+3} is being oxidized to C^{+4}.

Step 3: $MnO_4^- \ \rightarrow \ Mn^{2+}$

Step 4: $C_2O_4^{2-} \ \rightarrow \ CO_2$

Step 5: $MnO_4^- \ \rightarrow \ Mn^{2+}$ and $C_2O_4^{2-} \ \rightarrow \ 2\ CO_2$

Step 6:

$$MnO_4^- \ + \ 5\ e^- \ \rightarrow \ Mn^{2+}$$
$$C_2O_4^{2-} \ \rightarrow \ 2\ CO_2 \ + \ 2\ e^-$$

Step 7: All atoms other than H and O are balanced.

Step 8a: $MnO_4^- \ + \ 5\ e^- \ + \ 8\ H^+ \ \rightarrow \ Mn^{2+} \ + \ 4\ H_2O$

Step 8a: $C_2O_4^{2-} \ \rightarrow \ 2\ CO_2 \ + \ 2\ e^-$

Step 9: The lowest common multiple for the electrons is 10. Therefore, we will multiply the reduction half-reaction by 2 and the oxidation half-reaction by 5.

Step 10: $2\ MnO_4^- \ + \ 10\ e^- \ + \ 16\ H^+ \ \rightarrow \ 2\ Mn^{2+} \ + \ 8\ H_2O$

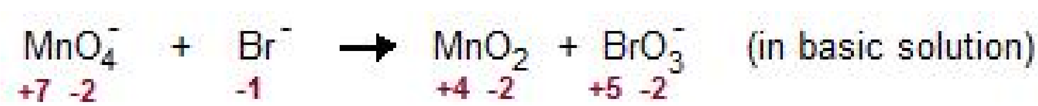
Step 10: $5\ C_2O_4^{2-} \ \rightarrow \ 10\ CO_2 \ + \ 10\ e^-$

Step 11 and 12: $2\ MnO_4^- \ + \ 16\ H^+ \ + \ 5\ C_2O_4^{2-} \ \rightarrow \ 2\ Mn^{2+} \ + \ 8\ H_2O \ + \ 10\ CO_2$

Example of an basic redox reaction balancing.

Given skeleton: $MnO_4^- \ + \ Br^- \ \rightarrow \ MnO_2 \ + \ BrO_3^-$ (in basic solution)

Step 1:



Step 2: Mn^{+7} is being reduced to Mn^{+4} and Br^- is being oxidized to Br^{+5}.

Step 3: $MnO_4^- \ \rightarrow \ MnO_2$

Step 4: $Br^- \ \rightarrow \ BrO_3^-$

Step 5: Both the atoms being oxidized and the atoms being reduced are balanced in the half-reactions.

Step 6: $MnO_4^- \ + \ 3\ e^- \ \rightarrow \ MnO_2$ and $Br^- \ \rightarrow \ BrO_3^- \ + \ 6\ e^-$

Step 7: All atoms other than H and O are balanced.

Step 8b: $MnO_4^- \ + \ 3\ e^- \ + \ 2\ H_2O \ \rightarrow \ MnO_2 \ + \ 4\ OH^-$

Step 8b: $Br^- \ + \ 6\ OH^- \ \rightarrow \ BrO_3^- \ + \ 6\ e^- \ + \ 3\ H_2O$

Step 9: The LCM for the electrons is 6. Therefore, we will multiply the reduction half-reaction by 2 and the oxidation half-reaction by 1.

Step 10: $2\ MnO_4^- \ + \ 6\ e^- \ + \ 4\ H_2O \ \rightarrow \ 2\ MnO_2 \ + \ 8\ OH^-$

Step 10: $Br^- \ + \ 6\ OH^- \ \rightarrow \ BrO_3^- \ + \ 6\ e^- \ + \ 3\ H_2O$

Steps 11 and 12 (Cancel electrons, H_2O, and OH^-):

$2\ MnO_4^- \ + \ Br^- \ + \ H_2O \ \rightarrow \ 2\ MnO_2 \ + \ 2\ OH^- \ + \ BrO_3^-$

Exercises

Balance the following redox equations.

1. $Br_2 \ + \ SO_2 \ \rightarrow \ Br^- \ + \ HSO_4^-$ (in acidic solution)
2. $PbO_2 \ + \ Mn^{2+} \ \rightarrow \ Pb^{2+} \ + \ MnO_4^-$ (in acidic solution)

3. $MnO_4^- + SO_3^{2-} \rightarrow MnO_2 + SO_4^-$ (in basic solution)
4. $Zn + NO_3^- \rightarrow NH_3 + Zn(OH)_4^{2-}$ (in basic solution)
5. $H_2O_2 + Cl_2O_7 \rightarrow ClO_2^- + O_2$ (in basic solution)

28.4 Lesson 28.4 Electrolysis

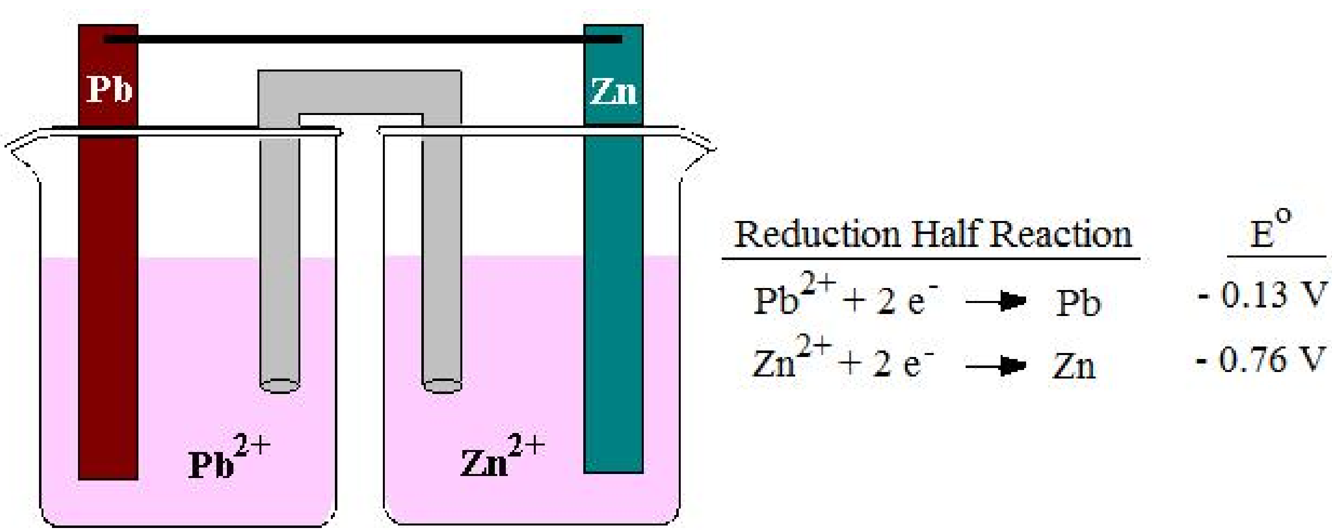
There are no worksheets for this lesson.

28.5 Lesson 28.5 Galvanic Cells

Electrochemical Cells Worksheet

CK-12 Foundation Chemistry

Name____ **Date**________________



Use the standard cell sketched above to answer questions 1 - 9.

1. Which electrode is the cathode?

A. Pb
B. Zn
C. Neither.

2. Which electrode is the anode?

A. Pb
B. Zn
C. Neither.

3. At which electrode will oxidation occur?

A. Pb
B. Zn
C. Neither.

4. What is the maximum voltage for this standard cell?

A. 0.89 V
B. 0.63 V
C. -0.89 V
D. -0.63 V
E. 0.50 V

5. Which way do the electrons flow in the external circuit?

A. From Pb to Zn.
B. From Zn to Pb.
C. No electron flow occurs.

6. Which way do cations flow through the salt bridge?

A. Toward the Pb electrode.
B. Toward the Zn electrode.
C. No cation flow occurs.

7. What happens to the cell voltage when the reaction reaches equilibrium?

A. Becomes maximum.
B. Drops to zero.
C. Becomes a positive value less than maximum.

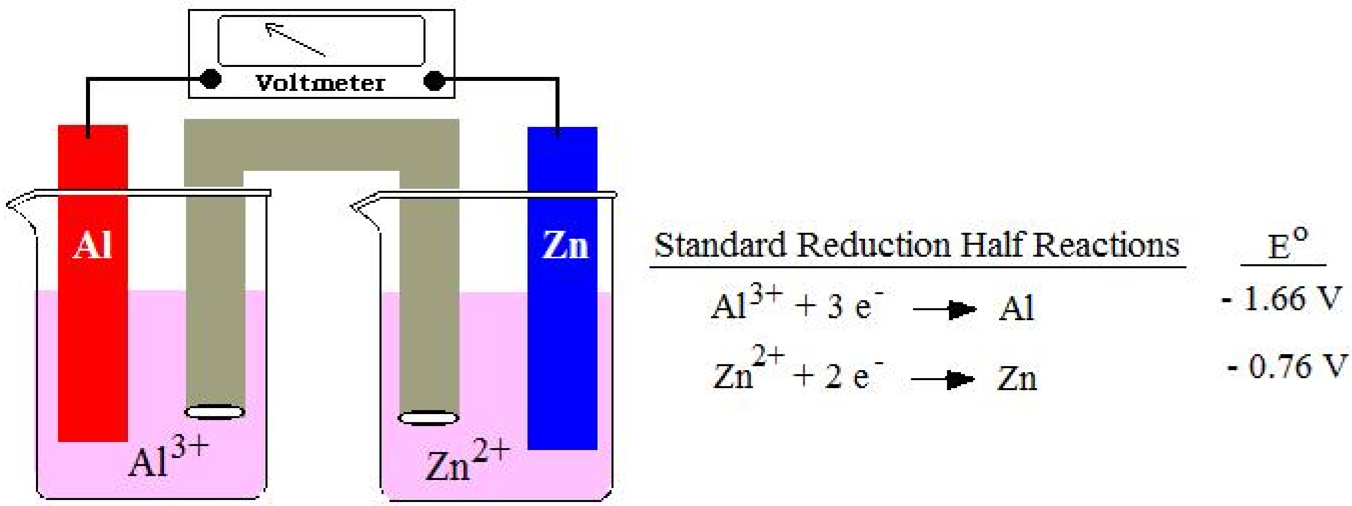
8. Which electrode will gain mass as the cell runs?

A. Pb

B. Zn
C. Neither.

9. What happens to the cell voltage as the cell runs?

A. Remains constant.
B. Increases.
C. Decreases.
D. May increase or decrease.



Use the standard cell sketched above to answer questions 10 - 21.

10. Which electrode is the anode?

A. Al
B. Zn
C. Neither.

11. At which electrode does reduction occur?

A. Al
B. Zn
C. Neither.

12. What is the voltage of this standard cell?

A. 2.42 V
B. -2.42 V
C. 0.90 V
D. -0.90 V
E. 1.80 V

13. Which way do the electrons flow in the external circuit?

A. From Al to Zn.
B. From Zn to Al.
C. No electron flow occurs.

14. Which way do anions flows through the salt bridge?

A. Toward the Al electrode.
B. Toward the Zn electrode.
C. No cation flow occurs.

15. Which electrode loses mass as the cell runs?

A. Al
B. Zn
C. Neither.

16. How many moles of electrons pass through the external circuit in order for 1.00 mole of atoms to be deposited on the cathode?

A. 6
B. 3
C. 4
D. 2
E. 1

17. If 24 electrons pass through the external circuit, how many atoms of zinc must react?

A. 24
B. 12
C. 8
D. 4
E. 0

18. If 24 electrons pass through the external circuit, how many atoms of aluminum must react?

A. 24
B. 12
C. 8
D. 4
E. 0

19. What will happen to the voltage of the cell if the molarity of Zn^{2+} is increased?

A. Increase.
B. Decrease.
C. Remain the same.

20. What will happen to the voltage of the cell if the molarity of Al^{3+} is increased?

A. Increase.
B. Decrease.
C. Remain the same.

21. What will happen to the voltage of the cell if the salt bridge is removed?

A. Increase slightly.
B. Decrease slightly.
C. Remain the same.
D. Drop to zero.

22. In the two cells in this worksheet, there are a total of three reduction half-reaction indicated, Al, Zn, and Pb. Which of these three metals is most easily oxidized?

A. Al
B. Zn
C. Pb

23. Will a reaction occur if aluminum metal is placed in a solution of Zn^{2+}?

A. Yes
B. No

24. Will a reaction occur if Pb metal is placed in a solution of Al^{3+}?

A. Yes
B. No

25. Will a reaction occur if aluminum metal is placed in a solution of Zn^{2+}?

A. Yes
B. No

Chapter 29

Nuclear Chemistry Worksheets

29.1 Lesson 29.1 Discovery of Radioactivity

There are no worksheets for this lesson.

29.2 Lesson 29.2 Nuclear Notation

There are no worksheets for this lesson.

29.3 Lesson 29.3 Nuclear Force

There are no worksheets for this lesson.

29.4 Lesson 29.4 Nuclear Disintegration

There are no worksheets for this lesson.

29.5 Lesson 29.5 Nuclear Equations

Nuclear Chemistry Worksheet

CK-12 Foundation Chemistry

Name________________________________ **Date**_____________

In questions 1 - 5, a single nuclear particle is missing. Fill-in the complete nuclear symbol for the missing particle.

1.

$$^{28}_{13}\mathrm{A}l \rightarrow ^{26}_{12}\mathrm{Mg} + ?$$

2.

$$^{210}_{84}\mathrm{P}o \rightarrow ^{210}_{85}\mathrm{At} + ?$$

3.

$$^{209}_{83}\mathrm{B}i \rightarrow ^{4}_{2}\mathrm{He} + ?$$

4.

$$^{242}_{96}\mathrm{C}m + ^{12}_{6}\mathrm{C} \rightarrow 3\ ^{1}_{0}\mathrm{n} + ?$$

5.

$$^{223}_{87}\mathrm{F}r + ? \rightarrow ^{226}_{88}\mathrm{R}a + ^{1}_{1}\mathrm{H}$$

6. Fill-in the following table with the mass number and the charge of the particles.

Table 29.1:

Particle	**Mass Number**	**Charge**
neutron		
proton		
electron		
alpha particle		
U-235 nucleus		

7. An isotope of bismuth, Bi-209, is bombarded with a proton. The product of the reaction is an isotope of element X and two neutrons. What is the mass number of this isotope of element X?

A. 206
B. 207
C. 208
D. 209
E. 210

8. Which of the following particles completes this equation?

$^{238}U \ + \ ^{4}He \ \rightarrow \ ^{241}Pu \ + \ ?$

A. Beta.
B. Alpha.
C. Proton.
D. Neutron.
E. None of these.

9. Which of the following particles completes this equation?

$^{241}Pu \ \rightarrow \ ^{241}Am \ + \ ?$

A. Beta.
B. Alpha.
C. Proton.
D. Neutron.
E. None of these.

10. Which of the following particles completes this equation?

$^{10}B \ \rightarrow \ ^{6}Li \ + \ ?$

A. Beta.
B. Alpha.
C. Proton.
D. Neutron.
E. None of these.

11. An atom contains 3 protons, 4 neutrons, and 3 electrons. What is its mass number?

A. 3

B. 6
C. 7
D. 10
E. None of these.

12. If Th-234 undergoes beta decay, the resultant particle will be ________________.

A. Ra-234
B. Th-230
C. Pa-234
D. U-235
E. None of these.

13. U-234 undergoes alpha decay and the resultant particle undergoes beta decay. What is the final particle after both decays?

A. Np-236
B. Pa-230
C. Ac-232
D. Np-239
E. Pa-233

14. 20.0 grams of a radioactive element is prepared in a nuclear reactor. The half-life of the isotope is 3 days. How many days will it take before there is only 2.50 grams of the substance remaining?

A. 1.5 days
B. 3 days
C. 6 days
D. 9 days
E. 12 days

15. Element X has only two isotopes. One of the isotopes has a mass number of 190 and the other has a mass number of 194. If the atomic mass of element X is 193.6, which of the two isotopes is most commonly found in nature?

A. 190
B. 193.6
C. 194
D. The two isotopes are equally common.
E. Insufficient data to determine.

29.6 Lesson 29.6 Radiation Around Us

There are no worksheets for this lesson.

29.7 Lesson 29.7 Applications of Nuclear Energy

There are no worksheets for this lesson.

Chapter 30

Organic Chemistry Worksheets

30.1 Lesson 30.1 Carbon, A Unique Element

There are no worksheets for this lesson.

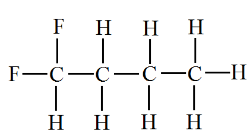
30.2 Lesson 30.2 Hydrocarbons

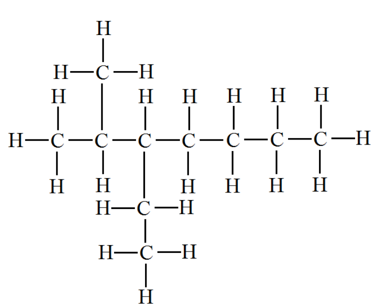
Organic Nomenclature Worksheet

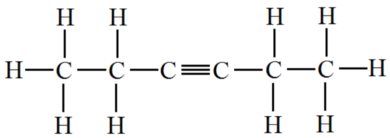
CK-12 Foundation Chemistry

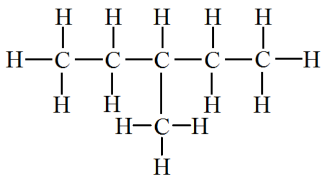
Name____________________________________ **Date**_______________

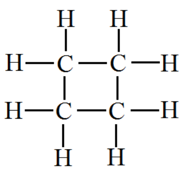
Name the following molecules.

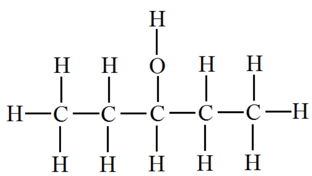
1.  ____

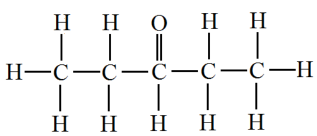
2.  ____

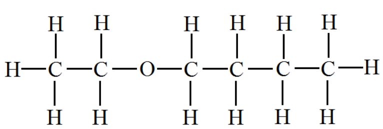
3. 

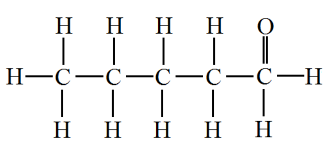
4. 

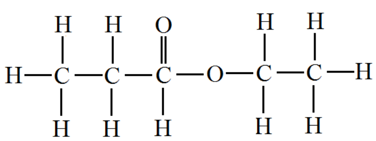
5. 

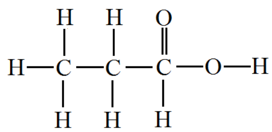
6. 

7. 

8. 

9. 

10. 

11.  ____

Draw the following molecules.

12. 1-Butyne
13. Methoxyethane
14. Butanal
15. 1,2-Dibromopropane

30.3 Lesson 30.3 Aromatics

There are no worksheets for this lesson.

30.4 Lesson 30.4 Functional Groups

Have students continue with the Organic Nomenclature worksheet started in lesson 30.2 .

30.5 Lesson 30.5 Biochemical Molecules

There are no worksheets for this lesson.